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JACARANDA
CHEMISTRY 1

VCE UNITS 1 AND 2 | SECOND EDITION

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CONTENTS

AREA OF STUDY 2 HOW CAN THE VERSATILITY OF NON-METALS BE EXPLAINED?

19 Analysis for acids and bases in water 522

AREA OF STUDY 2 REVIEW

AREA OF STUDY 3 PRACTICAL INVESTIGATION

20 Practical investigation 558

ABOUT THIS RESOURCE

Jacaranda Chemistry has been reimagined to provide students and teachers with the most comprehensive resource on the market. This engaging and purposeful suite of resources is fully aligned to the VCE Chemistry Study Design (2016–2021).

Formats

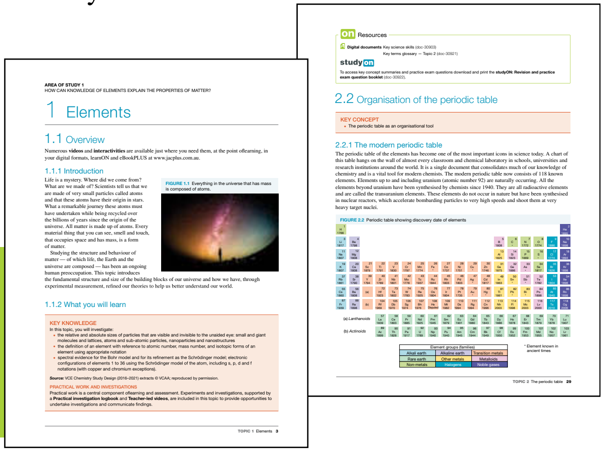
Jacaranda Chemistry is now available in print and a range of digital formats, including:

Print	**learnON**	**eBookPLUS**	**PDF**	**eGuidePLUS**
Printed textbook with free digital access code inside	learnON is our immersive and flexible digital learning platform	The eBookPLUS is an electronic version of the student text	Downloadable PDFs available with eBookPLUS	The eGuidePLUS includes everything from the eBookPLUS with additional resources designed for teachers

Fully aligned to the VCE Chemistry Study Design

Have confidence that you are covering the entire VCE Chemistry Study Design (2016–2021), with:

- tailored exercise sets at the end of each sub-topic
- additional background information easily distinguished from curriculum content
- practice SACs clearly linked to each outcome
- practice exams for each area of study.



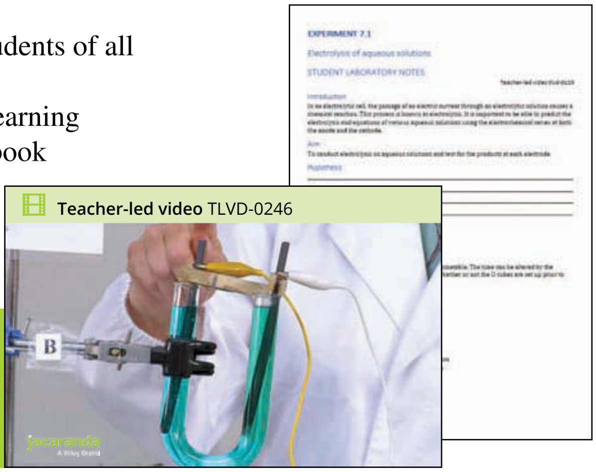
Key concepts approach

Students can easily understand which aspect of the Study Design they are studying with Key Knowledge dot points highlighted at the start of each topic and broken down into Key Concepts in each sub-topic.

Engaging resources and rich media

A variety of online and offline resources to encourage students of all learning styles, including:

- videos and interactivities embedded at the point of learning
- practical investigations complete with printable logbook
- a variety of worksheets and activities allow students to apply their knowledge of the content.



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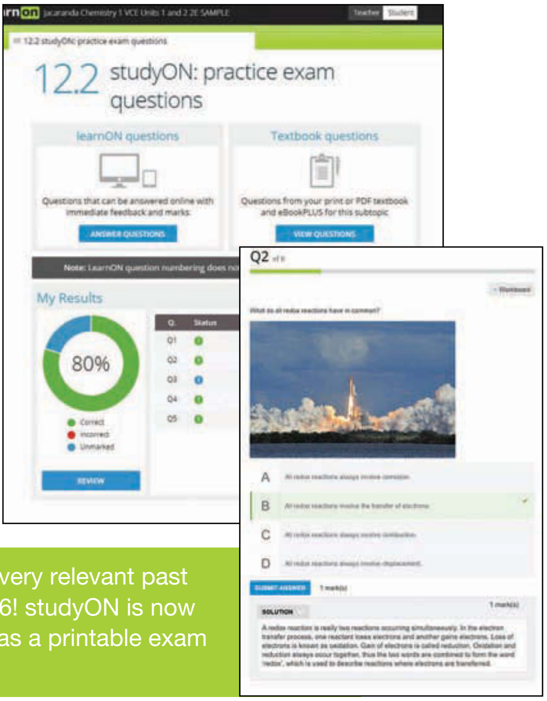
Videos of both sample problems and practical investigations led by experienced teachers allow students to better consolidate their learning.

Inspiring students to become independent learners

This resource, on the immersive digital platform learnON, encourages self-driven student learning through:

- fully worked solutions for every question providing students with immediate feedback
- progress tracked automatically in learnON allowing students to reflect on their learning, and enabling students to highlight areas of strength and weakness
- topic summary, key terms and practice questions at the end of every topic
- sample problems with fully worked solutions set out in the THINK–WRITE format.



12.2 studyON: practice exam questions

learnON questions	Textbook questions
Questions that can be answered online with immediate feedback and marks.	Questions from your print or PDF textbook and eBookPLUS for this subtopic
ANSWER QUESTIONS	VIEW QUESTIONS

Note: LearnON question numbering does no

My Results

80%

- 🟢 Correct
- 🔴 Incorrect
- 🔵 Unmarked

Q.	Status
Q1	🟢
Q2	🟢
Q3	🔵
Q4	🟢
Q5	🟢

What do all redox reactions have in common?

A All redox reactions always involve corrosion.

B All redox reactions involve the transfer of electrons. ✓

C All redox reactions always involve combustion.

D All redox reactions always involve displacement.

A redox reaction is really two reactions occurring simultaneously. In the electron transfer process, one reactant loses electrons and another gains electrons. Loss of electrons is known as oxidation. Gain of electrons is called reduction. Oxidation and reduction always occur together, thus the two words are combined to form the word 'redox', which is used to describe reactions where electrons are transferred.

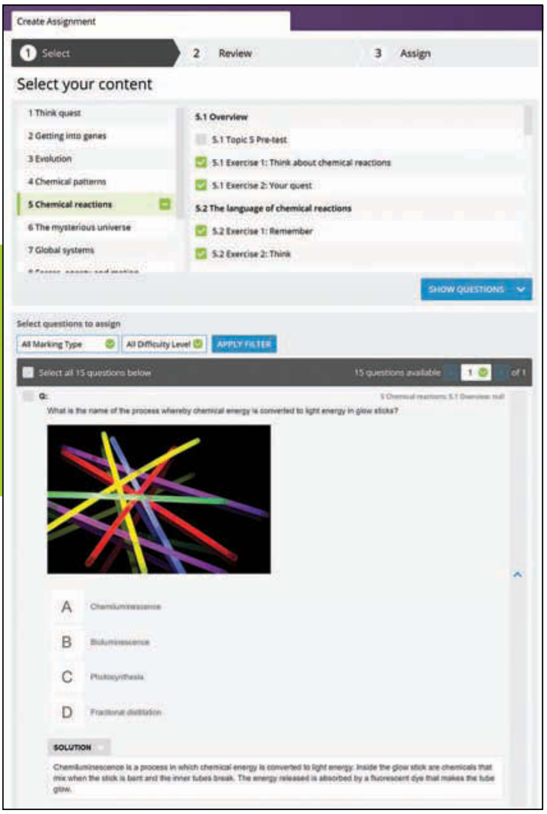
studyon

Prepare for exams with every relevant past exam question since 2006! studyON is now included in learnON and as a printable exam revision booklet.

A wealth of teacher resources

Jacaranda Chemistry empowers teachers to teach their class their way with the extensive range of teacher resources including:

- quarantined topic tests and SACs that are easily customisable
- practical investigation support with demonstrative videos, laboratory information, expected results and risk assessments
- work programs
- curriculum grids.



Select your content

- 1 Think quest
- 2 Getting into genes
- 3 Evolution
- 4 Chemical patterns
- 5 Chemical reactions
- 6 The mysterious universe
- 7 Global systems
- 8 Forces, energy and motion

5.1 Overview

- ☐ 5.1 Topic 5 Pre-test
- ☑ 5.1 Exercise 1: Think about chemical reactions
- ☑ 5.1 Exercise 2: Your quest

5.2 The language of chemical reactions

- ☑ 5.2 Exercise 1: Remember
- ☑ 5.2 Exercise 2: Think

Select questions to assign

☐ Select all 15 questions below — 15 questions available 1 of 1

☐ Q: 5 Chemical reactions: 5.1 Overview: null

What is the name of the process whereby chemical energy is converted to light energy in glow sticks?

A Chemiluminescence

B Bioluminescence

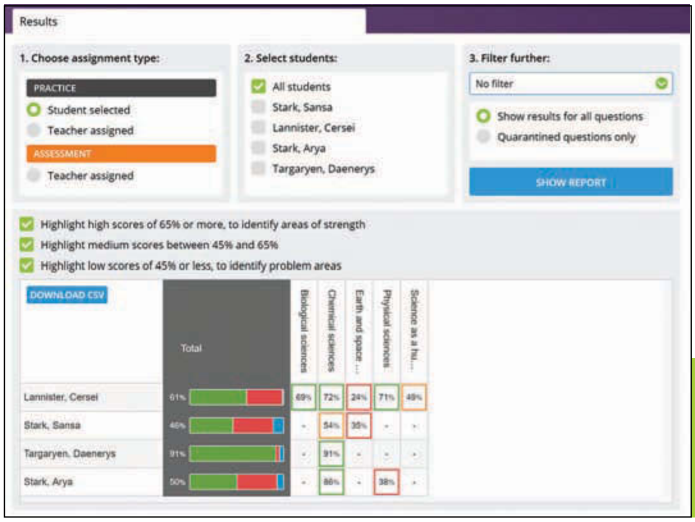
C Photosynthesis

D Fractional distillation

Chemiluminescence is a process in which chemical energy is converted to light energy. Inside the glow stick are chemicals that mix when the stick is bent and the inner tubes break. The energy released is absorbed by a fluorescent dye that makes the tube glow.

Test Maker

Create custom tests for your class from the entire question pool — including all subtopic, topic review and past VCAA exam questions



1. Choose assignment type:

PRACTICE
- ● Student selected
- ○ Teacher assigned

ASSESSMENT
- ○ Teacher assigned

2. Select students:

- ☑ All students
- ☐ Stark, Sansa
- ☐ Lannister, Cersei
- ☐ Stark, Arya
- ☐ Targaryen, Daenerys

3. Filter further:

No filter

- ● Show results for all questions
- ○ Quarantined questions only

- ☑ Highlight high scores of 65% or more, to identify areas of strength
- ☑ Highlight medium scores between 45% and 65%
- ☑ Highlight low scores of 45% or less, to identify problem areas

	Total	Biological sciences	Chemical sciences	Earth and space ...	Physical sciences	Science as a hu...
Lannister, Cersei	61%	69%	72%	24%	71%	49%
Stark, Sansa	46%	-	54%	35%	-	-
Targaryen, Daenerys	91%	-	91%	-	-	-
Stark, Arya	50%	-	86%	-	38%	-

Visibility of student results

Detailed breakdown of student results allows you to identify strengths and weaknesses across various topics and sub-topics.

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UNIT 1
HOW CAN THE DIVERSITY OF MATERIALS BE EXPLAINED?

Everything in the universe is composed of elements — from the furthest star, to the icing on your birthday cake, to the nanomaterials used to target individual cancer cells.

Understanding the properties of elements and their atomic structure allows an exploration of the relationships between them. These relationships relate to the properties of individual atoms, which in turn shape the bonding forces within and between particles. The periodic table is the primary tool used by chemists to predict the behaviour of elements, and the compounds they form with other elements.

To consider the properties of elements and the compounds they form, quantitative concepts are needed. These include the atomic number which identifies an element, the relative masses of elements and how these vary with isotopes, and the mole concept as the base unit of chemistry, which allows comparison of amounts of different substances. Chemical understanding is often gained through observations and data collection in experiments. To communicate this understanding, specialised terminology is required, including symbols, formulas, chemical nomenclature and equations.

Understanding how elements combine to form molecules is a fundamental and important skill. It allows us to explain the incredible diversity of materials in the world.

AREA OF STUDY	OUTCOME	TOPICS
1. How can knowledge of elements explain the properties of matter?	Relate the position of elements in the periodic table to their properties, investigate the structures and properties of metals and ionic compounds, and calculate mole quantities.	1. Elements 2. The periodic table 3. Metals 4. Ionic compounds 5. Quantifying atoms and compounds
2. How can the versatility of non-metals be explained?	Investigate and explain the properties of carbon lattices and molecular substances with reference to their structures and bonding, use systematic nomenclature to name organic compounds, and explain how polymers can be designed for a purpose.	6. Materials from molecules 7. Carbon lattices and carbon nanomaterials 8. Organic compounds 9. Polymers
3. Research investigation	Investigate a question related to the development, use and/or modification of a selected material or chemical and communicate a substantiated response to the question.	10. Research investigation

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AREA OF STUDY 1
HOW CAN KNOWLEDGE OF ELEMENTS EXPLAIN THE PROPERTIES OF MATTER?

1 Elements

1.1 Overview

Numerous **videos** and **interactivities** are available just where you need them, at the point of learning, in your digital formats, learnON and eBookPLUS at www.jacplus.com.au.

1.1.1 Introduction

Life is a mystery. Where did we come from? What are we made of? Scientists tell us that we are made of very small particles called atoms and that these atoms have their origin in stars. What a remarkable journey these atoms must have undertaken while being recycled over the billions of years since the origin of the universe. All matter is made up of atoms. Every material thing that you can see, smell and touch, that occupies space and has mass, is a form of matter.

Studying the structure and behaviour of matter — of which life, the Earth and the universe are composed — has been an ongoing human preoccupation. This topic introduces the fundamental structure and size of the building blocks of our universe and how we have, through experimental measurement, refined our theories to help us better understand our world.

FIGURE 1.1 Everything in the universe that has mass is composed of atoms.

1.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- the relative and absolute sizes of particles that are visible and invisible to the unaided eye: small and giant molecules and lattices, atoms and sub-atomic particles, nanoparticles and nanostructures
- the definition of an element with reference to atomic number, mass number, and isotopic forms of an element using appropriate notation
- spectral evidence for the Bohr model and for its refinement as the Schrödinger model; electronic configurations of elements 1 to 36 using the Schrödinger model of the atom, including s, p, d and f notations (with copper and chromium exceptions).

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of learning and assessment. Experiments and investigations, supported by a **Practical investigation logbook** and **Teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

on Resources

Digital documents Key science skills (doc-30903)
Key terms glossary — Topic 1 (doc-30918)
Practical investigation logbook (doc-30919)

studyon

To access key concept summaries and practice exam questions download and print the **studyON: Revision and practice exam question booklet** (doc-30920).

1.2 Atomic structure

KEY CONCEPT

- The relative and absolute sizes of particles that are visible and invisible to the unaided eye: small and giant molecules and lattices, atoms and sub-atomic particles, nanoparticles and nanostructures

1.2.1 The structure of atoms

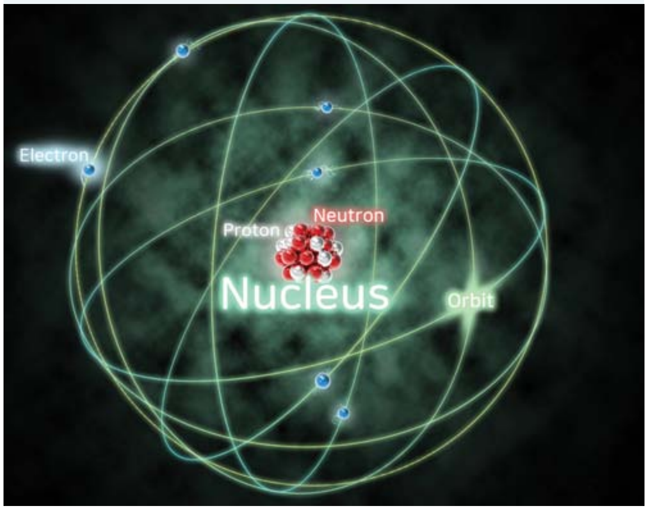
Atomic theory attempts to explain the structure of materials. According to this theory, all matter is made of **atoms**. Atoms are so small that it was not until 1981 that their images could finally be seen using the newly invented scanning tunnelling microscope. Due to their incredibly small size, models have developed to represent the internal structure of atoms.

One very useful model is the nuclear model of the atom proposed by Rutherford in 1911.

Rutherford's descriptions of an atom include:

- An atom is mostly empty space.
- An atom has a dense central structure called a nucleus.
- The nucleus, though its volume is very small relative to the atom as a whole, contains most of the mass of the atom.
- The nucleus is made up of positively charged particles called protons.
- The simplest nucleus is that of the hydrogen atom, which contains just one proton.
- All other atoms have nuclei that also contain neutrons.
- A neutron has no charge but has virtually the same mass as a proton.
- The empty space around the nucleus contains negatively charged particles called electrons.
- Electrons move very rapidly around the nucleus in orbits.
- Each electron has a definite energy and moves in a specific energy level. The mass of an electron is very much less than that of a proton or a neutron.

FIGURE 1.2 A modern take on New Zealand–born Ernest Rutherford's nuclear model of an atom.

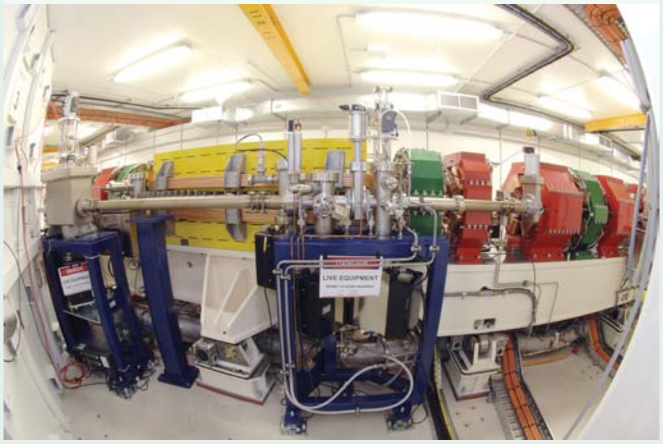


on Resources

Video eLesson Rutherford's gold foil experiment (eles-2486)

AUSTRALIA'S PARTICLE ACCELERATOR

FIGURE 1.3 The Australian Synchrotron



The Australian Synchrotron is a particle accelerator in Victoria that is used by scientists to investigate the structure of matter. With over 4000 research visits per year, the fields of investigation are diverse and include agricultural science, environmental science, minerals analysis, medical investigations, materials science, cultural heritage, nanotechnology and forensics.

From the outside, the Australian Synchrotron resembles a football stadium. Inside, however, instead of footballs going in different directions, electrons are accelerated around a large loop (with a circumference of 216 metres) at almost the speed of light. The light is produced by high-energy electrons that are deflected into circular orbit by the 'synchronised' application of strong magnetic fields. The light produced is 1 million times brighter than the sun. The light, X-rays and infrared radiation produced is directed to a number of experimental workstations where many different experiments take place.

Protons, electrons and neutrons are called **subatomic particles**. Table 1.1 summarises the properties of these particles. Atoms that are neutrally charged have the same number of electrons and protons. If an atom gains or loses an electron, it becomes charged and is then called an **ion**. You will study ions and their behaviour in more detail in later topics.

TABLE 1.1 Particles in an atom and their properties

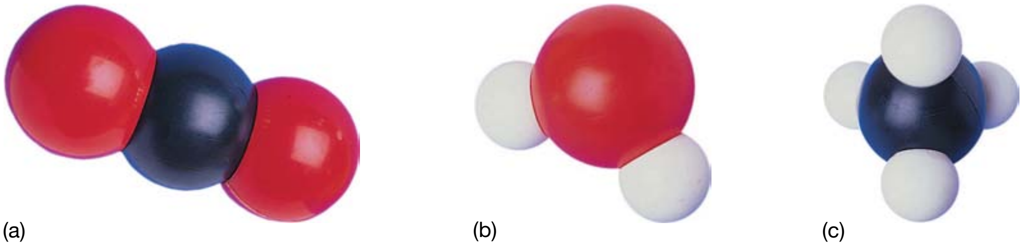
Subatomic particle	Relative mass	Relative charge
Electron	$\frac{1}{1837} = 0.0005$	−1
Proton	1	+1
Neutron	1	0

Atoms are not all the same. To date, chemists have identified 118 different types of atom. Some substances contain only one type of atom. These substances are called **elements**. For example, the element oxygen contains only oxygen atoms and pure lead contains only lead atoms. The atoms of each element have special characteristics, and these are used to classify them. Very few elements exist as individual atoms; examples are helium and neon. Many more substances consist of two or more atoms that are chemically combined.

These are called **molecules**.

Molecules can be represented by space-filling models as shown in figure 1.4. These models show the relative sizes of the centre-to-centre distances between the atoms and are useful in showing the dimensions of a molecule. They do not show the bonds involved.

FIGURE 1.4 Molecules of (a) carbon dioxide, CO_2, (b) water, H_2O, and (c) methane, CH_4 represented by space-filling models.



1.2.2 A matter of size

A matter of size

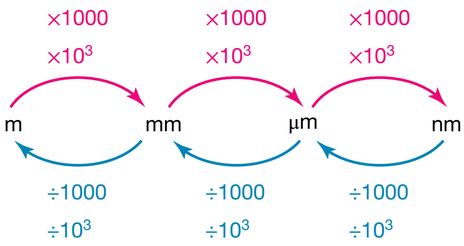
Understanding the dimensions of atoms can be quite a challenge. If the nucleus in the atom in figure 1.4 was 5 millimetres in diameter, the outer edge of the atom would be 50 000 millimetres or 50 metres away. When discussing the sizes of atoms, a more useful unit is the **nanometre**. In reality, the diameter of a hydrogen nucleus (a single proton) is 1.7566×10^{-15} m.

Nano- is a prefix like *milli*- and *micro*-:

- A millimetre is one thousandth of a metre: mm $= 10^{-3}$ m.
- A micrometre is one millionth of a metre: μm $= 10^{-6}$ m .
- A nanometre is one billionth of a metre: nm $= 10^{-9}$ m.

To compare two objects, their measurements must be converted to the same unit. For example:

FIGURE 1.5 Converting between metres and nanometres.



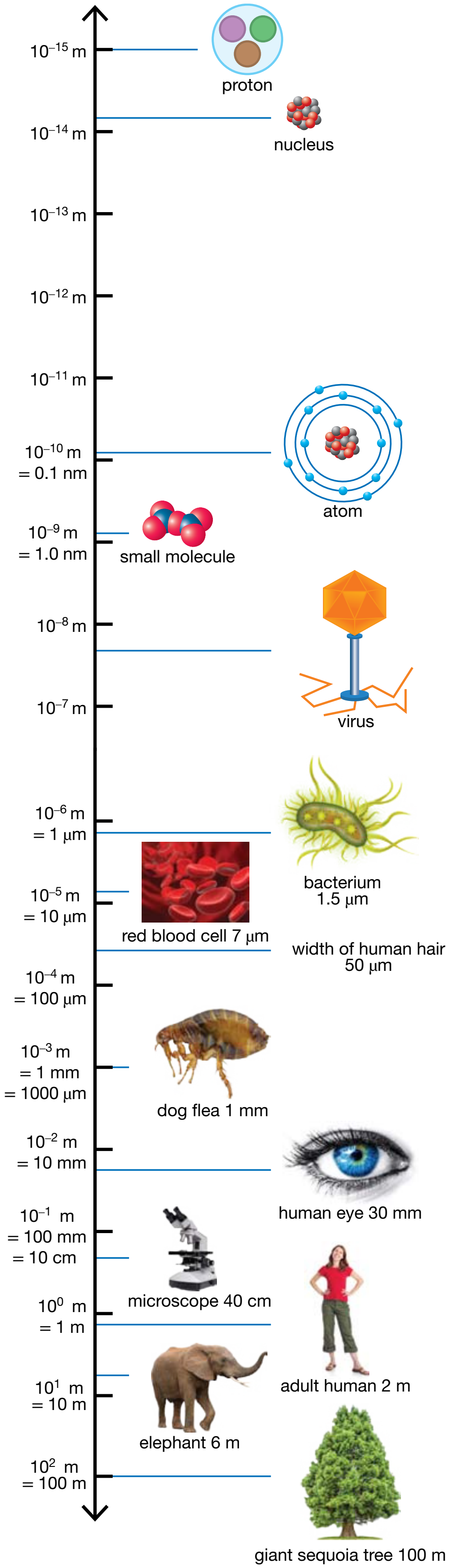
$$1 \text{ m} = 10^2 \text{ cm} = 10^3 \text{ mm} = 10^6\ \mu\text{m} = 10^9 \text{ nm}$$

Nanotechnology takes our ability to control things one step beyond the microworld, to the nanoworld. Special tools are needed to work at the nanoscale. Scientists use scanning probe microscopes (SPM) to make images of individual atoms that are much too small to see. SPMs work by 'feeling' the bumps caused by atoms on a surface. Any technology that involves manipulation, construction or control of objects with a size at the nanometre scale is a branch of nanotechnology. A large atom is about 0.1 nm in diameter. Molecules can range from about 0.15 nm (H_2) to a visible size (macromolecules). Bacteria range in size from about 500 nm to about 5000 nm, and viruses from about 10 to 50 nm. Figure 1.6 compares these scales, putting the nanoscale in context.

on Resources

Weblink Scale of the universe

FIGURE 1.6 Comparing scales — a human is about 2000 times larger than a flea, which in turn is 10 million times larger than an atom.



SAMPLE PROBLEM 1

A dog flea is 1 mm long whereas an African elephant can be 6 m long. How many times longer is the elephant than the flea?

Teacher-led video: SP1 (tlvd-0508)

THINK	WRITE
1. To compare sizes, the units must be the same. Convert both into mm *or* both into m. **TIP:** Table 6 in the VCE Chemistry Data Book has the metric conversion factors.	Convert to mm: $1\text{ m} = 1 \times 10 \times 10\text{ mm}$ $= 1000\text{ mm}$ Therefore, an elephant $= 6\text{ m} \times 1000$ $= 6000\text{ mm}$ *or* Convert to m: $1\text{ mm} = 1 \times 10^{-3}\text{ m}$ Therefore a flea $= 1 \times 10^{-3}\text{ m}$
2. Use the ratio of 1 mm to 6000 mm or 1×10^{-3} m to 6 m	$\frac{6000\text{ mm}}{1\text{ mm}} = 6000$ *or* $\frac{6\text{ m}}{1 \times 10^{-3}\text{ m}} = 6000$ The elephant is 6000 times longer than a flea.

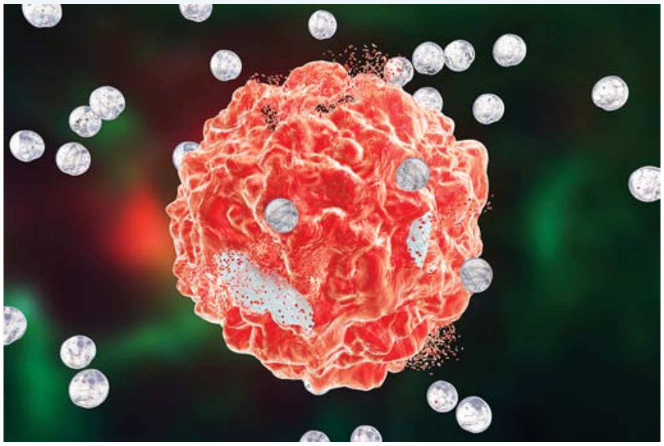
PRACTICE PROBLEM 1

If an atom is 1 nm wide, how many times larger is it than a single proton, which is 10^{-15} m wide?

1.2.3 Responsible use of nanotechnology

Nanoscience is an interdisciplinary field that involves combinations of knowledge from biology, chemistry, physics, engineering and computer science. It deals directly with atoms, molecules or macromolecules. Nanotechnology is likely to benefit many areas of our lives in the future, particularly in the area of improving and extending life and ensuring a cleaner and safer environment to live in. It has applications in items such as sunscreen, fabric protection and surface adhesion. However, potential risks also exist with nanotechnology, such as military use in creating miniaturised weapons and explosives, or in designing biological organisms at molecular levels. Who will own the technology? What risks may be generated as a result of the technology? These and other ethical and social questions have to be raised and discussed by scientists, policymakers and the general public so that we can responsibly develop and use nanotechnology. Metallic nanomaterials will be discussed in more detail in topic 3.

FIGURE 1.7 In the future, could nanoparticles be used to destroy cancer cells?



NANOTECHNOLOGY AND MEDICINE

The study of nanoparticles is dominated by surface effects such as electronic charge, dispersion forces, and hydrogen, ionic, metallic and covalent bonding. For chemists, nanotechnology will enable materials to be assembled literally atom by atom. This idea is referred to as 'molecular manufacturing', which aims to control assembly of individual atoms into specific molecules. This could mean production of materials that have no waste products, no defects in materials, no impurities, and no requirement for machining or assembly because the atoms are placed directly where required to make the final object. Nanotechnology has the potential to build cleaner, safer, longer lasting and smarter products for communication, medicine, transportation, industries, agriculture and the home.

The ability to design molecules atom by atom has immense potential in saving, improving and extending lives. Nanoscale technologies are helping medicine with devices in areas of drug discovery, diagnostics and therapeutics. For example, trials are being undertaken in which microcapsules with nanoscale pores containing pancreatic cells are used so that diabetic patients can obtain insulin without having to inject it. The pores in the microcapsules are big enough to allow insulin to flow out of the capsules and be released into the body but small enough to prevent the entry of harmful antibodies. Other research is developing cancer-fighting drugs that can be adsorbed onto nanoparticles so they can bypass the body's natural defence mechanisms and precisely target cancer cells.

Microarrays and labs-on-chips (also known as microfluidic devices) are tools at the microscale for analysing large numbers of DNA or protein samples at a single time. These are used for diagnosis and drug discovery. Microarrays perform one type of analysis thousands of times, and lab-on-a-chip devices can perform multiple types of analysis at once. Research is now being undertaken into producing nanoarrays, in which thousands of binding sites can be printed onto the area of a single conventional microarray spot. Coupled with computer technology, this will potentially enable the detection of diseases and the discovery of useful drugs to proceed much more quickly and at lower cost.

FIGURE 1.8 Nanoshells with antibodies attached can be delivered through the surface of a cell membrane. The antibodies are attracted to the protein receptor and are loaded into the cell membrane.

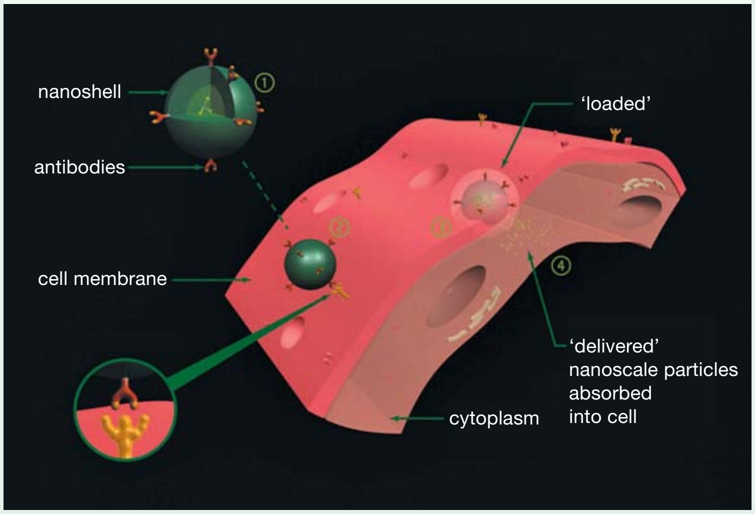
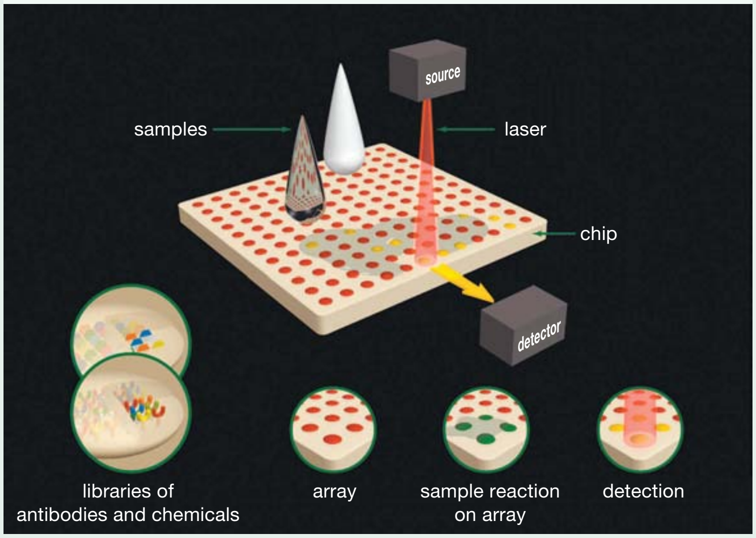


FIGURE 1.9 Samples are placed onto the microarray and may react with the antibodies or chemicals that are placed already onto the microarray surface; the detector then determines if the reaction has occurred.



1.2 EXERCISE

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1. Reorder the following from smallest to largest: bacteria, flea, width of a human hair, DNA molecule, helium atom.
2. A human hair is about 50 μm thick. A pet flea is about 1 mm in size. How many times smaller is the thickness of a human hair than a flea?
3. A red blood cell is about 7 μm in diameter. A virus particle is 50 nm wide. How many times smaller is the virus particle than a red blood cell?

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1.3 Representing atoms

KEY CONCEPT

- The definition of an element with reference to atomic number, mass number, and isotopic forms of an element using appropriate notation

1.3.1 Atomic number (symbol *Z*)

Each of the 118 elements known to chemists has its own **atomic number**. The atomic number (symbol Z) of an element is defined as the number of protons in the nucleus of an atom of that element. When an atom is neutrally charged, the atomic number of the atom corresponds to the number of electrons, because the number of positive charges must be the same as the number of negative charges. For example, oxygen has an atomic number of 8 and, therefore, has 8 protons and 8 electrons.

1.3.2 Mass number (symbol *A*)

Protons have approximately the same mass as neutrons. The electron's mass is negligible. Therefore, the mass of an atom depends on the number of particles in the nucleus. The **mass number** (symbol A) is defined as the total number of protons and neutrons in an atom of an element.

We can use the periodic table to identify each element by its atomic number. The relative atomic mass of each element is also shown on the table (this is discussed in topic 2).

An element is commonly represented as follows:

$$\begin{array}{l} \textbf{mass number} \longrightarrow \\ \textbf{atomic number} \longrightarrow \end{array} {}^{A}_{Z}\boldsymbol{E} \longleftarrow \textbf{symbol for element}$$

This is known as the **isotopic symbol** of an element. We can determine the number of neutrons in an atom by subtracting the atomic number, Z, from the mass number, A. For example, sodium, Na, has atomic number 11 and mass number 23, and can be represented as ${}^{23}_{11}\text{Na}$. An atom of sodium, therefore, has 11 protons and 12 neutrons.

1.3.3 Isotopes

FIGURE 1.10 The oldest reliably dated rock art in Australia is 28 000 years old.

All atoms of a particular element contain the same number of protons. However, atoms of the same element can contain different numbers of neutrons, and these atoms are called **isotopes**. Isotopes have similar chemical properties because their electron structure is the same. They do, however, have different physical properties due to their different masses.

Naturally occurring oxygen consists of three isotopes: $^{16}_{8}O$, $^{17}_{8}O$ and $^{18}_{8}O$. Isotopes are named by their element name followed by their mass number to distinguish them; for example, the isotopes of oxygen are oxygen-16, oxygen-17 and oxygen-18. Aluminium has only one isotope, aluminium-27, $^{27}_{13}Al$.

Isotopes can be used to date archaeological and geological features. Radiocarbon dating using the decay of carbon-14 isotopes is used to date organic material within (or nearby) Indigenous Australian rock art, for example.

SAMPLE PROBLEM 2

a. Write the symbols for the atoms nitrogen-14 and nitrogen-15.
b. How many protons does each atom have?
c. How many neutrons does each atom have?
d. What are these atoms called?
e. Write the complete symbol/quantities for each isotope.

Teacher-led video: SP2 (tlvd-0509)

THINK	WRITE
a. The symbol will not change, regardless of the atomic masses.	N
b. The number of protons of an element will not change, regardless of the atomic masses. The atomic number of N is 7, hence it has 7 protons.	7
c. The number of neutrons can be determined using the following: Number of neutrons = mass number (*Z*) – atomic number (*A*)	nitrogen-14: 14 – 7 = 7 nitrogen-15: 15 – 7 = 8
d. Name for atoms with the same number of protons but with different masses (different number of neutrons)	Isotopes
e. This must include the symbol, mass and atomic numbers in the correct format.	nitrogen-14: $^{14}_{7}N$ nitrogen-15: $^{15}_{7}N$

PRACTICE PROBLEM 2

a. **Write the isotopic symbol for the atoms of hydrogen-1, hydrogen-2 and hydrogen-3.**
b. **How many protons does each atom have?**
c. **How many neutrons does each atom have?**
d. **Write the complete symbol/quantities for each isotope.**

1.3 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Look up your periodic table to find the atomic number of each of the following elements.
   - (a) H
   - (b) Ne
   - (c) Ag
   - (d) Au
2. An atom has 13 protons and 14 neutrons. What is its
   - (a) atomic number?
   - (b) mass number?
   - (c) name?
3. Find the symbols for elements with the following atomic numbers.
   - (a) 5
   - (b) 12
   - (c) 18
   - (d) 20
4. In the element argon, $Z = 18$ and $A = 40$. For argon, state the following.
   - (a) The number of neutrons
   - (b) The number of electrons
   - (c) The isotopic symbol for this element
5. How many protons and neutrons are in $^{79}_{35}\text{Br}$?
6. An atomic nucleus consists of one proton and one neutron. What is its symbol?
7. Complete the following table.

Element	Number of protons	Number of electrons	Number of neutrons
$^{12}_{6}\text{C}$			
$^{56}_{26}\text{Fe}$			
$^{40}_{18}\text{Ar}$			
$^{235}_{92}\text{U}$			
$^{238}_{92}\text{U}$			
$^{19}_{9}\text{F}$			

   - (a) Identify any isotopes in the table.
   - (b) Explain the difference between the isotopes.
8. Why do we identify an element by its atomic number rather than its mass number?

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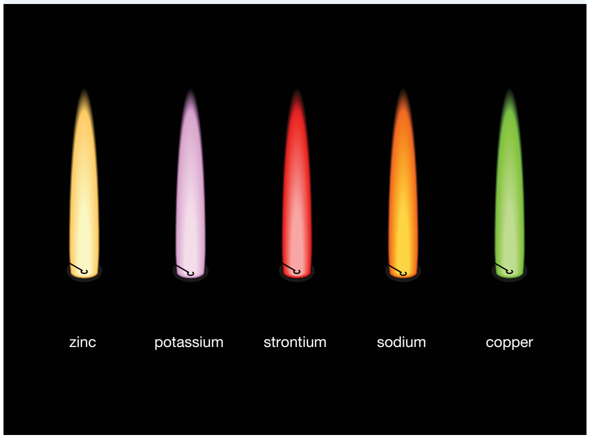
1.4 Electrons

KEY CONCEPT

- Spectral evidence for the Bohr model and for its refinement as the Schrödinger model; electronic configurations of elements 1 to 36 using the Schrödinger model of the atom, including s, p, d and f notations (with copper and chromium exceptions)

1.4.1 Exciting electrons

FIGURE 1.11 When wires with small amounts of different metal salts are placed in a flame, the electrons are excited and emit characteristic coloured light.



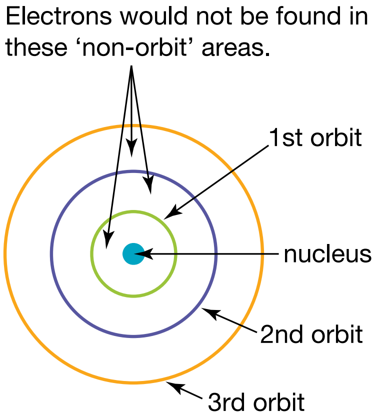
What causes rainbows? Why is it that when you look into a fire you see different coloured flames? The answers lie in the way the electrons are arranged around the nucleus of the atom. This arrangement largely determines the properties and the behaviour of elements and the materials made from them.

When white light is separated by a prism, a *continuous spectrum* of colour is observed. Every element emits light if it is heated by passing an electric discharge through its gas or vapour. This happens because the atoms of the element absorb energy, and then lose it, emitting it as light. Passing the light emitted by an element through a prism produces an **atomic emission spectrum** for that element.

The emission spectra of elements are quite different from the spectrum of white light. White light gives a continuous spectrum, whereas atomic emission spectra consist of separate lines of coloured light. Each line in an emission spectrum corresponds to one particular frequency of light being given off by the atom; therefore, each line corresponds to one exact amount of energy being emitted.

1.4.2 Bohr's energy levels electrons

FIGURE 1.12 The Bohr model of an atom

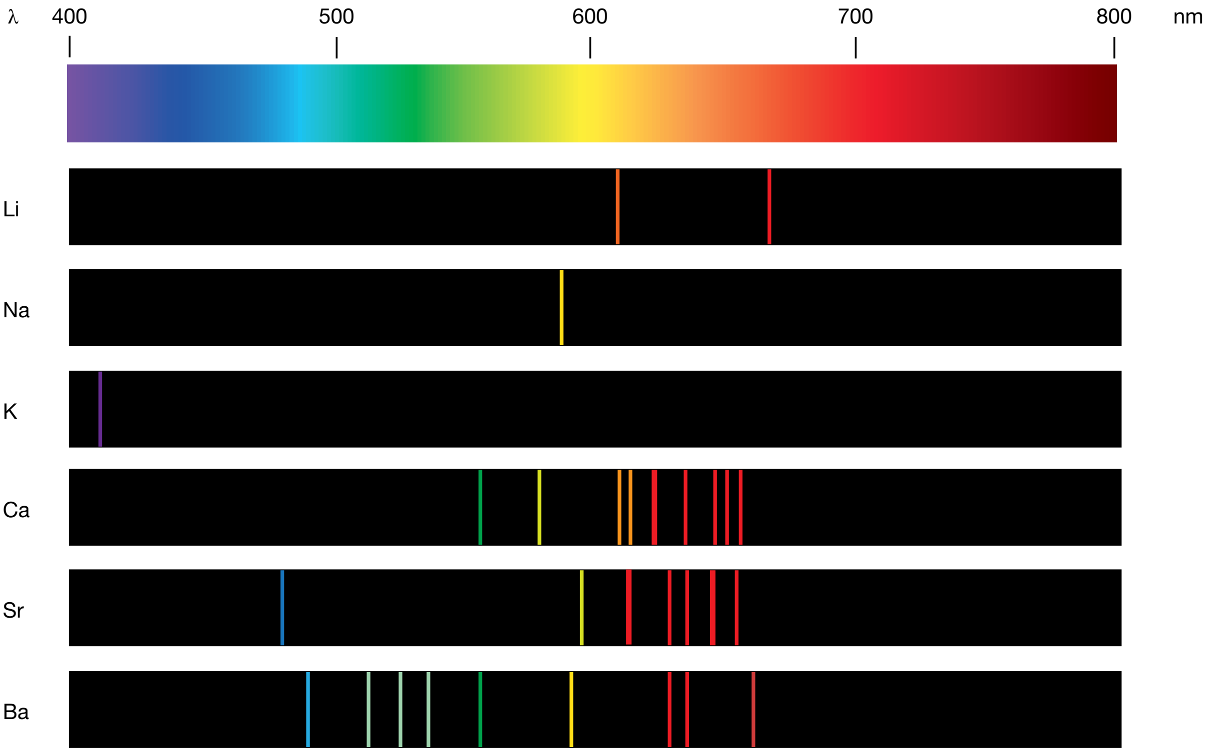


In 1913, Niels Bohr suggested an explanation for the emission spectrum by proposing a model for the hydrogen atom. He proposed that electrons of specific energy move around the central nucleus in circular orbits or energy levels. Electrons cannot exist between these orbits. Although an electron cannot lose energy while orbiting a nucleus, it could be given excess energy (by a flame or electric current) and then move to a higher orbit. If this happens, the electron has moved from the **ground state** (lowest energy level) to an **excited** state. When it drops back down to a lower, more stable orbit, this excess energy is given out as a **photon** or quantum of light. This is seen as a line of a particular colour on the visible spectrum.

The energy given out is the difference in energy between the two energy levels. Since only certain allowed energy levels are possible, the energy released has specific allowed values, each corresponding to a line in the emission spectrum. This spectrum is different for each element, so it is often called the 'fingerprint' by which an element may be identified.

Some metallic elements can be identified simply by their characteristic flame colours when heated. Copper burns with a blue-green flame, and sodium burns with a yellow–orange flame.

FIGURE 1.13 White light is a continuous spectrum (top). The emission spectra of various atomic elements consist of distinct lines that correspond to differences in energy levels.



on Resources

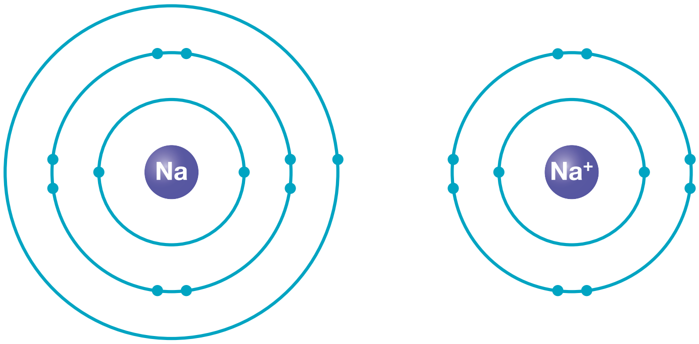
Digital document Experiment 1.1 Flame tests of metal cations (doc-30847)

Teacher lead video Experiment 1.1 Flame tests of metal cations (tlvd-0614)

Electron shells

Electrons may be visualised as moving within a region of space surrounding the nucleus. The regions are called electron shells and are numbered 1, 2, 3 and 4. A definite energy level is associated with each shell; the innermost shell has the lowest energy level. An electron has to gain energy to move further away from the nucleus. If it gains enough energy to completely leave the atom, the particle that is left is no longer neutral and is called a positive ion. Sodium, Na has 11 protons and 11 electrons. If it loses an electron, it becomes the positive ion Na^+ because it now has 11 protons and only 10 electrons.

FIGURE 1.14 Sodium easily loses an outer shell electron to become a sodium ion.



Further studies of line spectra in the 1910s and 1920s led to the prediction that a maximum number of electrons could be present in a given energy level.

1.4.3 Electron configuration

The arrangement of electrons in the shells is called the atom's **electron configuration**. The electron capacity of each shell is limited. The maximum number of electrons that each shell can hold is $2n^2$ where n is the shell number or energy level.

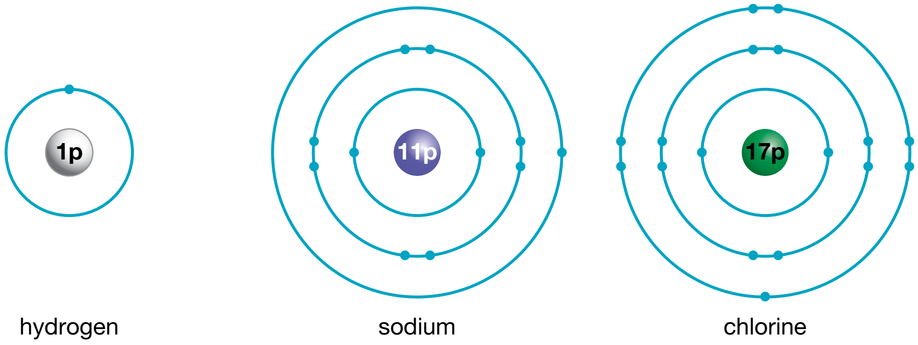
Electron shells are filled in order from the nucleus, starting with the innermost shell, so that the electrons are in their lowest possible energy levels (or ground state). The one electron of a hydrogen atom would be in the first shell, and the electron configuration is written as 1. Sodium has 11 electrons, so two go into the first shell, eight in the second and the last electron in the third shell. The electron configuration of sodium, therefore, is written as 2, 8, 1. Chlorine has 17 electrons and an electron configuration of 2, 8, 7. Note that for the first 20 elements *the third shell never has more than eight electrons*. Potassium, for example, has 19 electrons and an electron configuration of 2, 8, 8, 1 rather than 2, 8, 9. This means that the fourth shell is the outer shell for potassium electrons, rather than the third.

Ions are atoms that have lost or gained one or more electrons. For example, a sodium atom has 11 electrons, so its electron configuration is 2, 8, 1. A sodium ion, Na^+, has lost an electron so its electron configuration is 2, 8. Chemists are particularly interested in the electrons in the highest energy level of an atom since it is these outer-shell electrons that mainly determine the chemical properties of elements. These electrons are called **valence electrons**.

Shell model diagrams

The electron configuration of an atom can be represented using shell model diagrams, like those in figure 1.15. These show the electron shells as concentric rings around the nucleus, with the electrons marked on each ring, and help us to visualise the structure and behaviour of atoms.

FIGURE 1.15 Shell model diagrams of hydrogen, sodium and chlorine. Dots represent the electrons in the shells.



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Interactivity Shell-shocked (int-0676)

Limitations of the shell model

The shell model represents only part of the story of the atom. More discoveries are always being made that cause scientists to reconsider their models and their understanding of the atom. The shell model doesn't really explain the various differences in energies between the electron shells. It seems to imply that all the electrons orbit the nucleus in exactly circular paths, like planets around a sun. We know from looking at molecules with electron tunnelling microscopes that they come in many different shapes and sizes, so this model does not fully explain every aspect of every atom.

The order of the electrons filling the electron shells is not really explained by this model either. For example, compare calcium (2, 8, 8, 2) with scandium (2, 8, 9, 2) — why isn't the electron configuration of scandium 2, 8, 8, 3? Other models have been developed that are more complex and explain more of the data scientists have gathered.

1.4.4 Erwin Schrödinger: the quantum-mechanical model of an atom

FIGURE 1.16 Schrödinger developed the quantum-mechanical model of the atom.

Although Bohr's model of an atom could account for the lines seen in the emission spectra of hydrogen, it did not explain and could not mathematically predict the lines seen in the emission spectra of the more complex atoms. Furthermore, it did not explain why electrons moving around the nucleus and emitting electromagnetic radiation did not fall into the nucleus of the atom, causing it to collapse.

In 1923 the French scientist Louis de Broglie (1892–1987) proposed, using Albert Einstein's and Max Planck's quantum theory, that electrons show both particle and wave behaviour. In 1926, the Austrian physicist Erwin Schrödinger (1887–1961) made use of the new quantum theory to refine Bohr's model of the atom. He wrote and developed a wave equation describing the location and energy of an electron in a hydrogen atom. From this, he developed the currently accepted quantum-mechanical model of the atom — a complex mathematical model based on particles such as electrons showing wave-like behaviour. This model features electrons arranged in shells, subshells and orbitals within an atom.

According to quantum mechanics, the electron is not considered as moving along a definite path. Instead, the electron is found in a region of space around a nucleus called an **orbital**. An orbital may be visualised as a blurry cloud of negative charge; the cloud is most dense where the probability of finding the electron is large and less dense where the probability of finding the electron is small. As in the Bohr model, the electron is attracted to the nucleus by electrostatic forces and moves in such a way that its total energy has a specific value.

FIGURE 1.17 The quantum-mechanical model showing (a) the charge cloud for the 1*s* electron in hydrogen and (b) charge clouds for the electrons in the 1*s* and 2*s* subshells.

In the quantum-mechanical model of the atom:

1. The energy levels of electrons are designated by **principal quantum numbers, *n***, and are assigned specific values: $n = 1, 2, 3, 4, 5$ and so forth. These principal quantum numbers may be referred to as shells.
2. Within each shell, several different energy levels called subshells may be found. The number of subshells equals the shell number; for example, if the shell number is 2, two subshells are present at that energy level. Each subshell corresponds to a different electron cloud shape. Subshells are represented by the letters *s, p, d, f* and so on.

TABLE 1.2 Energy levels within shells of an atom

Shell number (n)	Number of subshells	Subshell symbol	Maximum number of electrons in subshell
1	1	*s*	2
2	2	*s*	2
		p	6
3	3	*s*	2
		p	6
		d	10
4	4	*s*	2
		p	6
		d	10
		f	14

1.4.5 Electron configuration

The way in which electrons are arranged around the nucleus of an atom is called the electron configuration of the atom.

Generally, the order of subshell filling is from lowest energy first:

$$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p...$$

Notice that the 4*s* subshell is filled before the 3*d* subshell, which is of a higher energy than the 4*s* subshell. Likewise, the 4*d* subshell is higher in energy than the 5*s* subshell, and so on. This is demonstrated in figure 1.19, which shows the order of filling of subshells may be found by following the arrows from tail to head, starting with the top arrow.

FIGURE 1.18 The energy levels of atomic orbitals

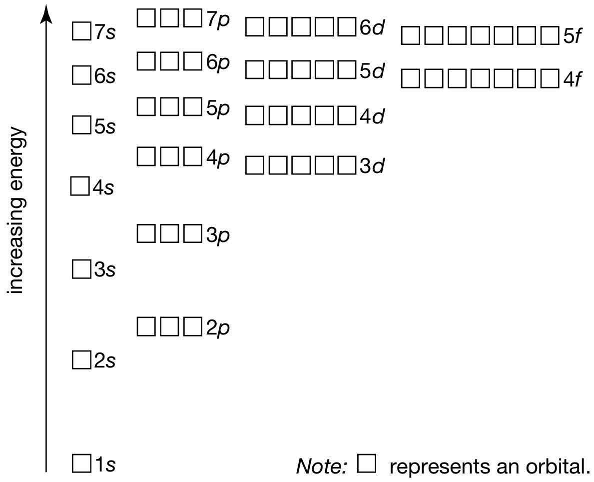
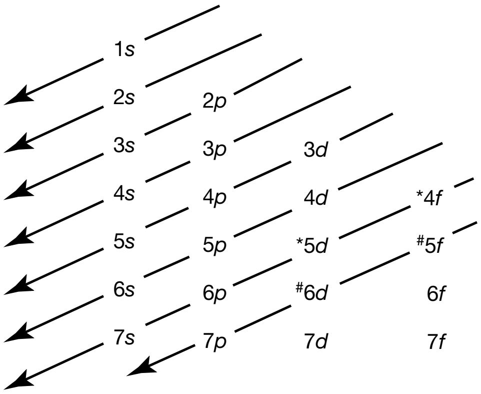


FIGURE 1.19 The order of filling of subshells



* and #: In both these cases it is found that the two subshells fill at the same time.

Excited states

When an atom moves to a higher energy level than the ground state by absorbing energy, its electron configuration changes. The outermost electron moves to a higher energy level subshell.

For example, neon has two electrons in the first shell and eight in the second shell, $1s^2 2s^2 2p^6$; when the outermost electron gains energy, this becomes $1s^2 2s^2 2p^5 3s^1$.

Note that, once the order of filling subshells has been determined, the subshells are written in increasing numerical order, rather than the order of increasing energy.

So $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$ becomes $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 4s^2$.

The electron configuration for the elements hydrogen, sodium and scandium, for example, can be written as:

H (Z = 1)	$1s^1$
Na (Z = 11)	$1s^1 2s^2 2p^6 3s^1$
Sc (Z = 21)	$1s^1 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$

The electron configuration for scandium can also be written as $1s^2 2s^2 2p^6 3s^2 3p^6 3d^1 4s^2$ to show the last occupied shell is the fourth one; nevertheless, the fourth shell starts to fill before the third shell is complete.

The position of an element on the periodic table can also be easily ready from the electron configuration. The periodic table is divided into groups (the horizontal rows of the table) and periods (the vertical columns of the table). The significance of the groups and periods will be discussed in topic 2.

on Resources

Weblink Interactive periodic table

SAMPLE PROBLEM 3

Find the ground state electron configuration of a fluorine atom.

Teacher-led video: SP3 (tlvd-0510)

THINK	WRITE
1. The atomic number of fluorine is 9. Fluorine has nine electrons. According to the order of subshell filling, the 1*s* subshell in a fluorine atom fills first, and contains two of fluorine's nine electrons.	$1s^2$
2. The next energy level is the 2*s* subshell — this holds another two electrons.	$1s^2 2s^2$
3. The 2*p* subshell can hold six electrons; however, since only five electrons remain to be placed ($9 - 2 - 2 = 5$), one of the 2*p* subshell orbitals is incomplete.	$1s^2 2s^2 2p^5$

PRACTICE PROBLEM 3

Find the ground state electron configuration of a potassium atom.

SAMPLE PROBLEM 4

Find the electron configuration of an aluminium ion, Al^{3+}.

Teacher-led video: SP4 (tlvd-0511)

THINK	**WRITE**
1. An aluminium ion, Al^{3+}, is an aluminium atom ($Z = 13$) that has lost three electrons. Hence $13 - 3 = 10$ electrons need to be placed in shells. The 1*s* shell contains two electrons.	$1s^2$
2. The next energy level is the 2*s* subshell — this holds another two electrons.	$1s^22s^2$
3. The 2*p* subshell can hold six electrons and because six electrons remain to be placed ($10 - 2 - 2 = 6$), the 2*p* subshell orbitals is complete.	$1s^22s^22p^6$

PRACTICE PROBLEM 4

The following electron configuration represents an atom or ion.

$$\mathbf{1s^22s^22p^63p^1}$$

Is it in an excited state?

1.4.6 Chromium and copper: atypical electron configurations

A few elements have electron configurations that do not follow the usual pattern. Chromium (atomic number 23) and copper (atomic number 29), for example, may be expected to be written as follows:

Chromium $1s^22s^22p^63s^23p^63d^44s^2$

Copper $1s^22s^22p^63s^23p^63d^94s^2$

The correct electron configurations are:

Chromium $1s^22s^22p^63s^23p^63d^54s^1$

Copper $1s^22s^22p^63s^23p^63d^{10}4s^1$

These arrangements give chromium a half-filled *d* subshell and copper a filled *d* subshell. Filled subshells are more stable than half-filled subshells. However, half-filled subshells are more stable than other partly filled subshells.

FIGURE 1.20 Chromite is an ore of chromium.

FIGURE 1.21 Chalcopyrite is an ore of copper.

1.4 EXERCISE

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1. Complete the table. The first row has been done for you.

Energy level (n)	**Maximum number of electrons ($2n^2$)**
1	$2 \times 1^2 = 2$
2	
3	
4	

2. (a) What was the importance of Bohr's contribution to our understanding of the atom?
   (b) What are the limitations of Bohr's model?
3. Complete the following table to show the electron configuration of the first twenty elements. Some of the configurations have been provided for you.

Element	**Symbol**	**Atomic number**	**Electron configuration** ***1, 2, 3, 4***
Hydrogen	H	1	1
Helium	He	2	
Lithium	Li	3	
Beryllium	Be	4	2, 2
Boron	B	5	
Carbon	C	6	
Nitrogen	N	7	
Oxygen	O	8	
Fluorine	F	9	
Neon	Ne	10	
Sodium	Na	11	2, 8, 1
Magnesium	Mg	12	
Aluminium	Al	13	
Silicon	Si	14	
Phosphorus	P	15	
Sulfur	S	16	
Chlorine	Cl	17	2, 8, 7
Argon	Ar	18	
Potassium	K	19	2, 8, 8, 1
Calcium	Ca	20	

4. Using the periodic table, produce shell model diagrams for the first 10 elements.

5. (a) What is the electron configuration of nitrogen?
   (b) Which element has atoms with the electron configuration of 2, 8, 8, 2?
6. Write the full electron configuration of the following elements: lithium, potassium, carbon, chlorine, argon and nitrogen.
7. The isotope $^{32}_{15}\text{P}$ is used in the treatment of leukaemia.
   (a) Write the full electron configuration of this isotope.
   (b) How does it differ from $^{31}_{15}\text{P}$?
8. An element *X* has configuration $1s^22s^22p^63s^23p^4$.
   (a) What group is it in?
   (b) What period is it in?
   (c) Give its name and symbol.
9. Name the elements with the following configurations.
   (a) $1s^22s^1$
   (b) $1s^22s^22p^6$
   (c) $1s^22s^22p^63s^1$
   (d) $1s^22s^22p^63s^23p^1$
   (e) $1s^22s^22p^63s^23p^6$
   (f) $1s^22s^22p^63s^23p^64s^2$
10. For each of the elements in the following two sets of atomic numbers, write the electron configuration.
   (a) $Z = 20, 12, 4, 9$
   (b) $Z = 5, 6, 8, 16$
11. Distinguish between the terms ‘shell’ and ‘subshell’.
12. Write the ground state electron configurations of the first 20 elements of the periodic table.
13. Which of the following electron configurations are ground state configurations and which are excited state configurations?
   (a) $1s^22s^22p^63s^1$
   (b) $1s^22s^22p^63s^23p^43d^1$
   (c) $1s^22s^22p^63s^23p^63d^54s^2$
14. A neutral magnesium atom has an electron configuration of $1s^22s^22p^63s^16s^1$.
   (a) How can you tell that the atom is in an excited state?
   (b) Describe what would happen if the atom changed its electron configuration to the ground state.
   (c) Write the ground state electron configuration of the magnesium ion Mg^{2+}.
15. Identify the following elements with ground state electron configurations of
   (a) $1s^22s^22p^63s^23p^63d^54s^2$.
   (b) $1s^22s^22p^63s^23p^64s^2$.

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1.5 Review

1.5.1 Summary

Atomic structure

- The nuclear model of the atom proposes that atoms are composed of the following subatomic particles:
  - protons, which are positively charged particles found in the nucleus
  - neutrons, which are neutrally charged particles found in the nucleus
  - electrons, which are negatively charged particles found in energy levels around the nucleus.
- The atom is mainly empty space.

- Each element is represented by a specific chemical symbol and is composed of only one kind of atom.
- The atoms of an element:
  - all contain the same number of protons
  - are neutral, because the number of electrons is equal to the number of protons.
- A molecule is a group of atoms bonded together.
- Particles at the atomic level are measured using the nanometre (nm); 1 nm = 10^{-9} m.
- Nanotechnology works at the scale of atoms and molecules.
- Nanotechnology is likely to benefit many areas of our lives in the future, particularly in the area of improving and extending life and ensuring a cleaner and safer environment to live in. It has many potential benefits in biotechnology and drug delivery.

Representing atoms

- An isotope is an atom of an element with a different number of neutrons and hence a different mass number.
- The isotopic symbol of an element E may be written as ${}^{A}_{Z}E$ where:
  - atomic number (Z) = number of protons
  - mass number (A) = number of protons + number of neutrons.

Electrons

- Niels Bohr's model suggested that electrons orbited around the nucleus in definite paths of energy called quantum levels or shells. A quantum of energy or photon is emitted when an electron that has gained enough energy to move to a higher energy level then falls back to its ground state position. Bohr determined the different energy levels by using mathematical formulas that measured the wavelengths of the different energy levels.
- Erwin Schrödinger's most important contribution to modern atomic theory was his development of the mathematical description of the paths that electrons would most likely follow in their orbits around the nucleus. The formulas he developed formed the basis of the quantum-mechanical model of the atom. In it he proposed that, instead of Bohr's idea of electrons following predetermined paths, they move around in regions of space called orbitals.
- The quantum-mechanical model has the following features:
  - Orbitals are of various shapes and are found within subshells, which in turn are found in shells.
  - Each shell has a different energy level, with the shell furthest away from the nucleus having the highest energy level and the one closest to the nucleus having the lowest energy level.
  - Each energy shell has a principal quantum number, n. The first four shells have n = 1, 2, 3 and 4.
  - Subshells are energy levels found within shells. There are four different types of subshells — the *s, p, d* and *f* subshells.
  - The maximum number of electrons the different subshells can hold are:
    s subshell: 2 electrons
    p subshell: 6 electrons
    d subshell: 10 electrons
    f subshell: 14 electrons.
  - Electrons fill shells and subshells of lowest energy first. The order of filling is:
    $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d$.
- The ground state electron configuration of an atom refers to electrons in their lowest energy level. Any other configurations represent the atom in an excited state and in a higher energy level.

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1.5.2 Key terms

atomic emission spectrum a spectrum emitted as distinct bands of light of diagnostic frequencies by elements or compounds
atom a neutral particle with a nucleus; the smallest sample of an element
atomic number the number of protons in the nucleus of an atom of a particular element
electron configuration the arrangement of electrons in the shells of an atom
element chemical species consisting of atoms of a single type
excited state raised to a higher than ground-state energy level
ground state the least excited energy level of an atom
ion a charged atom
isotopes forms of an element with the same number of protons but different numbers of neutrons in the nucleus
isotopic symbol representation of an element as $^{A}_{Z}E$, where E is the symbol for the element, A is the mass number and Z is the atomic number
mass number the total number of protons and neutrons in the nucleus of a particular isotope of an element
molecule group of atoms bonded together covalently
nanometre one billionth of a metre, 0.000000001 m (10^{-9} m)
orbital three-dimensional wave describing a bound electron
photon particle of light
principal quantum number, *n* quantum number that indexes the energy and size of an atomic orbital
subatomic particles electrons, protons and neutrons
valence electrons electrons in the outermost shell (highest energy level) of an atom; largely determine chemical properties of an element and contribute to chemical bond formation

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Digital document Key terms glossary — Topic 1 (doc-30918)

1.5.3 Practical work and investigations

Experiment 1.1

Flame tests of metal cations

Aim: To observe the characteristic flame colours of the metal ions K^+, Na^+, Li^+, Sr^{2+}, Cu^{2+}, Ca^{2+}, Ba^{2+}, and to identify an unknown metal ion

Digital document: doc-30847
Teacher-led video: tlvd-0614

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Digital document Practical investigation logbook (doc-30919)

1.5 Exercises

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1.5 Exercise 1: Multiple choice questions

1. An atomic particle has a net charge of zero and is found in the nucleus. Which of the following particles is it?
   - A. Proton
   - B. Electron
   - C. Neutron
   - D. Positron
2. Atoms
   - A. have a nucleus that occupies most of the atom's volume.
   - B. are all the same size.
   - C. contain only protons in their nucleus.
   - D. have most of their mass in the nucleus.
3. Which of the following particles have approximately the same mass?
   - A. A proton and an electron
   - B. A proton and a neutron
   - C. A neutron and an electron
   - D. An electron and a hydrogen atom
4. A nanometre is
   - A. one billionth of a metre.
   - B. one hundredth of a metre.
   - C. one thousandth of a metre.
   - D. one hundred thousandth of a metre.
5. What is the nanometre scale in centimetres?
   - A. 10^{-5}
   - B. 10^{-9}
   - C. 10^{-7}
   - D. 10^{-12}
6. Which of the following is the correct ranking of prefixes from largest to smallest when they refer to particle size?
   - A. Nano, micro, kilo
   - B. Nano, kilo, micro
   - C. Micro, kilo, nano
   - D. Kilo, micro, nano
7. In the nucleus of an atom are 11 protons and 12 neutrons. What is its mass number?
   - A. 11
   - B. 12
   - C. 1
   - D. 23
8. What atomic number does the element ${}^{16}_{8}O$ have?
   - A. 24
   - B. 16
   - C. 8
   - D. 2

9. What is the number of valence electrons in the element calcium?
   - A. 1
   - B. 2
   - C. 20
   - D. 40

10. Which of the following statements regarding subatomic particles is correct?
    - A. Protons are positively charged particles and neutrons are negatively charged.
    - B. The relative masses of an electron, a proton and a neutron are all about 1 unit.
    - C. In a neutral atom, the number of neutrons is equal to the number of protons.
    - D. Isotopes of an element have the same number of protons but different numbers of neutrons.

11. What did Bohr's theory of the atom propose?
    - A. Electrons orbit the nucleus like planets move around the sun.
    - B. No more than two electrons are allowed in any energy level.
    - C. Electrons move around the nucleus in fixed orbits, each of which has a different energy level.
    - D. Energy shells have subshells that contain regions of space called orbitals.

12. What is the maximum number of electrons that can be placed in the shell $n = 3$?
    - A. 18
    - B. 8
    - C. 2
    - D. 32

13. Which of these electron configurations represents an atom in an excited energy state?
    - A. $1s^2 2s^2 2p^6 3s^2$
    - B. $1s^2 2s^2 2p^6 3s^2 3p^4 3d^1$
    - C. $1s^2 2s^2 2p^6$
    - D. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$

14. What is the electron configuration of ${}^{35}_{17}\mathrm{Cl}^-$?
    - A. $1s^2 2s^2 2p^6 3s^2 3p^5 3d^1$
    - B. $1s^2 2s^2 2p^6 3s^2 3p^4 4s^2$
    - C. $1s^2 2s^2 2p^6 3s^2 3p^6$
    - D. $1s^2 2s^2 2p^6 3s^2 3p^4 4s^1 3d^1$

15. The electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6$ represents which of the following ions?
    - A. O^{2-}
    - B. S^{2-}
    - C. Al^{3+}
    - D. Na^+

16. The electron configuration of an atom X is $1s^2 2s^2 2p^6 3s^2 3p^1$. Which of the following formulas is most likely to be a compound formed with X?
    - A. XF_2
    - B. CaX
    - C. XCl_3
    - D. MgX_2

1.5 Exercise 2: Short answer questions

1. Match each object with its correct size.

Object	Size
Red blood cell	0.15 nm
Virus	7000 nm
Hydrogen molecule	30 nm

2. For each of the atoms below state the:

- **i.** atomic number
- **ii.** mass number
- **iii.** number of protons
- **iv.** number of neutrons
- **v.** number of electrons
- **vi.** name of element.

a. $^{23}_{11}\text{Na}$ **b.** $^{19}_{9}\text{F}$ **c.** $^{28}_{14}\text{Si}$ **d.** $^{56}_{26}\text{Fe}$ **e.** $^{197}_{79}\text{Au}$ **f.** $^{235}_{92}\text{U}$

3. Complete the following table.

Name of atom	Atomic number	Mass number	Protons	Neutrons	Electrons
	18	36			
		34		18	
	18			20	
		31			15
Lead				126	82
			19	20	
				20	16

4. What is the general name for the group of atoms that includes carbon-12, carbon-13 and carbon-14?

5. From the information in the table, complete the following.

- **a.** Identify those species that are isotopes of the same element.
- **b.** Write the isotopic symbol for each atom.

Atom	Neutrons	Protons	Electrons
A	20	20	20
B	20	17	17
C	22	20	20
D	26	20	20

6. The following isotopes belong to three elements. Identify the elements and list the isotopes next to the name of each element.

$$^{37}_{17}A \quad ^{26}_{12}B \quad ^{59}_{27}C \quad ^{35}_{17}D \quad ^{25}_{12}E \quad ^{60}_{27}F \quad ^{24}_{12}G$$

7. **a.** What is required for an electron in an atom to move from a lower energy level to a higher energy level?

- **b.** What happens when an electron returns to a lower level?

8. If the flame spectra of calcium chloride, calcium carbonate and calcium nitrate were observed, would you observe one main colour or three?

9. An atom has several excited states but only one ground state. What does this mean?

10. Explain the difference between the following.

- **a.** A ‘shell’ and a ‘subshell’
- **b.** An atomic ‘orbit’ and an ‘orbital’

11. Chromium (atomic number 23) and copper (atomic number 29) do not follow the usual pattern of subshell electron configuration.

- **a.** Why does this occur?
- **b.** Write the subshell electron configuration for chromium.
- **c.** Write the subshell electron configuration for copper.

1.5 Exercise 3: Exam practice questions

Question 1 (1 mark)

Explain why a neutral atom cannot have the atomic configuration of 6 protons, 13 neutrons and 13 electrons.

Question 2 (9 marks)

Nanoparticles have very large surface area to volume ratios compared to the same material in bulk, as lumps, powders or as sheets. This often effects their properties.

a. A titanium oxide, TiO_2, nanoparticle has a diameter of 25 nm. It is often added to paper to enhance its brightness. If a piece of paper is typically 1 mm thick, how many times smaller is a single nanoparticle of TiO_2 compared to the thickness of one piece of paper? **1 mark**

b. If a cube-shaped nanomaterial has a side length of 100 nm, what would be its

  i. surface area?

  ii. volume?

  iii. surface area to volume ratio? **3 marks**

c. If the 100 nm side length cube was reduced to 10 nm, how would this effect the surface area to volume ratio? **5 marks**

Question 3 (6 marks)

a. Write the electron configuration of each of the following in their ground states.

  i. Argon atom, $Z = 18$

  ii. Calcium ion, Ca^{2+}, $Z = 20$ **2 marks**

b. By referring to the electron configuration of the argon atom, explain how its group and period on the periodic table can be determined. **4 marks**

1.5 Exercise 4: studyON Topic Test online only

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2 The periodic table

2.1 Overview

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2.1.1 Introduction

The periodic table of chemical elements is one of the most significant achievements in science, capturing the essence not only of chemistry, but also of physics and biology. It is a unique tool, enabling scientists to predict the appearance and properties of matter on the Earth and in the rest of the Universe. The year 1869 is considered the time Dmitri Mendeleev synthesised centuries of previous experiments and constructed the modern Periodic System. In 2019 we saw the 150th anniversary of the Periodic Table of Chemical Elements and so this year was proclaimed the International Year of the Periodic Table of Chemical Elements by the United Nations General Assembly and UNESCO.

Dmitri Mendeleev (1834–1907), a Russian chemist, spent years collecting and sorting information about each of the 63 elements known at the time and constructed a set of data cards (one data card for each element). On each card he noted the atomic mass and other properties of the element and its compounds. He noticed the existence of 'groups' of different elements with similar chemical properties. He then produced a periodic table according to increasing order of relative atomic mass and the periodicity of their properties; that is, similar physical and chemical properties occurring at regular intervals. He even left gaps for elements that, he correctly predicted, had not yet been discovered.

FIGURE 2.1 Dmitri Mendeleev evolved earlier versions of the periodic table by organising groups of different elements with similar chemical properties into a chart, forming the basis for today's periodic table.

2.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- the periodic table as an organisational tool to identify patterns and trends in, and relationships between, the structures (including electronic configurations and atomic radii) and properties (including electronegativity, first ionisation energy, metallic/non-metallic character and reactivity) of elements.

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2.2 Organisation of the periodic table

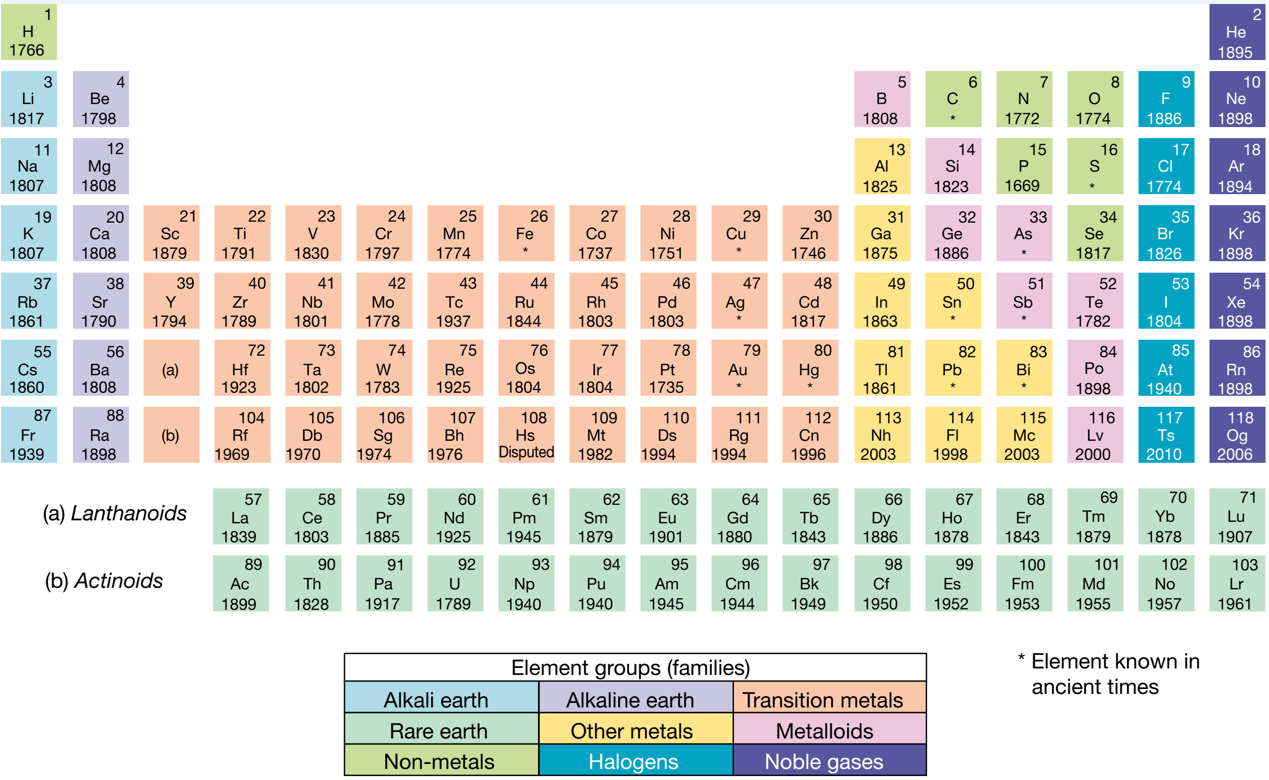
KEY CONCEPT

- The periodic table as an organisational tool

2.2.1 The modern periodic table

The periodic table of the elements has become one of the most important icons in science today. A chart of this table hangs on the wall of almost every classroom and chemical laboratory in schools, universities and research institutions around the world. It is a single document that consolidates much of our knowledge of chemistry and is a vital tool for modern chemists. The modern periodic table now consists of 118 known elements. Elements up to and including uranium (atomic number 92) are naturally occurring. All the elements beyond uranium have been synthesised by chemists since 1940. They are all radioactive elements and are called the transuranium elements. These elements do not occur in nature but have been synthesised in nuclear reactors, which accelerate bombarding particles to very high speeds and shoot them at very heavy target nuclei.

FIGURE 2.2 Periodic table showing discovery date of elements



1 H 1766																	2 He 1895
3 Li 1817	4 Be 1798											5 B 1808	6 C *	7 N 1772	8 O 1774	9 F 1886	10 Ne 1898
11 Na 1807	12 Mg 1808											13 Al 1825	14 Si 1823	15 P 1669	16 S *	17 Cl 1774	18 Ar 1894
19 K 1807	20 Ca 1808	21 Sc 1879	22 Ti 1791	23 V 1830	24 Cr 1797	25 Mn 1774	26 Fe *	27 Co 1737	28 Ni 1751	29 Cu *	30 Zn 1746	31 Ga 1875	32 Ge 1886	33 As *	34 Se 1817	35 Br 1826	36 Kr 1898
37 Rb 1861	38 Sr 1790	39 Y 1794	40 Zr 1789	41 Nb 1801	42 Mo 1778	43 Tc 1937	44 Ru 1844	45 Rh 1803	46 Pd 1803	47 Ag *	48 Cd 1817	49 In 1863	50 Sn *	51 Sb *	52 Te 1782	53 I 1804	54 Xe 1898
55 Cs 1860	56 Ba 1808	(a)	72 Hf 1923	73 Ta 1802	74 W 1783	75 Re 1925	76 Os 1804	77 Ir 1804	78 Pt 1735	79 Au *	80 Hg *	81 Tl 1861	82 Pb *	83 Bi *	84 Po 1898	85 At 1940	86 Rn 1898
87 Fr 1939	88 Ra 1898	(b)	104 Rf 1969	105 Db 1970	106 Sg 1974	107 Bh 1976	108 Hs Disputed	109 Mt 1982	110 Ds 1994	111 Rg 1994	112 Cn 1996	113 Nh 2003	114 Fl 1998	115 Mc 2003	116 Lv 2000	117 Ts 2010	118 Og 2006
(a) *Lanthanoids*			57 La 1839	58 Ce 1803	59 Pr 1885	60 Nd 1925	61 Pm 1945	62 Sm 1879	63 Eu 1901	64 Gd 1880	65 Tb 1843	66 Dy 1886	67 Ho 1878	68 Er 1843	69 Tm 1879	70 Yb 1878	71 Lu 1907
(b) *Actinoids*			89 Ac 1899	90 Th 1828	91 Pa 1917	92 U 1789	93 Np 1940	94 Pu 1940	95 Am 1945	96 Cm 1944	97 Bk 1949	98 Cf 1950	99 Es 1952	100 Fm 1953	101 Md 1955	102 No 1957	103 Lr 1961

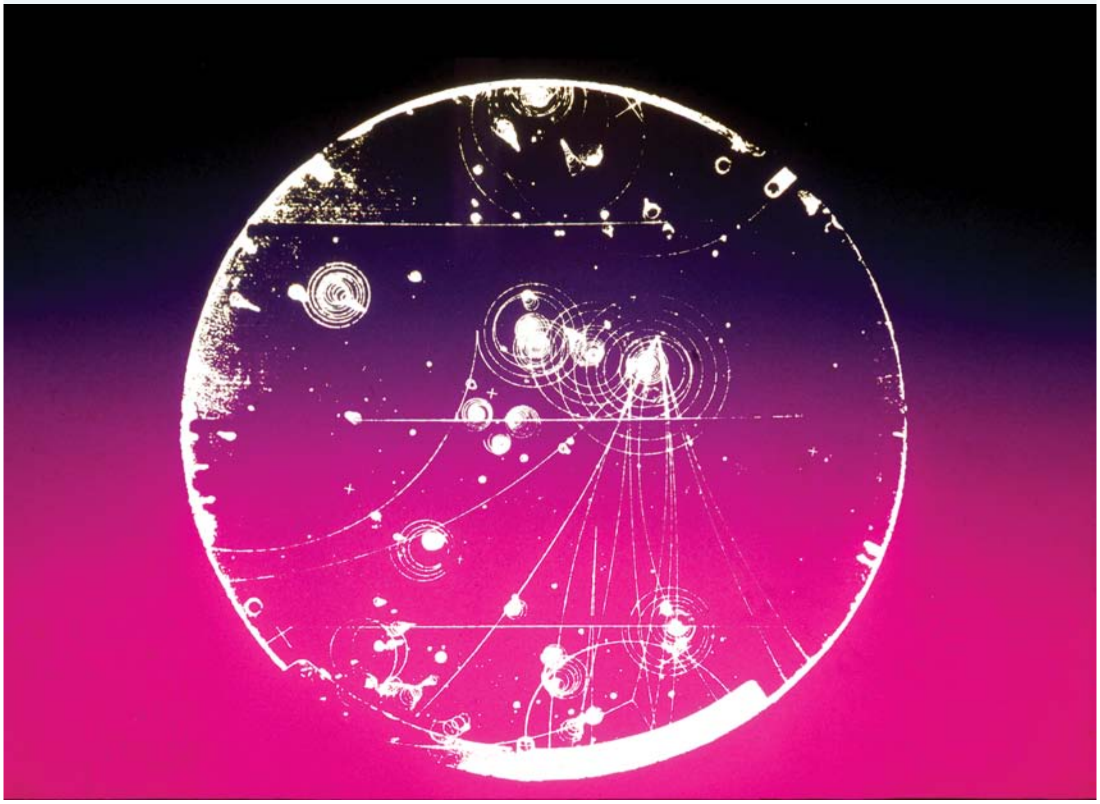
Element groups (families)		
Alkali earth	Alkaline earth	Transition metals
Rare earth	Other metals	Metalloids
Non-metals	Halogens	Noble gases

* Element known in ancient times

In the modern periodic table, all the chemical elements are arranged in order of increasing atomic number (the number of protons in a nucleus of an atom of that element). The elements are arranged in rows and columns in relation to their electronic structures and also their chemical properties. Modern understanding of the periodic table arose from the recognition of four principles:

1. Atomic number, rather than atomic mass, was the basic property that determined the order of the elements in the periodic table.
2. Repeating patterns of electron configuration were observed when the electrons around the nucleus of an atom were arranged in order of increasing energy level.
3. The arrangement of the outer-shell electrons was most important in determining the chemical properties of an element.
4. The periodic recurrence of similar properties was seen to result from the periodic change in the electronic structure.

FIGURE 2.3 Particle tracks like this one are part of the evidence to show that a new particle has been produced. Darmstadtium, Ds, element 110, and roentgenium, Rg, element 111, were first discovered in 1994 using a heavy ion accelerator from the fusion of lead and other elements.



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Digital document Investigation: History of the periodic table (doc-30923)

2.2.2 Periods and groups in the periodic table

The seven horizontal rows in the periodic table are called **periods**. Each period corresponds to the filling of a shell. The location of an element in a period tells you the number of shells each atom of that element has. Elements in the third period, for example, have three shells.

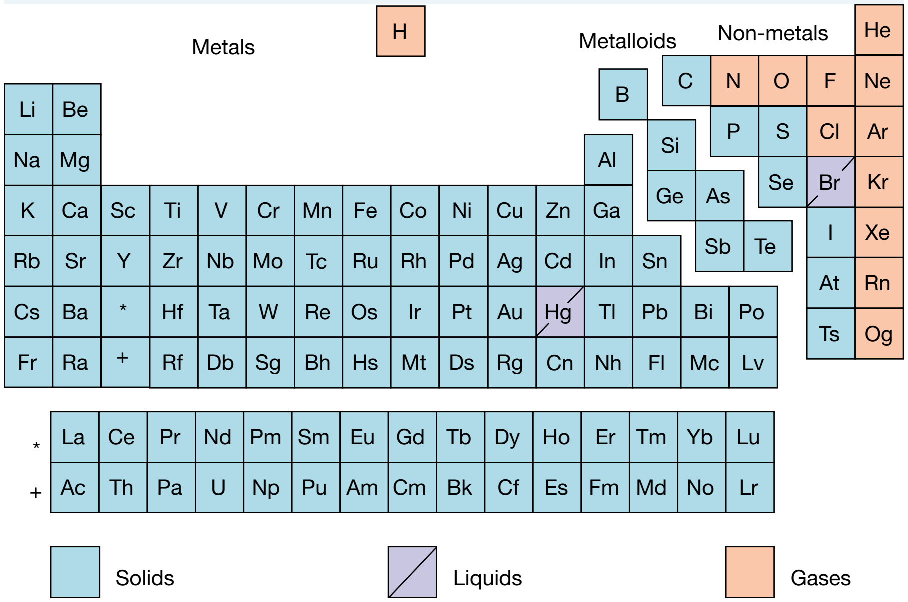
Vertical columns of elements are called **groups**. For example, all atoms of group 2 elements have two electrons in their outer shell. Groups 1, 2 and 13 to 18 are also referred to as main-group elements. These include not only the most abundant elements on Earth but also the most abundant elements in the universe.

A periodic table labelled with periods and groups can be found in the appendix.

2.2.3 Metals and non-metals in the periodic table

Elements may be classified as **metals** or **non-metals**. In the periodic table, the metals are found towards the left side and the non-metals are found towards the right side, as shown in figure 2.4. Most elements are metals, although some elements show both metallic and non-metallic characteristics. These elements are known as *metalloids*, and are also shown in figure 2.4.

FIGURE 2.4 Metals, non-metals and metalloids in the periodic table



The elements in groups 3 to 12 include some familiar metals, such as iron, which is commonly used in construction, copper, which is particularly useful for electrical wiring, and gold, which is used in jewellery. These groups are known as the transition elements. Transition elements contain atoms with filling *d* subshells. These elements are metals and have the properties listed in table 2.1 but are less reactive than *s*-block metals, and most form coloured compounds. In addition, most can form ions with different charges; for example, copper can form Cu^+ ions and Cu^{2+} ions.

TABLE 2.1 General properties of metals and non-metals

Metals	Non-metals
High melting and boiling points	Low melting and boiling points
Good conductors of heat and electricity	Poor conductors of heat and electricity
Opaque	Transparent in a thin sheet
Shiny appearance	Dull colour
Ductile and malleable	Brittle when solid
Strong	Weak
Form positive ions (cations)	Form negative ions (anions)

2.2 EXERCISE

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1. Find out why element 117 tennessine was previously called ununseptium and how it was synthesised.
2. Explain why the classification of elements into a periodic table may be useful to chemists.

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2.3 Patterns and trends in the periodic table

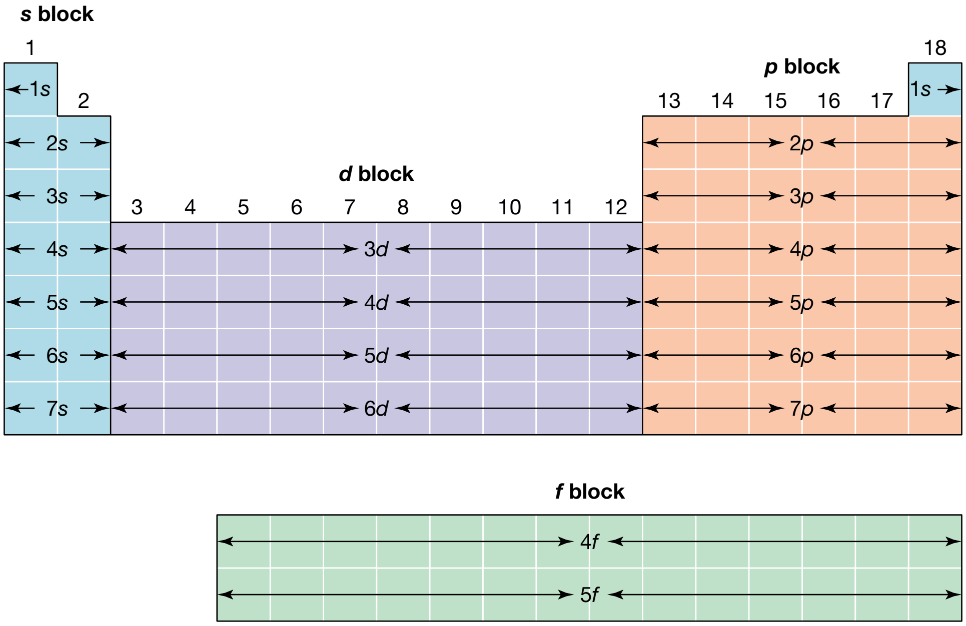
KEY CONCEPT

- The periodic table as an organisational tool to identify patterns and trends in, and relationships between, the structures (including electronic configurations and atomic radii) and properties (including electronegativity, first ionisation energy, metallic/non-metallic character and reactivity) of elements

2.3.1 Periodic patterns and trends

Elements in the periodic table can be divided into four main blocks according to their electron configurations.

FIGURE 2.5 Each block of the periodic table corresponds to a particular subshell.



s-block elements

The elements in group 1 and group 2 form a block of reactive metals and are known as the *s*-block elements. These elements have their outermost electrons in the *s* subshell. Group 1 elements have outer shells of s^1 and group 2 elements have outer shells of s^2. Helium is a group 2 element with a filled *s* subshell of the innermost shell of the atom, rendering it unreactive. It is often grouped with the group 18 noble elements with similar properties.

p-block elements

The elements in groups 13 to 18 form the *p* block, in which elements have their outermost electrons in the *p* subshells. These elements have outermost electron configurations of s^2p^1 to s^2p^6.

d-block elements

The *d*-block elements, from group 3 to group 12, are the transition metals or transition elements. All elements in this block fill their electrons subshells in the order $1s < 2s < 2p < 3s < 3p < 4s < 3d$. For example, period 4 transition metal elements fill the 4*s* subshell before the 3*d* subshell. Similarly, electrons are lost from the 4*s* subshell before the 3*d* subshell. Their outermost shell electron configurations are d^1s^2 to $d^{10}s^2$.

f-block elements

The lanthanoids and actinoids form a block of elements within the transition metals and are sometimes known as the inner transition elements. These elements form the *f* block of the periodic table and have their *f* subshells progressively filled.

The periodic table in figure 2.6 is based on the IUPAC periodic table and shows relative atomic masses to two decimal places. It is common to use only one decimal place in calculations. Relative atomic mass (A_r) is based on the carbon-12 atom, the most common isotope of carbon. This isotope is assigned a mass of exactly 12. On this scale, 1 is therefore equal to $\frac{1}{12}$ of the mass of a carbon-12 atom. Values in brackets are for the most stable or best-known isotopes. This is explained further in topic 5.

Using electron shell configuration to identify an element

The best way to identify an element based on its electron shell configuration is to simply add up all of the electrons, because this is equal to the atomic (proton) number of the element, *Z*. Using the atomic number, you can then find both the period and group.

To determine the position of an atom, consider the following examples:

- $1s^22s^22p^63s^23p^64s^23d^6$ has 26 electrons, and so 26 protons. The element is iron. From the periodic table, we can see this is period 4 and group 8.
- $1s^22s^22p^63s^23p^3$ has 15 electrons, and so 15 protons. The element is phosphorus. From the periodic table, we can see this is period 3 and group 15.

To determine the position of an ion, simply add or remove the corresponding amount electrons (depending on whether it is a positive or negative ion) to give the electronic configuration of the atom; this is then equal to the atomic (proton) number of the element. Using the atomic number, you can then again find both the period and group.

For example, consider an ion with a charge of 2– and an electron configuration of $1s^22s^22p^6$. Because the ion has gained two electrons (to become 2–), simply subtract 2 electrons to give the electronic configuration of the atom. Therefore, $1s^22s^22p^6$, becomes $1s^22s^22p^4$ which has 8 electrons and therefore 8 protons, and the element is oxygen (period 3 and group 15). The original ion would be O^{2-} (oxide).

If an ion has a charge of 3+ and an electron configuration of $1s^22s^22p^63s^23p^63d^5$, because the ion has lost 3 electrons to become 3+, simply add 3 electrons to give the electronic configuration of the atom. Remember that the 4*s* subshell fills and empties before the 3*d* subshell. Therefore, $1s^22s^22p^63s^23p^63d^5$ becomes $1s^22s^22p^63s^23p^64s^23d^6$, which has 26 electrons, and therefore 26 protons, and the element is iron (period 4 and group 8). The original ion would be Fe^{3+}.

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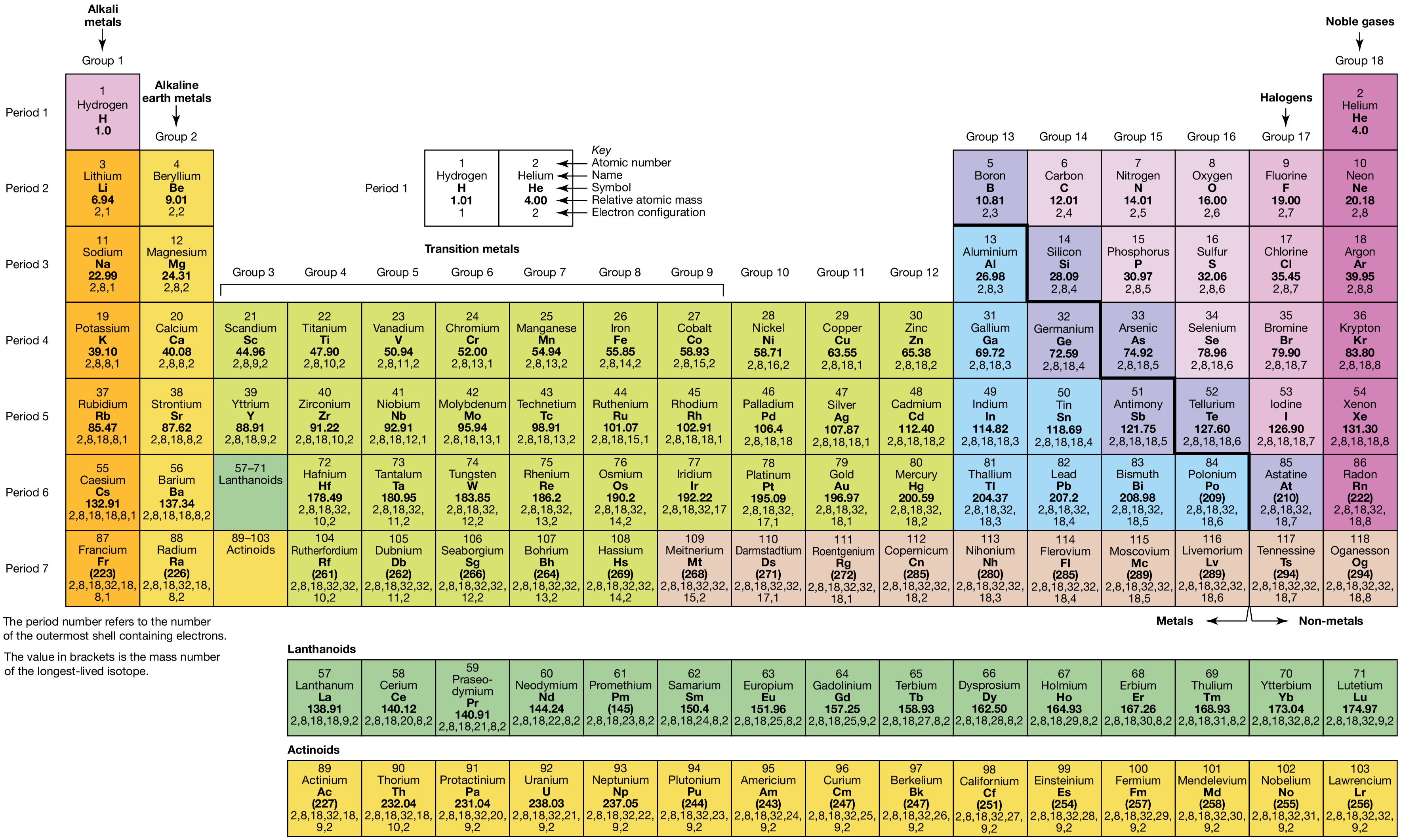
Weblink Electron configuration

FIGURE 2.6 Periodic table showing relative atomic masses to two decimal places



Key

Period 1	1 Hydrogen **H** **1.01** 1	2 Helium **He** **4.00** 2

	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	Group 9	Group 10	Group 11	Group 12	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
Period 1	1 Hydrogen **H** **1.0**																	2 Helium **He** **4.0**
Period 2	3 Lithium **Li** **6.94** 2,1	4 Beryllium **Be** **9.01** 2,2											5 Boron **B** **10.81** 2,3	6 Carbon **C** **12.01** 2,4	7 Nitrogen **N** **14.01** 2,5	8 Oxygen **O** **16.00** 2,6	9 Fluorine **F** **19.00** 2,7	10 Neon **Ne** **20.18** 2,8
Period 3	11 Sodium **Na** **22.99** 2,8,1	12 Magnesium **Mg** **24.31** 2,8,2											13 Aluminium **Al** **26.98** 2,8,3	14 Silicon **Si** **28.09** 2,8,4	15 Phosphorus **P** **30.97** 2,8,5	16 Sulfur **S** **32.06** 2,8,6	17 Chlorine **Cl** **35.45** 2,8,7	18 Argon **Ar** **39.95** 2,8,8
Period 4	19 Potassium **K** **39.10** 2,8,8,1	20 Calcium **Ca** **40.08** 2,8,8,2	21 Scandium **Sc** **44.96** 2,8,9,2	22 Titanium **Ti** **47.90** 2,8,10,2	23 Vanadium **V** **50.94** 2,8,11,2	24 Chromium **Cr** **52.00** 2,8,13,1	25 Manganese **Mn** **54.94** 2,8,13,2	26 Iron **Fe** **55.85** 2,8,14,2	27 Cobalt **Co** **58.93** 2,8,15,2	28 Nickel **Ni** **58.71** 2,8,16,2	29 Copper **Cu** **63.55** 2,8,18,1	30 Zinc **Zn** **65.38** 2,8,18,2	31 Gallium **Ga** **69.72** 2,8,18,3	32 Germanium **Ge** **72.59** 2,8,18,4	33 Arsenic **As** **74.92** 2,8,18,5	34 Selenium **Se** **78.96** 2,8,18,6	35 Bromine **Br** **79.90** 2,8,18,7	36 Krypton **Kr** **83.80** 2,8,18,8
Period 5	37 Rubidium **Rb** **85.47** 2,8,18,8,1	38 Strontium **Sr** **87.62** 2,8,18,8,2	39 Yttrium **Y** **88.91** 2,8,18,9,2	40 Zirconium **Zr** **91.22** 2,8,18,10,2	41 Niobium **Nb** **92.91** 2,8,18,12,1	42 Molybdenum **Mo** **95.94** 2,8,18,13,1	43 Technetium **Tc** **98.91** 2,8,18,13,2	44 Ruthenium **Ru** **101.07** 2,8,18,15,1	45 Rhodium **Rh** **102.91** 2,8,18,18,1	46 Palladium **Pd** **106.4** 2,8,18,18	47 Silver **Ag** **107.87** 2,8,18,18,1	48 Cadmium **Cd** **112.40** 2,8,18,18,2	49 Indium **In** **114.82** 2,8,18,18,3	50 Tin **Sn** **118.69** 2,8,18,18,4	51 Antimony **Sb** **121.75** 2,8,18,18,5	52 Tellurium **Te** **127.60** 2,8,18,18,6	53 Iodine **I** **126.90** 2,8,18,18,7	54 Xenon **Xe** **131.30** 2,8,18,18,8
Period 6	55 Caesium **Cs** **132.91** 2,8,18,18,8,1	56 Barium **Ba** **137.34** 2,8,18,18,8,2	57–71 Lanthanoids	72 Hafnium **Hf** **178.49** 2,8,18,32,10,2	73 Tantalum **Ta** **180.95** 2,8,18,32,11,2	74 Tungsten **W** **183.85** 2,8,18,32,12,2	75 Rhenium **Re** **186.2** 2,8,18,32,13,2	76 Osmium **Os** **190.2** 2,8,18,32,14,2	77 Iridium **Ir** **192.22** 2,8,18,32,17	78 Platinum **Pt** **195.09** 2,8,18,32,17,1	79 Gold **Au** **196.97** 2,8,18,32,18,1	80 Mercury **Hg** **200.59** 2,8,18,32,18,2	81 Thallium **Tl** **204.37** 2,8,18,32,18,3	82 Lead **Pb** **207.2** 2,8,18,32,18,4	83 Bismuth **Bi** **208.98** 2,8,18,32,18,5	84 Polonium **Po** **(209)** 2,8,18,32,18,6	85 Astatine **At** **(210)** 2,8,18,32,18,7	86 Radon **Rn** **(222)** 2,8,18,32,18,8
Period 7	87 Francium **Fr** **(223)** 2,8,18,32,18,8,1	88 Radium **Ra** **(226)** 2,8,18,32,18,8,2	89–103 Actinoids	104 Rutherfordium **Rf** **(261)** 2,8,18,32,32,10,2	105 Dubnium **Db** **(262)** 2,8,18,32,32,11,2	106 Seaborgium **Sg** **(266)** 2,8,18,32,32,12,2	107 Bohrium **Bh** **(264)** 2,8,18,32,32,13,2	108 Hassium **Hs** **(269)** 2,8,18,32,32,14,2	109 Meitnerium **Mt** **(268)** 2,8,18,32,32,15,2	110 Darmstadtium **Ds** **(271)** 2,8,18,32,32,17,1	111 Roentgenium **Rg** **(272)** 2,8,18,32,32,18,1	112 Copernicum **Cn** **(285)** 2,8,18,32,32,18,2	113 Nihonium **Nh** **(280)** 2,8,18,32,32,18,3	114 Flerovium **Fl** **(285)** 2,8,18,32,32,18,4	115 Moscovium **Mc** **(289)** 2,8,18,32,32,18,5	116 Livermorium **Lv** **(289)** 2,8,18,32,32,18,6	117 Tennessine **Ts** **(294)** 2,8,18,32,32,18,7	118 Oganesson **Og** **(294)** 2,8,18,32,32,18,8

The period number refers to the number of the outermost shell containing electrons.

The value in brackets is the mass number of the longest-lived isotope.

Lanthanoids

57 Lanthanum **La** **138.91** 2,8,18,18,9,2	58 Cerium **Ce** **140.12** 2,8,18,20,8,2	59 Praseo-dymium **Pr** **140.91** 2,8,18,21,8,2	60 Neodymium **Nd** **144.24** 2,8,18,22,8,2	61 Promethium **Pm** **(145)** 2,8,18,23,8,2	62 Samarium **Sm** **150.4** 2,8,18,24,8,2	63 Europium **Eu** **151.96** 2,8,18,25,8,2	64 Gadolinium **Gd** **157.25** 2,8,18,25,9,2	65 Terbium **Tb** **158.93** 2,8,18,27,8,2	66 Dysprosium **Dy** **162.50** 2,8,18,28,8,2	67 Holmium **Ho** **164.93** 2,8,18,29,8,2	68 Erbium **Er** **167.26** 2,8,18,30,8,2	69 Thulium **Tm** **168.93** 2,8,18,31,8,2	70 Ytterbium **Yb** **173.04** 2,8,18,32,8,2	71 Lutetium **Lu** **174.97** 2,8,18,32,9,2

Actinoids

89 Actinium **Ac** **(227)** 2,8,18,32,18,9,2	90 Thorium **Th** **232.04** 2,8,18,32,18,10,2	91 Protactinium **Pa** **231.04** 2,8,18,32,20,9,2	92 Uranium **U** **238.03** 2,8,18,32,21,9,2	93 Neptunium **Np** **237.05** 2,8,18,32,22,9,2	94 Plutonium **Pu** **(244)** 2,8,18,32,23,9,2	95 Americium **Am** **(243)** 2,8,18,32,24,9,2	96 Curium **Cm** **(247)** 2,8,18,32,25,9,2	97 Berkelium **Bk** **(247)** 2,8,18,32,26,9,2	98 Californium **Cf** **(251)** 2,8,18,32,27,9,2	99 Einsteinium **Es** **(254)** 2,8,18,32,28,9,2	100 Fermium **Fm** **(257)** 2,8,18,32,29,9,2	101 Mendelevium **Md** **(258)** 2,8,18,32,30,9,2	102 Nobelium **No** **(255)** 2,8,18,32,31,9,2	103 Lawrencium **Lr** **(256)** 2,8,18,32,32,9,2

SAMPLE PROBLEM 1

An atom has the following electron configuration: $1s^22s^22p^63s^23p^64s^23d^3$. Identify the period, group and name of the element.

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THINK	WRITE
1. This is an atom, so the number of protons is equal to the number of electrons. Add the electrons to determine the atomic (proton) number, which can easily be found on the periodic table.	$2 + 2 + 6 + 2 + 6 + 2 + 3 = 23$ electrons
2. Identify the period and group by recognising the element in the periodic table.	The element is vanadium. The atom belongs to period 4 and group 5.

PRACTICE PROBLEM 1

An ion with a 3− charge has the following electron configuration: $1s^22s^22p^6$. Identify the period, group and name of the element.

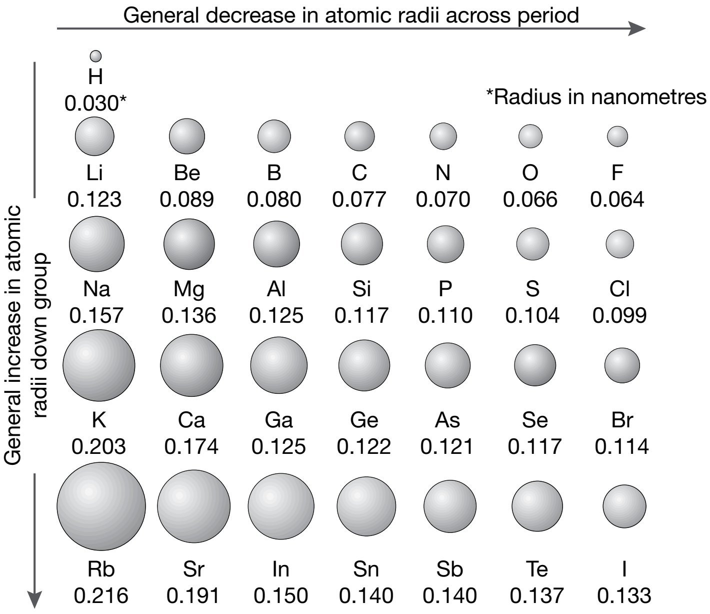
2.3.2 Atomic radii

Since an atom does not have a sharply defined boundary to set the limit of its size, the radius of an atom cannot be measured directly. However, several methods are available to gain an estimate of the relative sizes of atoms.

Atomic size generally increases down a group of the periodic table. Going down a group, electrons are added to successively higher energy levels, or main shells, further out from the nucleus. Because the number of positive charges in the nucleus also increases as you go down a group, the nuclear charge (attraction of positive charges in the nucleus to the electrons) increases. The inner electrons, however, create a 'shielding' effect, thereby decreasing the pull of the nucleus on the outermost electrons.

Atomic size generally decreases from left to right across a period. Across a period, each atom maintains the same number of main shells. Each element has one proton and one electron more than the preceding element. The electrons are added to the same main shell so the effect of the increasing nuclear charge on the outermost electrons is to pull them closer to the nucleus. Atomic size, therefore, decreases (see figure 2.7).

FIGURE 2.7 Atomic radii in nanometres of selected elements



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Interactivity Periodic table trends (int-6352)

2.3.3 Electronegativity

The **electronegativity** of an element is a measure of the degree to which an atom can attract an electron to itself. This is most evident when it is chemically combined with another element. The extent of attraction is expressed in arbitrary units on the Pauling electronegativity scale (see figure 2.8).

FIGURE 2.8 The Pauling scale of electronegativities

Increasing electronegativity →

Increasing electronegativity ↑

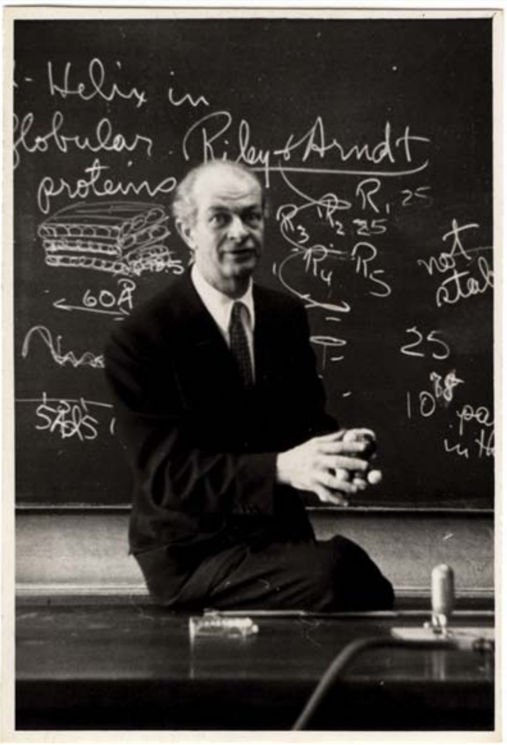
Group

1	2	13	14	15	16	17	18
H 2.1							He —
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne —
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar —
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr —
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe —
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn —
Fr 0.7	Ra 0.9						

Each element except the noble gases, which do not readily form compounds, is assigned an electronegativity number. Caesium and francium, the least electronegative elements, have a value of 0.7, whereas fluorine, the most electronegative element, has a value of 4.0.

Going across a period from left to right, the electronegativity of the main-group elements increases. This is because, as you move from one element to the next across a period, the nuclear charge increases by one unit, as one electron is added to the outer shell. As the positive charge in the nucleus increases, the atom has an increasing electron-attracting power and, therefore, an increasing electronegativity. This is the same reason the atomic radius tends to decrease across a period, as described previously. Moving down a group, the electronegativity decreases because the outer electrons are further away from the nucleus and the shielding effect of the inner electrons decreases the electron-attracting power of the atom. Similarly, this results in a general increase in atomic radii down a group.

FIGURE 2.9 Linus Carl Pauling (1901–1994) received the Nobel Prize for chemistry in 1954, and the Nobel Peace Prize in 1962.



2.3.4 First ionisation energy

When an atom gains or loses an electron, it forms an ion. The energy required to remove an electron from a gaseous atom is known as the first **ionisation energy**. Since the amount of energy required to do this is very small, it is more realistic to compare the amount of energy required to ionise one mole of atoms simultaneously. Therefore, the unit used is kilojoules per mole (kJ mol^{-1}).

Removal of one electron from a gaseous metal atom, M(g), results in the formation of a positive ion with a 1+ charge:

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

The energy required to remove this first (outermost) electron is called the first ionisation energy.

To remove the outermost electron from the gaseous 1+ ion results in a positive ion with a 2+ charge:

$$M^{+}(\mathrm{g}) \rightarrow M^{2+}(\mathrm{g}) + \mathrm{e}^{-}$$

The amount of energy required is called the second ionisation energy, and so on. Table 2.2 shows the first three ionisation energies of the first 20 elements in the periodic table.

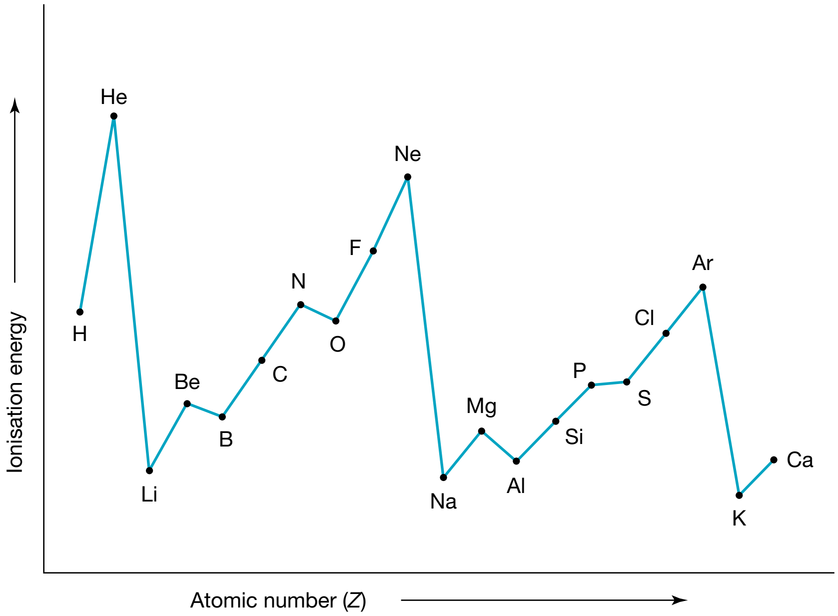
TABLE 2.2 Ionisation energies of the first 20 elements (kJ mol^{-1}). The red letters and numbers indicate the elements and first ionisation energies for group 1.

Symbol of element	Ionisation energy (kJ mol^{-1})		
	First	Second	Third
H	1312		
He	2371	5247	
Li	520	7297	11810
Be	900	1757	14840
B	800	2430	3659
C	1086	2352	4619
N	1402	2857	4577
O	1314	3391	5301
F	1681	3375	6045
Ne	2080	3963	6276
Na	495.8	4565	6912
Mg	737.6	1450	7732
Al	577.4	1816	2744
Si	786.2	1577	3229
P	1012	1896	2910
S	999.6	2260	3380
Cl	1255	2297	3850
Ar	1520	2665	3947
K	418.8	3069	4600
Ca	589.5	1146	4941

In general, the first ionisation energy decreases moving down a group of the periodic table. Since the size of the atoms is increasing when moving down a group, the outermost electrons are further from the nucleus. The nucleus, therefore, does not hold these electrons as strongly, so they are more easily removed. The atom, therefore, has a lower ionisation energy.

The first ionisation energy generally increases as we move from left to right across a period. The nuclear charge is increasing, whereas the shielding effect is relatively constant. A greater attraction of the nucleus for the electron, therefore, leads to an increase in ionisation energy. Ionisation energy is high for noble gases such as He, Ne and Ar. Table 2.2 shows the change in ionisation energies for Li, Na and K.

FIGURE 2.10 Trends in ionisation energy of the first 20 elements



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Weblink Periodic table patterns

2.3.5 Metallic/non-metallic characteristics

In terms of electronic structure, the metallic characteristics of an element are determined by its ease in losing electrons. As elements move across a period, they lose their metallic characteristics. This is because, as the number of electrons in the same shell increases across a period and the nuclear charge also increases, the electrons become less easily lost to form positive ions. As elements move down a group, they become more metallic because their outermost shell electrons are further away from the nucleus (due to an increased number of shells) and are less strongly attracted. Hence, the elements lose their outermost shell electrons more easily. These metallic/non-metallic characteristics are illustrated in the variation in melting and boiling points (metals typically have high melting and boiling points in comparison with non-metals), as shown in table 2.3.

TABLE 2.3 The melting points (t_m) and boiling points (t_b) of selected elements

Group 1			Group 2			Group 13			Group 14			Group 15			Group 16			Group 17		
	t_m (°C)	t_b (°C)		t_m (°C)	t_b (°C)		t_m (°C)	t_b (°C)		t_m (°C)	t_b (°C)		t_m (°C)	t_b (°C)		t_m (°C)	t_b (°C)		t_m (°C)	t_b (°C)
Li	180	1320	**Be**	1283	3000	**B**	2030	2550*	**C**	3600	4800	**N**	–210	–196	**O**	–218	–183	**F**	–220	–188
Na	98	890	**Mg**	650	1100	**Al**	660	2500	**Si**	1400	2400	**P**	44	280	**S**	113	444	**Cl**	–101	–35
K	63	770	**Ca**	850	1500	**Ga**	30	2400	**Ge**	940	2800	**As**	820	613*	**Se**	220	685	**Br**	–7	59
Rb	39	690	**Sr**	770	1400	**In**	157	2000	**Sn**	232	2300	**Sb**	630	1380	**Te**	450	1390	**I**	113	184*
Cs	29	690	**Ba**	710	1140	**Tl**	304	1460	**Pb**	327	1750	**Bi**	271	1560						
Low melting points that decrease down the group. Group 1 metals are soft and have low densities.			Much higher melting points of the first elements due to some covalent character in the bonds between atoms. Boron is a metalloid made up of a giant ring structure with covalent bonds.						High melting points that decrease down the group. C, Si and Ge exist as giant covalent networks of great hardness and high melting point.			Elements at the top of groups 15 and 16 and all the elements of group 17 exist as small covalent molecules. The melting points and boiling points increase down the groups as metallic bonding takes over in groups 15 and 16 elements.								

*Sublimes

2.3.6 Reactivity of elements

The oxidising strength of an element can be defined as how readily an element gains electrons. Elements that gain electrons easily are strong oxidising agents (also called oxidants) and are themselves reduced. Likewise, the reducing strength of an element is defined as how readily an element loses electrons.

The more readily an element gives up its electrons, the more easily it is oxidised, making it a stronger reducing agent (reductant) because it has more reducing strength. As elements move across a period, the reducing strength decreases as the atoms give up their outermost shell electrons less readily, and the oxidising strength of these elements increases as elements gain electrons more readily. The extreme in oxidising/reducing behaviour of elements across the periods can be seen in examples such as sodium and potassium metals, which give up their electrons very readily, and the non-metals fluorine and chlorine, which prefer to hold on to their electrons. Hence, sodium and potassium are strong reducing agents while fluorine and chlorine are strong oxidising agents. Going down a group, the elements release their electrons more readily, making them stronger reducing agents (the reducing strength increases). For example, potassium is a stronger reducing agent than sodium and is therefore more reactive.

2.3.7 Summary of trends in the periodic table

TABLE 2.4 Trends in the periodic table

	Across a period		Down a group	
Metallic character	↓	The attraction force increases with the increasing number of protons.	↑	Electrons are less attracted to the nucleus because they are shielded by the increasing number of shells.
Atomic size	↓	The attraction force increases with the increasing number of protons and so pulls the electrons closer.	↑	There is an increasing number of electron shells.
Reactivity	↓	Reactivity is high at the start of the period, less in the middle and more reactive at the end.	↑	Outer electrons are less strongly attracted.
Electronegativity	↑	Fluorine is the most electronegative element. The further away an element is from fluorine in the periodic table, the less electronegative it is.	↓	Increasing numbers of inner shells shield outermost shell electrons from the nucleus.

2.3 EXERCISE

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1. Identify the period and group to which each of the following elements belongs and name the element.
   (a) $1s^22s^22p^4$
   (b) $1s^22s^1$
   (c) $1s^22s^22p^63s^23p^64s^23d^6$
   (d) $1s^22s^22p^63s^23p^64s^23d^{10}4p^6$
   (e) $1s^22s^22p^63s^23p^64s^23d^1$
   (f) $1s^22s^22p^63s^23p^64s^23d^{10}$
2. Identify the period and group to which each of the following ions belongs and name the element.
   (a) An ion with a charge of 7+ and the electronic configuration of $1s^22s^22p^63s^23p^6$
   (b) An ion with a charge of 1– and the electronic configuration of $1s^22s^22p^63s^23p^6$

3. Write the ground-state electron configuration for elements with the following atomic numbers, and state the group and period of each element.
   (a) 8
   (b) 13
   (c) 20
   (d) 35
4. (a) Zinc is an element that forms mostly colourless compounds. Write the subshell configuration for zinc ($Z = 30$).
   (b) State two reasons zinc does not fit into the usual pattern for transition elements.
5. Describe and explain the general trends in the following.
   (a) Electronegativity across the periodic table
   (b) The atomic radii of elements down a group
   (c) The reducing strength of elements across a period
6. Explain why the periodic table includes the following.
   (a) Two elements in the first period
   (b) Eight elements in the second period
   (c) No transition elements in the first three periods
7. For each of the following pairs of elements, state which element is the more electronegative.
   (a) K, Ca
   (b) Be, Ca
   (c) Cl, Br

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2.4 Review

2.4.1 Summary

Organisation of the periodic table

- The periodic table is a method of organising all the known elements to show their similarities and differences.
- Historically, the development of the periodic table was based on the classification of elements according to their chemical and physical properties.
- The organisation of elements on the periodic table was changed from an order of increasing atomic weight to increasing atomic number.
- Elements after uranium (atomic number 92) are artificially synthesised and radioactive. They are known as the transuranium elements.
- To date, elements up to 118 have been made. Transuranium elements are extremely unstable.

Patterns and trends in the periodic table

- Elements arranged down the same vertical columns (groups) in the modern periodic table display similar physical and chemical properties.
- Elements arranged along the same horizontal rows (periods) are placed in order of increasing atomic number.
- The main features of the periodic table are the:
  - eight main groups, which progressively fill both the *s* and *p* subshells
  - transition elements, which progressively fill the *d* subshells
  - rare earth elements, which progressively fill the *f* subshells. Elements of the *f* block are made up of the lanthanoids and the actinoids.

- The fundamental structure of the periodic table as developed by Mendeleev has remained largely unchanged, despite the discoveries and developments of new theories of atomic structure.
- Metals are mostly found on the left side and the middle of the periodic table, separated by the metalloids from the non-metals, which are found on the right.
- Atomic size decreases across the periodic table and increases down the groups.
- Electronegativity generally increases across a period and decreases down a group.
- Metallic character decreases across a period and increases down a group.
- Reducing strength decreases across a period and increases down a group. Oxidising strength increases across a period and decreases down a group.

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2.4.2 Key terms

electronegativity the electron-attracting power of an atom
groups the vertical columns of elements in the periodic table
ionisation energy the energy required to remove an electron from a gaseous atom or ion
metal ductile, malleable, conducting element
non-metal non-ductile, non-malleable, non-conducting element
periods the seven horizontal rows in the periodic table

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Digital document Key terms glossary — Topic 2 (doc-30921)

Note: No practical investigation logbook

2.4 Exercises

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2.4 Exercise 1: Multiple choice questions

1. In constructing his initial forms of the periodic table, Mendeleev placed sodium and potassium in the same group because these two elements
   A. have the same atomic mass.
   B. have the same number of electrons.
   C. react violently with water.
   D. have metallic looks about them.
2. Which one of the following statements about the periodic table is correct?
   A. All the elements listed on the periodic table are naturally occurring.
   B. The periodic table can be used to predict the physical and chemical properties of undiscovered elements.
   C. Elements with atomic number over 95 are radioactive.
   D. The modern periodic table is arranged in order of atomic mass.

3. Which of the following elements is expected to be most similar in properties to magnesium?
   A. Aluminium
   B. Sulfur
   C. Potassium
   D. Strontium
4. Based on trends in the periodic table, which of the following elements would have the smallest atomic radius?
   A. Silicon
   B. Fluorine
   C. Calcium
   D. Beryllium
5. As you move across the periodic table, which of the following is generally true?
   A. Electronegativity decreases
   B. Ionisation energy decreases
   C. Metallic characteristics decrease
   D. Atomic radius increases
6. What is a trend as you go down the periodic table?
   A. The size of atoms increases
   B. The metallic characteristics decrease
   C. Oxidising strength increases
   D. Electronegativity increases

Refer to the following table to answer questions 7–10.

Element	Charge on atom or ion	Atomic number	Electron configuration of atom or ion
A	*A*		$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$
B	B^-	9	
C	C^{2-}	16	
D	D^{3+}		$1s^2 2s^2 2p^6$

7. Element *A* is in
   A. period 4 and is a transition element.
   B. period 4, group 1.
   C. period 4, group 13.
   D. period 3, group 13.
8. What is element *B*?
   A. Na
   B. Mg
   C. F
   D. O
9. What is the ground-state electron configuration for the element *C*?
   A. $1s^2 2s^2 2p^6 3s^2 3p^2 4s^2$
   B. $1s^2 2s^2 2p^6 3s^2 3p^4$
   C. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
   D. $1s^2 2s^2 2p^6 3s^2 3p^6$
10. What is the atomic number of the ion of element *D*?
   A. 10
   B. 7
   C. 13
   D. 6

2.4 Exercise 2: Short answer questions

1. Use the list of elements to answer the questions that follow: Cl, Ni, N, Mg, I, B, Na.
   a. Choose the two elements that have the most similar chemical properties and explain your choice.
   b. Which of the elements is the most reactive metal? Explain why.
   c. Which element is a transition metal?
   d. Which of the elements is the most reactive non-metal?
   e. Which element has the highest electronegativity? Explain why.
2. a. Outline the contribution of Dimitri Mendeleev to the development of the first periodic table.
   b. Mendeleev arranged his periodic table according to increasing atomic mass, yet he placed tellurium before iodine, which has a smaller atomic mass. Explain why he made this decision.
3. Why are elements in a periodic table arranged in order of atomic *number* rather than atomic *mass?*
4. Predict which pair of elements in each of the following sets would have the greatest similarities in chemical properties and which would have the greatest differences.
   a. Na and Cl, Na and K, Na and Ca
   b. Cl and I, Cl and S, Cl and Mg
5. Explain why it is easier for elements to lose electrons going down a group.
6. Consider the element phosphorus ($Z = 15$).
   a. What is its ground state electron configuration?
   b. Is it a metal or non-metal?
   c. In what group is it?
   d. In what period would you find it?
   e. In what block would you find it?
   f. Name an element that would have similar properties to phosphorus.
   g. Would it have a higher or lower electronegativity than chlorine?
   h. Would it have a larger or smaller atomic size than nitrogen?
7. The relative atomic mass of argon is 39.9. It is placed before potassium in the periodic table even though the relative atomic mass of potassium is 39.1. Explain why they are placed in these positions.
8. a. Explain what is meant by the term ‘nuclear charge’.
   b. Which has the higher nuclear charge: potassium or calcium?
9. a. Going across the periodic table, the numbers of protons and electrons increase. Why then does the size of the atoms decrease?
   b. Explain the trend in atomic radius going down a group.
10. How is an element’s outer electron configuration related to its position in the periodic table? Give three examples that illustrate your answer.

2.4 Exercise 3: Exam practice questions

Question 1 (3 marks)

a. What is meant by the term ‘electronegativity’? **1 mark**

b. Explain how electronegativity is related to ionisation energy **2 marks**

Question 2 (4 marks)

Why do metals generally have low electronegativities, whereas non-metals have high electronegativities?

Question 3 (7 marks)

The following table represents part of the periodic table. The letters shown represent some of the elements but they may not be the symbols of those elements.

																	E
A	*T*											*M*	*X*	*L*		*G*	
												Q	*D*		*R*	*J*	
				B													

Selecting only from the elements labelled on the periodic table shown, write the letter (*A, B, D, E, L, G, J, M, Q, R, T* or *X*) corresponding to the element with the following.

a. A non-metal with four electrons in its outer shell **1 mark**

b. In group 14, period 2 **1 mark**

c. A transition element **1 mark**

d. A halogen with two occupied shells **1 mark**

e. In period 2 and with a total of three electrons **1 mark**

f. A halogen **1 mark**

g. An alkaline earth metal **1 mark**

2.4 Exercise 4: studyON Topic Test online only

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3 Metals

3.1 Overview

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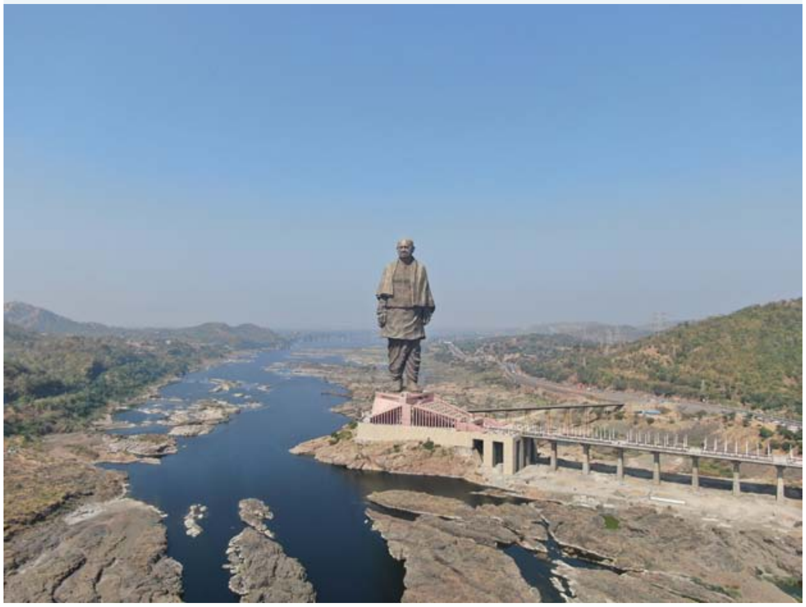
3.1.1 Introduction

It has been known for centuries that elements can be divided into two broad groups: metals and non-metals. The largest group is the metals. The first metal to gain widespread use was copper, which was first mined and used on the island of Cyprus around 5000 years ago. Copper's popularity was mainly due to its relatively low melting point (1084 °C) and the ease with which it could be extracted from its ores. In contrast, aluminium, which is the most common metal in the Earth's crust, was difficult and expensive to extract from its ore and has become widespread only in the last century due to technological advances.

FIGURE 3.1 The Statue of Unity in Gujurat, India, is the world's largest statue at 182 metres tall. It contains 25 000 tonnes of steel to re-enforce the concrete structure and 3550 tonnes of bronze to decorate the outer surface.



3.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- the common properties of metals (lustre, malleability, ductility, heat and electrical conductivity) with reference to the nature of metallic bonding and the structure of metallic crystals, including limitations of representations, and general differences between properties of main group and transition group metals

- experimental determination of the relative reactivity of metals with water, acids and oxygen
- the extraction of a selected metal from its ore/s, including relevant environmental, economic and social issues associated with its extraction and use
- experimental modification of a selected metal related to the use of coatings or heat treatment or alloy production
- properties and uses of metallic nanomaterials and their different nanoforms, including comparison with the properties of their corresponding bulk materials.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of learning and assessment. Experiments and investigations, supported by a **Practical investigation logbook** and **Teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

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Digital documents Key science skills (doc-30903)
Key terms glossary — Topic 3 (doc-30924)
Practical investigation logbook (doc-30925)

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3.2 Properties of metals

KEY CONCEPT

- The common properties of metals (lustre, malleability, ductility, heat and electrical conductivity) with reference to the nature of metallic bonding and the structure of metallic crystals, including limitations of representations, and general differences between properties of main group and transition group metals

3.2.1 The common properties of metal

Since more than 80 metal elements exist, it is not surprising that some of their properties vary widely. Group 1 metals, such as lithium, sodium and potassium, must be stored in oil (paraffin or mineral) because they react violently with water, whereas metals such as copper, silver and gold are chemically unreactive and are used in coins and jewellery. At room temperature, mercury is a liquid, but magnesium is a solid. Zinc is brittle, iron is hard, and lead sheet can be bent easily. Nevertheless, most metals have the following properties.

Metals:
- are lustrous (have a shiny surface when polished)
- are malleable (can be hammered, bent or rolled into sheets or other shapes)
- are ductile (can be drawn out into wires)
- are good conductors of heat
- are good conductors of electricity
- generally have a high density (mass per unit volume)
- have a range of melting points but most are quite high
- are often hard (have high resistance to denting, scratching and bending).

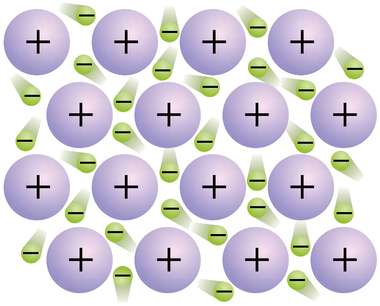
3.2.2 Crystal nature of metallic bonding

Metallic atoms have low electronegativities, which means they tend to lose their outer shell electrons easily. Once a metallic atom has lost its outer shell electron, it becomes a **cation** (positively charged). This produces electron configurations similar to the group 18 elements (noble gases) and hence a more stable state.

X-ray studies show that metals occur as crystal lattices. Chemists imagine these **metallic lattice structures** as being made up of a patterned array of cations (figure 3.2). The electrons from each metallic atom overlap each other forming a sea of mobile electrons that can flow between all the cations. They are referred to as **delocalised** electrons (figure 3.3).

FIGURE 3.2 This box of oranges gives a representation of a repeated pattern array of cations in a lattice structure.

FIGURE 3.3 Metal cations vibrate about a fixed position surrounded by a sea of delocalised electrons.



Electrostatic forces of attraction between the cations and the negatively charged electrons occur in all directions and this holds the lattice together. This type of non-directional bonding is known as **metallic bonding**. This means that metal atoms are hard to separate but relatively easy to move.

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Weblinks Structure of metals

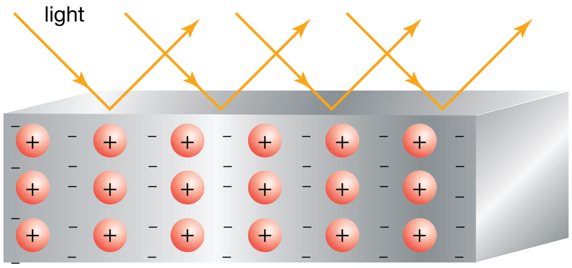
Lattice structure

3.2.3 The structure and properties of metallic crystals

The lattice structure of metals may be used to explain many of their properties. It is these properties that give metals their many applications in our society.

Lustre

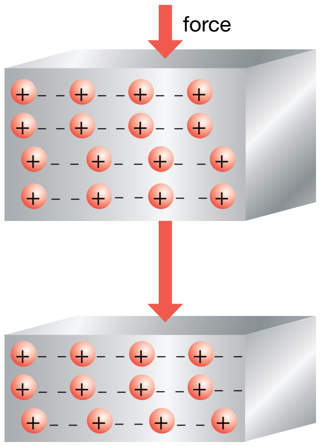
The **lustrous** appearance of a metal is due to the mobile electrons within the lattice being able to reflect light back into to your eye, causing the metal to look shiny.



Malleability and ductility

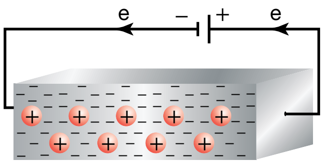
Metals are **malleable** and **ductile**, rather than brittle, as a result of the non-directional nature of metallic bonds.

The attractive forces exerted by the cations for the mobile electrons occur in all directions. This means that layers of atoms can move past one another without disrupting the force between the cations and the negative sea of electrons. The nature of the metal does not change when the metal becomes thinner.



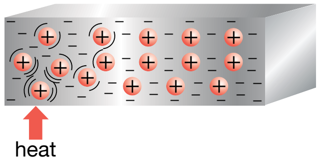
Electrical conductivity

Metals are good conductors of electricity. When an electric field is applied to a metal, one end of the metal becomes positive and the other becomes negative. All the electrons experience a force towards the positive end. This flow of charge (movement of electrons) is what we call an electric current.



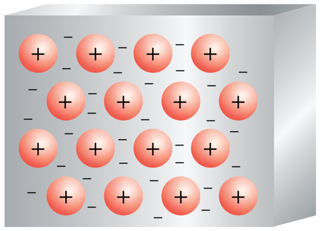
Heat conductivity

Metals are good conductors of heat. Electrons gain kinetic energy in hotter areas of the metal and quickly transfer it to other colder parts of the metal lattice due to the electron's freedom of movement. The heat causes the electrons to move faster, and the 'bumping' of these electrons against each other and the protons transfers the heat throughout the metal.



Density

Most metals have relatively high densities because metallic lattices are close-packed.



Melting points and hardness

The generally high melting points and hardness of metals indicate that metallic bonding is quite strong (although not as strong as covalent or ionic bonding). Melting points and hardness increase with an increase in the number of outer shell electrons, since a greater attractive force exists between the cations and the electrons.

FIGURE 3.4 Steel can be pressed and rolled into sheets, as at this factory. What metallic property is illustrated here?

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Weblink Treatment of steel

3.2.4 Limitations of models

Although many of the properties of metals can be explained using the lattice model, other properties are not explained. For example:

- Why is iron magnetic and copper non-magnetic?
- Why is platinum twenty times as dense as sodium?
- Why is lead malleable whereas iron is tough?

Different models may be used to explain some of these properties. One model is the **ball bearing model**.

Metal strength — the ball bearing model

Some metals bend easily whereas others are strong. When a metal bends, layers of atoms slide over each other. The more easily they slide, the more easily the metal bends. A metal with perfect rows of atoms bends easily. If the rows are distorted, they cannot slide over each other as freely, so the metal does not bend as easily.

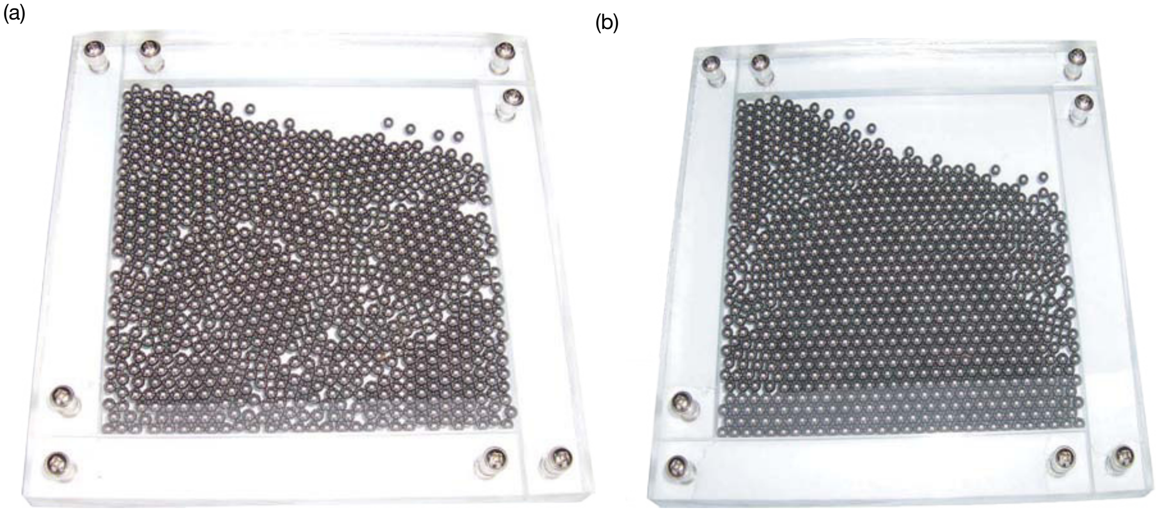
Metals do not crystallise with perfect rows of atoms throughout the lattice. They form areas of perfect close-packing called grains. The boundaries between the grains are called grain boundaries. Although the atoms in metal grains are packed in a regular order, the grains themselves are irregularly shaped crystals of the metal pushed tightly together.

The structure of a metal may be modelled by using ball bearings to show the arrangement of cations into grains. The model shows that:

- Small areas of perfect packing simulate the arrangement of metal cations in a single crystal grain. This metal would not bend easily (see figure 3.5a).
- Grains are large areas of regular arrangement; they represent a single metal crystal. Grain boundaries are narrow areas of disorder seen between one crystal grain and another. Vacant sites occur when an atom is missing from the crystal structure. Dislocations happen when a row of atoms is displaced and the regular packing stops. Metals with large grains have fewer dislocations and they bend more easily (see figure 3.5b).

The crystal structure, and therefore the properties, of metals can be changed in a number of ways. The next section discusses some of these.

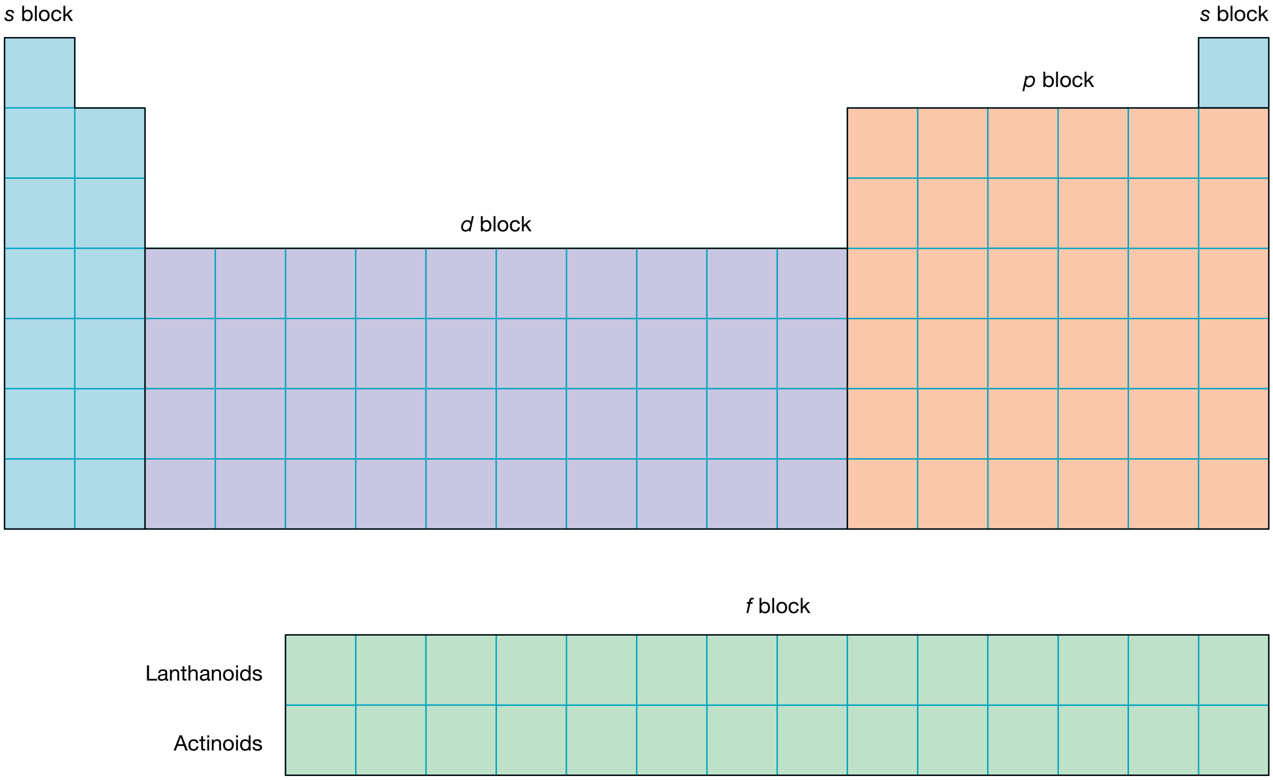
FIGURE 3.5 Demonstration of the ball-bearing model: (a) small areas of perfect packing and (b) large areas of regular arrangement; they represent a single metal crystal.



3.2.5 Properties of *s*-block metals

Metals are widely used in our society. Their different properties make them useful for particular purposes. (Figure 3.6 shows the blocks of the periodic table.)

FIGURE 3.6 The blocks of the periodic table



The *s*-block metals include the alkali metals. They include the elements lithium, sodium and potassium, which are very reactive; they are not found in elemental form and so most uses involve their compounds. A knife can easily cut through these metals because they have such a low density. They have relatively low melting points due to having only one outershell electron to participate in bonding. Also, the atoms are larger than other atoms in the same period, and they are not as efficiently packed together. Alkaline earth metals, such as beryllium, magnesium and calcium, are less reactive, harder and have higher melting points than alkali metals. Usually, these reactive elements are found as compounds. The elements and compounds have a variety of applications. The *s*-block metals usually form compounds that are white.

Sodium is used as a coolant in nuclear reactors, in the manufacture of titanium and for street lights. Lithium is used in rechargeable batteries and, because of its lightness, it is combined with other metals to make alloys. Magnesium is alloyed because it produces superior qualities, and can then be used for aircraft and guided missile parts where lightness and high tensile strength are essential. The compound calcium carbonate is used to make calcium oxide for neutralising soils, glass making and as a component of cement. The properties of the *s*-block metals are summarised in table 3.1.

FIGURE 3.7 Sodium has a low density (as do other group 1 and 2 metals) and can be cut with a knife.

TABLE 3.1 Properties of metals in the *s* block

Metal	Symbol	Density (g mL^{-1})	Melting point (°C)	Thermal conductivity at 25 °C (watts $m^{-1}K^{-1}$)	Electrical conductivity at 25 °C ($ohm^{-1}m^{-1} \times 10^7$)	Hardness
Lithium	Li	0.53	181	85	1.1	Soft
Sodium	Na	0.97	98	140	2.1	Soft
Potassium	K	0.66	64	100	1.4	Soft
Beryllium	Be	1.86	1278	190	3.1	Hard
Magnesium	Mg	1.75	649	160	2.3	Medium
Calcium	Ca	1.55	839	200	3.0	Medium

3.2.6 Properties of *p*-block metals

Cans are manufactured from aluminium or tin-coated steel. Aluminium and tin are metals found in the *p* block of the periodic table, which also contains metalloids and non-metals. Aluminium is the most abundant metal in the Earth's crust and is very versatile. When exposed to air it develops a coating of aluminium oxide, Al_2O_3, which prevents further reactions with oxygen, and hence further corrosion (see topic 14). Aluminium is used in transport, building and food storage such as drink cans and saucepans. The good electrical conductivity and low density of aluminium make it ideal for use in overhead wires. Tin is used as a coating to prevent corrosion, such as in 'tin' cans, which are made from tin-coated steel. The alloys solder, bronze and pewter all contain tin.

FIGURE 3.8 Bullets have an outer layer of copper alloy and a core of lead. Lead is used because it is dense, readily available and easy to shape.

Lead, another *p*-block metal, has been used since Roman times because it is easily worked and resistant to corrosion. Lead was used for water pipes and to line aqueducts to deliver water to homes and public baths. It has been used previously in pottery glazes and paint pigments, in jewellery and as a petrol additive. Due to health hazards, these uses for lead have significantly declined in recent decades. Current uses include car batteries, roofing, exercise weights and bullets.

TABLE 3.2 Properties of metals in the *p* block

Metal	Symbol	Density (g mL^{-1})	Melting point (°C)	Thermal conductivity at 25 °C (watts $m^{-1}K^{-1}$)	Electrical conductivity at 25 °C ($ohm^{-1}m^{-1} \times 10^7$)	Hardness
Aluminium	Al	2.7	660	236	3.8	Medium
Tin	Sn	7.3	232	66.8	0.9	Medium
Lead	Pb	11.3	327	35.3	0.5	Soft

3.2.7 Properties of *d*-block metals

The presence of transition metal compounds is responsible for the colourful stained glass windows in churches and other buildings. The formation of coloured compounds is an interesting characteristic of the *d*-block elements; exceptions are scandium and zinc. The metals that display magnetism — iron, cobalt and nickel — are transition metals. As with other metals, most are used in the form of alloys. They can form ions with different charges; for example, iron forms two compounds with chlorine — $FeCl_2$, which is a green compound, and $FeCl_3$, which is an orange compound. Another valuable feature of transition metals is that they are very useful catalysts. A catalyst is a substance that increases the rate of a chemical reaction without itself being chemically changed.

FIGURE 3.9 Transition elements form coloured solutions.

Transition metals are an ever-present part of our daily lives, from the iron in the oxygen transport molecule, haemoglobin, in our blood, to the $2 coins in our pockets, which are 92% copper. Iron is widely used in construction, often in alloy form as steel, which is less brittle and more resistant to corrosion than elemental iron. It is used in the manufacture of tools and vehicles and is a catalyst in the production of ammonia. Copper is used in electrical wiring and as bases on saucepans because of its good electrical and heat conduction. Interestingly, copper forms the basis of blood in crustaceans, which results in it being a blue colour. Silver and gold, as well as being used for jewellery, are good conductors of electricity and are used in electrical contacts. Apart from its presence in baked beans, the element nickel is used in making stainless steel, in batteries, in coins and as a catalyst in the manufacture of margarine.

TABLE 3.3 Properties of transition metals

Metal	Symbol	Density (g mL^{-1})	Melting point (°C)	Thermal conductivity at 25 °C (watts $m^{-1}K^{-1}$)	Electrical conductivity at 25 °C ($ohm^{-1}m^{-1} \times 10^7$)	Hardness
Iron	Fe	7.9	1535	75	1.0	Hard
Copper	Cu	8.9	1083	401	6.0	Hard
Silver	Ag	10.5	961	429	6.3	Medium
Gold	Au	19.3	1059	318	4.3	Soft

3.2 EXERCISE

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1. Lithium exists as a crystalline solid at temperatures lower than 180 °C.
   (a) Write the electron configuration of a lithium atom.
   (b) Write the symbol and electron configuration of its cation.
   (c) Describe how lithium atoms are bonded together in the solid.
2. Determine which of the atomic numbers below belong to a metal.
   (a) 2
   (b) 11
   (c) 17
   (d) 19
   (e) 25
3. Why is the structure of a metal stable? Illustrate your answer using the example of aluminium and its electron configuration.
4. Explain the following properties of iron.
   (a) Hard
   (b) Malleable and ductile
   (c) Able to conduct electricity in solid and molten form
5. Magnesium is produced from the mineral dolomite, $CaMg(CO_3)_2$.
   (a) Describe the bonding in magnesium metal crystals.
   (b) Magnesium crystals are silver-white and shiny. Explain this property of lustre in terms of structure and bonding.
6. What are some of the limitations of the 'delocalised sea' of electrons lattice model for metals?
7. How does the ball bearing model explain why some metals bend more easily than others?
8. (a) State the charge on ions formed from group 1 elements.
   (b) State the charge on ions formed from group 2 elements.
9. Use the properties of sodium and potassium to predict the approximate values of the melting point and thermal conductivity of rubidium, the next element in group 1.
10. Explain why aluminium is used for drink cans whereas 'tin' cans are used for other foods.
11. Give three reasons silver and gold are used for jewellery.
12. How do (a) gold, (b) mercury and (c) zinc differ from the majority of transition metals?
13. Choose a metal that would be best suited for fabrication of the following.
    (a) Car bodies
    (b) Beer cans
    (c) Household electrical wiring
    (d) Fishing sinkers
    Justify your choices.
14. Electricity transmission lines for industry are usually made from aluminium even though copper is a better electrical conductor. Suggest two reasons aluminium is used.
15. Explain the following properties of metals in terms of the metallic structure.
    (a) Metals are conductors of heat and electricity.
    (b) Metals are malleable.
    (c) Metals are ductile.
    (d) Metals are shiny.
    (e) Metals have high density.

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3.3 Reactivity of metals

KEY CONCEPT

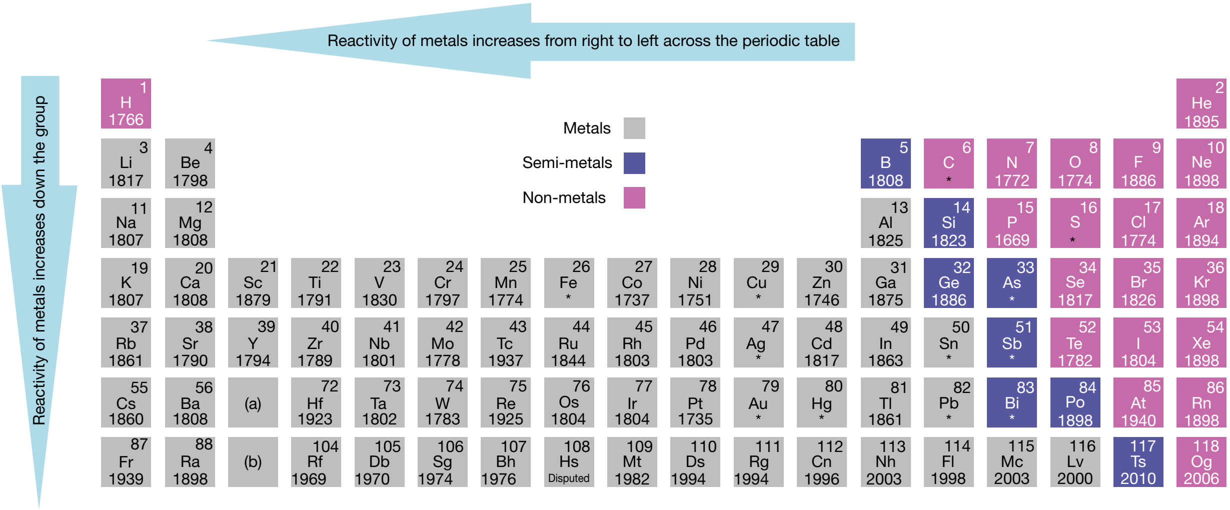
- Experimental determination of the relative reactivity of metals with water, acids and oxygen

The majority of elements found on the periodic table (around 80%) are metallic. The reactivity of metals depends on their ability to lose electrons and form positive ions (cations). The energy required to form cations is measured as ionisation energy (see section 2.3.4). This chemical property distinguishes metals from semi-metals and non-metals because these do not form cations. Being able to predict the products of reactions between metals and their compounds is key to understanding how different metals are extracted from their natural sources (ores).

3.3.1 Reactivity of metals and the periodic table

Figure 3.10 shows the general trends in reactivity of the metallic elements found in the periodic table.

FIGURE 3.10 General trends in reactivity in the metallic elements.



1 H 1766																	2 He 1895
3 Li 1817	4 Be 1798											5 B 1808	6 C *	7 N 1772	8 O 1774	9 F 1886	10 Ne 1898
11 Na 1807	12 Mg 1808											13 Al 1825	14 Si 1823	15 P 1669	16 S *	17 Cl 1774	18 Ar 1894
19 K 1807	20 Ca 1808	21 Sc 1879	22 Ti 1791	23 V 1830	24 Cr 1797	25 Mn 1774	26 Fe *	27 Co 1737	28 Ni 1751	29 Cu *	30 Zn 1746	31 Ga 1875	32 Ge 1886	33 As *	34 Se 1817	35 Br 1826	36 Kr 1898
37 Rb 1861	38 Sr 1790	39 Y 1794	40 Zr 1789	41 Nb 1801	42 Mo 1778	43 Tc 1937	44 Ru 1844	45 Rh 1803	46 Pd 1803	47 Ag *	48 Cd 1817	49 In 1863	50 Sn *	51 Sb *	52 Te 1782	53 I 1804	54 Xe 1898
55 Cs 1860	56 Ba 1808	(a)	72 Hf 1923	73 Ta 1802	74 W 1783	75 Re 1925	76 Os 1804	77 Ir 1804	78 Pt 1735	79 Au *	80 Hg *	81 Tl 1861	82 Pb *	83 Bi *	84 Po 1898	85 At 1940	86 Rn 1898
87 Fr 1939	88 Ra 1898	(b)	104 Rf 1969	105 Db 1970	106 Sg 1974	107 Bh 1976	108 Hs Disputed	109 Mt 1982	110 Ds 1994	111 Rg 1994	112 Cn 1996	113 Nh 2003	114 Fl 1998	115 Mc 2003	116 Lv 2000	117 Ts 2010	118 Og 2006

It is possible to determine qualitatively the reactivity of samples of different metals by investigating experimentally how readily they react with oxygen, cold water, steam and dilute acids (a summary of the reactions can be found in section 3.3.7).

3.3.2 Relative reactivity of metals with water

As figure 3.10 indicates, group 1 (alkali metals) and group 2 (alkaline earth metals), are highly reactive, and some of these metals react with water at room temperature to produce alkalis and hydrogen gas. (Alkalis are metal hydroxides that are soluble in water and form solutions that neutralise acids; see figure 3.11 topic 13.) These highly reactive metals must be stored under oil to prevent their reaction with atmospheric water vapour.

FIGURE 3.11 Although the lattice structure of metals is strong, they can still react — as sodium does on contact with water.

For example: Group 1 metal + water → group 1 metal hydroxide + hydrogen gas
potassium + water → potassium hydroxide + hydrogen gas
$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$

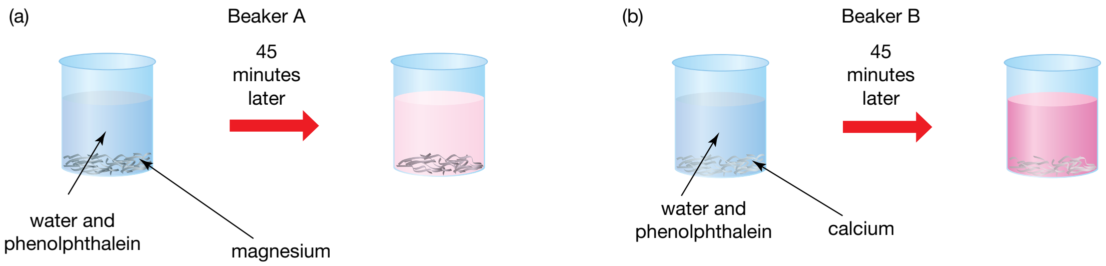
For example: Group 2 metal + steam → group 2 metal hydroxide + hydrogen gas
magnesium + steam → magnesium hydroxide + hydrogen gas
$Mg(s) + 2H_2O(g) \rightarrow Mg(OH)_2(aq) + H_2(g)$

on Resources

Weblink Introduction to the periodic table: alkali metals
Video eLesson Reaction of potassium with water (eles-3209)

The experiment illustrated in figure 3.12 shows how the relative reactivity of magnesium and calcium can be observed. When equal amounts of freshly sanded magnesium ribbon and strips of oil-free calcium react with water in beakers A and B as shown, hydrogen gas and a hydroxide compound are formed. The presence of the hydroxide compound is indicated by phenolphthalein. After 45 minutes, beaker B has a more intense pink colour, suggesting that more hydroxide compound has been produced by the reaction. We can deduce qualitatively from this that calcium is more reactive with water than is magnesium.

FIGURE 3.12 Demonstration of relative reactivity of (a) magnesium and (b) calcium.



In general terms, if a metal reacts with cold water, a hydroxide and hydrogen gas is formed, whereas if they only react with steam, they tend to form metal oxides as the high temperatures decompose the hydroxides into the oxide.

SAMPLE PROBLEM 1

Why is it necessary for magnesium to be freshly sanded for a reaction to occur with cold water?

Teacher-led video: SP1 (tlvd-0513)

THINK	WRITE
Magnesium, and some other metals, appear to be unreactive, but this is not always the case. Consider what happens to the magnesium when it is exposed to air that may affect the reaction with cold water.	Metals like magnesium (and especially aluminium) react readily with atmospheric oxygen, forming a protective coating of the metal oxide. This metal oxide coating acts as a barrier, preventing water from coming into direct contact with the metal and, therefore, no reaction or a slow reaction occurs. Sanding the magnesium allows the water to come into contact with the pure metal, allowing a reaction to occur.

PRACTICE PROBLEM 1

Why is it necessary for calcium to be oil-free for a reaction to occur?

3.3.3 Relative reactivity of metals with acids

Hydrogen is also produced when the more reactive metals react with acids.

For example: Group 1 metal + acid → salt + hydrogen
lithium + hydrochloric acid → lithium chloride + hydrogen
$2Li(s) + 2HCl(aq) \rightarrow 2LiCl(aq) + H_2(g)$

For example: Group 2 metal + acid → salt + hydrogen
magnesium + sulfuric acid → magnesium sulfate + hydrogen
$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

on Resources

Video eLesson Reaction of iron with acid (eles-3210)

3.3.4 Relative reactivity of metals with oxygen

The most reactive metals are found on the very left side of the periodic table and, as figure 3.10 shows, reactivity increases down the groups. Some metals, such as sodium, are so reactive that they react with the oxygen in air and need to be stored under oil to prevent this reaction (see figure 3.13). See topic 14 for more on redox ('reduction–oxidation') reactions.

FIGURE 3.13 Sodium reacting violently with pure oxygen.

For example: Group 1 metal + oxygen → metal oxide
sodium + oxygen → sodium oxide
$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$

For example: Group 2 metal + oxygen → metal oxide
calcium + oxygen → calcium oxide
$2Ca(s) + O_2(g) \rightarrow 2CaO(s)$

SAMPLE PROBLEM 2

Figure 3.13 demonstrates the reaction of sodium and oxygen to produce sodium oxide. Assuming a small amount of air can enter the glass jar, what is the equation that shows the formation of carbon when sodium reacts with a very small amount of air?

Teacher-led video: SP2 (tlvd-0514)

THINK	**WRITE**
Pure oxygen is trapped in the combustion jar, but when the molten sodium is placed inside the jar, air has a chance to enter.	
Even though air contains only 0.04% CO_2, the sodium is more reactive and is oxidised (loses electrons) and the carbon is reduced (gains electrons).	$4Na(l) + CO_2(g) \rightarrow 2Na_2O(s) + C(s)$

PRACTICE PROBLEM 2

A small spatula of magnesium powder was added to a beaker that contained dilute nitric acid solution, HNO_3. Bubbles were initially observed. When bubbles were no longer observed, more magnesium powder was added and the reaction again produced bubbles. This was repeated until magnesium powder remained in the beaker and no further bubbles were observed.

When a flame was placed at the mouth of the test tube, the gas burned with a squeaky pop.

Answer the following:

a. **Identify the gas produced.**
b. **Give the name of the colourless solution formed in this reaction.**
c. **Suggest why the reaction stopped.**

3.3.5 Summary of the reactivity of metals

TABLE 3.4 Reactivity series of metals

Element	Appearance of metal	Reaction with oxygen	Reaction with water or steam	Reaction with dilute acids
K	Dull, stored under oil	Oxidises in air at room temperature to give oxides	Hydrogen formed from cold water	Hydrogen formed violently
Na				
Ca	Generally dull			Hydrogen formed with cold acid
Mg			Hydrogen formed with steam	
Al				
Zn				
Fe				
Sn		Oxidises when heated in air or oxygen to give oxides	No reaction at bunsen burner temperatures	
Pb				No visible reaction
Cu	Generally shiny			No reaction
Hg			No reaction	
Ag		No reaction		
Au				

Note: aluminium metal often appears to be less reactive than the reactivity series indicates because it has a coating of aluminium oxide and this protects it from reacting further with oxygen. This coating can be thickened by a process called anodising.

on Resources

Digital documents Experiment 3.1 Metallic trees (doc-30869)

Experiment 3.2 The underwater garden (doc-30870)

Video eLesson Experiment 3.2 The underwater garden (tlvd-0616)

3.3 EXERCISE

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1. Write the balanced equation (including state symbols) for the reaction of a sample of pure calcium metal with room temperature deionised water.
2. Write fully balanced equations, including states, for the following reactions.
   (a) Potassium + oxygen
   (b) Caesium + water
   (c) Tin + nitric acid, HNO_3

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3.4 Extraction of metal from its ore

KEY CONCEPT

- The extraction of a selected metal from its ore/s, including relevant environmental, economic and social issues associated with its extraction and use

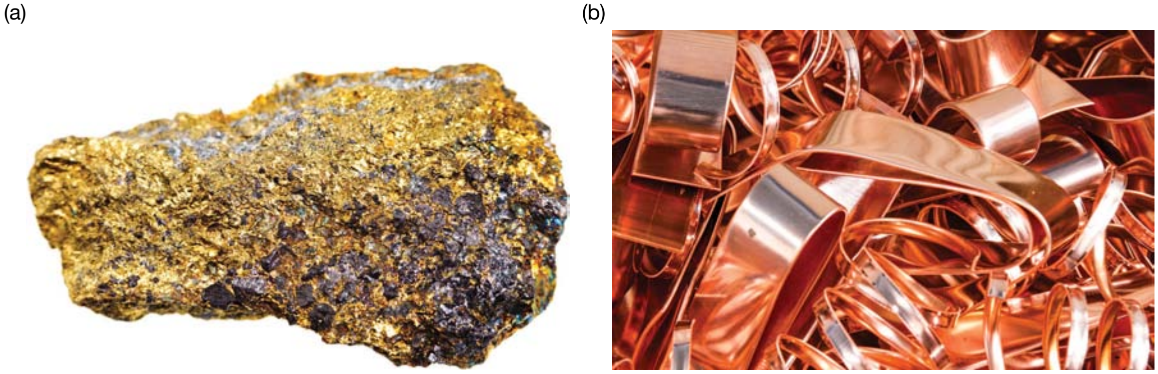
3.4.1 Reactivity and extraction method

The method used to extract metals from their ores depends on their reactivity. The reactivity of a metal also influences how easily it is extracted (see table 3.4). Silver and gold, for example, are unreactive and can often be found naturally in their elemental state, which is why they have been known and used since ancient times. For example, gold may be panned in stream beds. More reactive metals must be extracted from their ores. Ores are rocks that are mixtures of different compounds and from which it is economically viable to separate out the metal. Copper, for example, which ranks second only to iron in its importance in human history, is extracted from the ore chalcopyrite (see figure 3.14)

TABLE 3.5 Reactivity metals and type of extraction

Reactivity	Metal	Extraction
Very reactive	K, Na, Ca, Mg, Al	Electrolysis of molten compound, using electricity to cause a chemical reaction
Reactive	Zn, Fe, Sn, Pb	Heating the metal ore with carbon
Less reactive	Cu, Hg, Ag	Heating the metal ore in air
Unreactive	Ag, Pt, Au	Can occur as free elements

FIGURE 3.14 (a) Copper can be extracted from the ore chalcopyrite. (b) Today the main use of copper is in electrical cables, pipes and wires, shown here ready for recycling.



3.4.2 The extraction of iron metal from its ore

Our society is highly dependent on iron. Iron is needed to manufacture steel for transport, construction, machinery, containers and appliances. About 5% of the Earth's crust is iron; it is the fourth most abundant element after oxygen, silicon and aluminium. Because iron is a metal of medium reactivity, it is extracted from its ores by reaction with carbon in a blast furnace. Iron ore is composed of a number of different minerals (iron-rich compounds), from which iron is extracted. Iron ores include magnetite, Fe_3O_4, haematite, Fe_2O_3, pyrite, FeS, goethite, FeO(OH), limonite, $FeO_3(OH).n(H_2O)$, and siderite, $FeCO_3$.

The blast furnace

FIGURE 3.15 Molten iron from a blast furnace.

The first stage in the manufacture of iron or steel is preparing the ore (the rock that contains the iron) for the blast furnace. It has to be crushed to fist-sized lumps of rock. These rocks are heated to drive off any water or other impurities that will evaporate.

The second stage involves the reduction of iron oxides to iron in the blast furnace. To do this, carbon in the form of coke and limestone (calcium carbonate) are added.

The furnace itself is a tapered cylindrical tower about 40 metres high. It is made of steel and is lined with bricks.

A mixture of iron ore, coke and limestone is added at the top of the furnace in such a way that no gas escapes while it is being added. At the same time, blasts of hot air at about 720 °C are blown into the furnace through small holes near the base. The purpose of the air is to burn the coke to form carbon monoxide. The carbon monoxide reduces the iron oxide to iron in a reaction that also produces carbon dioxide and great quantities of heat.

The iron falls to the bottom of the furnace and eventually melts at about 1500 °C. It is drained off every few hours.

The role of the limestone is to remove impurities inside the furnace. Impurities such as silica, SiO_2, and alumina, Al_2O_3, are removed. The substances formed are known as 'slag'. The slag floats on top of the iron and is drained off at a different level.

The iron obtained from a blast furnace is not completely pure and is called 'pig-iron'. It is hard and brittle and melts at about 1300 °C.

The following outlines the steps in this process.

1. The coke reacts with oxygen in the air, giving carbon dioxide:

$$C\,(s) + O_2\,(g) \rightarrow CO_2\,(g)$$

2. The limestone decomposes to calcium oxide and carbon dioxide:

$$CaCO_3\,(s) \rightarrow CaO\,(s) + CO_2\,(g)$$

3. The carbon dioxide reacts with more coke, giving carbon monoxide:

$$C\,(s) + CO_2\,(g) \rightarrow 2CO\,(g)$$

4. This reacts with iron oxide in the ore, giving liquid iron:

$$Fe_2O_3\,(s) + 3CO\,(g) \rightarrow 2Fe\,(l) + 3CO_2\,(g)$$

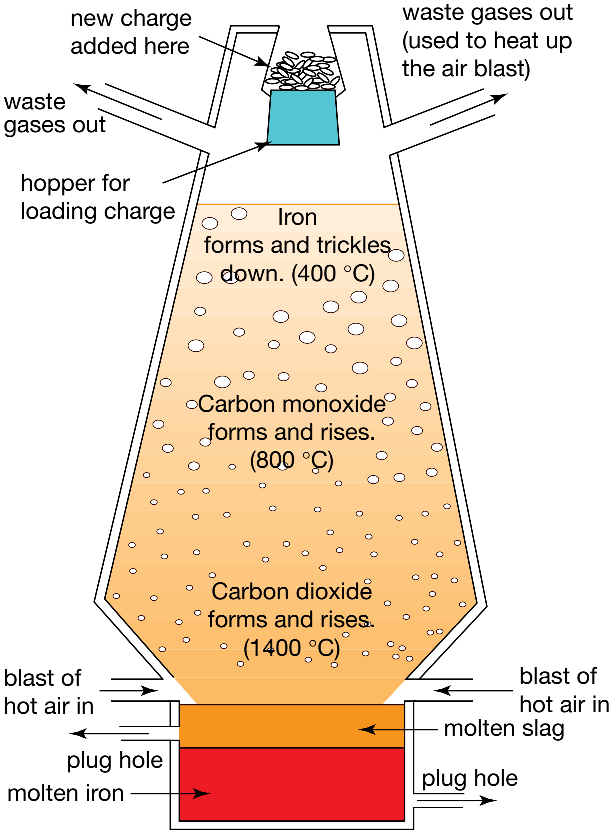
   The iron trickles to the bottom of the furnace.

5. Calcium oxide from step 2 reacts with sand in the ore, to form calcium silicate or slag:

$$CaO\,(s) + SiO_2 \rightarrow CaSiO_3\,(l)$$

The slag runs down the furnace and floats on the iron. The slag and iron are drained from the bottom of the furnace.

FIGURE 3.16 Iron production in a blast furnace



3.4.3 Impacts of iron production on society, the economy and social issues

The benefits to society of iron production include not only providing essential materials but also offering employment in mining, production, transport and sales. Exporting iron ore contributes to the economy and, therefore, the wealth of the nation. These benefits must be balanced with the costs of obtaining and transporting the raw materials, and the energy requirements for the process. The cost to the environment may also be considerable. The destruction of the landscape for open-cut mining, the significant waste materials that must be disposed of, and the particulate matter and waste gases all contribute to environmental damage. The waste carbon dioxide contributes to the greenhouse effect, and sulfur dioxide gas contributes to acid rain. Recycling iron can reduce some of these problems. Less energy is used to recycle metals than the original process, and resources are conserved, resulting in lower costs and less pollution. Collecting waste gases and focusing on revegetation after mining can also reduce some of the environmental impact.

FIGURE 3.17 Australia is one of the world's largest producers of iron ore; most of it is mined in Western Australia.

3.4 EXERCISE

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1. State the chemical name for the iron compounds obtained from the different iron ores.
2. Find out what methods are used by industry to minimise the effects on the environment of iron production.
3. What differences would you notice in your daily life if iron were not available?

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3.5 Modifying metals

KEY CONCEPT

- Experimental modification of a selected metal related to the use of coatings or heat treatment or alloy production

3.5.1 Work-hardening metals

Bending or hammering cold metals causes their crystal grains to become smaller. Since bending is now made more difficult, the metal is toughened, or **work hardened**. This is easily seen by bending a metal coathanger and then trying to bend it back to its original shape. It does not bend back in the area that has been work hardened.

Work-hardened metals are usually more brittle, owing to the increased number of dislocations in the crystal structure. Constantly bending a piece of metal backwards and forwards causes it to snap since more and more dislocations are being produced in the metal crystal.

FIGURE 3.18 Hammering cold metals makes them harder.

3.5.2 Modifying metals using heat

The neatness of the rows of atoms in a metal may be disrupted by heat. This can occur in three main ways:

- **Annealed** metals are heated until they are red hot and then cooled slowly. Larger metal crystals are formed so that the metal produced is softer. Annealing is an intermediate step in metal-forming processes. It is used to restore ductility lost due to work hardening in items such as the coiled copper pipes in hot water systems.
- **Quenched** metals are heated until they are red hot and then allowed to cool quickly in cold water. Smaller crystals are formed, which makes the metal harder but more brittle. For example, blacksmiths formerly quenched horseshoes by plunging them into cold water.
- **Tempered** metals are produced when quenched metals are warmed again to a lower temperature and then allowed to cool slowly. This procedure reduces the brittleness of the metal while retaining the hardness. For example, metals used to make special wood-carving tools such as lathes and chisels, as well as axes and rock drills, are tempered.

FIGURE 3.19 The rate at which a roll of hot iron cools will determine its structural properties.

Quenching, annealing and tempering alter the properties of a metal because these processes disrupt the metal lattice.

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Digital document Experiment 3.3 The effect of heat on metals (doc-30871)

3.5.3 Coating metals

Metals may be coated for decorative or protective reasons, or to make them more useful, such as non-stick frypans, chrome-plated pipes and jewellery. Generally, however, metals, particularly iron, are coated to prevent corrosion.

Surface protection

Surface protection is used to prevent the air and water from coming into contact with the metal. A number of methods may be used for surface protection of iron and steel:

- Plastic is the most common form of surface protection. It is used to cover items ranging from simple household coathangers to commercial cables and wiring.
- Paint is used for many large objects, such as ships and bridges (figure 3.20).
- Grease or oil may be used to coat the moving parts of machinery.
- Metal coatings are used on steel.

FIGURE 3.20 Exposed surfaces on the Sydney Harbour Bridge must be repainted every five years. But with a surface area of nearly 500 000m^2, painters are permanently employed to ensure that the Sydney Harbour Bridge is protected against corrosion.

Two types of such coatings are used: one that is less reactive than the metal it is coating and one that is more reactive.

A coating using a metal that is less reactive than steel is called **noble coating**, and it serves purely as a covering. Any scratch or break in the coating results in rapid corrosion of the exposed steel. Noble coating is used where scratching is unlikely and where zinc coating is not suitable (because zinc salts are poisonous). Steel food containers are plated with tin, for example, and chromium is used as a noble coating on steel parts in cars and household items.

A coating using a metal that is more reactive than steel is called a **sacrificial coating** (see topic 14). A break in a sacrificial coating results in the formation of an electrochemical cell. The coating corrodes and the steel is protected. Zinc is the most common form of sacrificial coating. Objects coated in this way are said to be **galvanised**. When used on an iron roof, for example, the zinc coating reacts with carbon dioxide in air to form a layer of basic zinc carbonate, $Zn(OH)_2.ZnCO_3$, over its surface, thus protecting it and slowing its rate of corrosion. Prepainted galvanised steel sheets for roofs, walls and household appliances are now commercially available. Australians make extensive use of galvanised iron for roofing, guttering, pipes, rubbish bins, fencing wire and nails.

FIGURE 3.21 A worker pressing galvanised sheet steel at a steelworks. The zinc prevents the steel from rusting even if the surface is scratched or broken. This is because the more reactive zinc loses electrons, rather than the iron in the steel.

3.5.4 Alloys

Alloys are used because not all the properties of a metal may be suited to a specific use. By mixing a metal with other metals or some non-metals, we can change its properties. This process is called **alloying**.

Alloy production

Alloys are prepared by melting metals, with or without other materials together, and then cooling the mixture. The metals are a physical mixture because metals do not form any metallic bonds with different metals. Most of the metallic materials in everyday use are alloys rather than pure metals. Sometimes, non-metallic atoms such as carbon and silicon are used to form an alloy.

FIGURE 3.22 Polarised light micrograph of a thin section of common brass, a copper–zinc alloy, showing the distinct grain structure of this metal alloy. The grain boundaries — the three-dimensional surfaces that enclose grains — show up as fine lines around the polygonal grains.

Alloying metals can change their properties to make them more suitable for different purposes. A common alloy is gold. Pure gold (24 carat) is a relatively soft element and is easily deformed. We, therefore, use the alloy 18 carat gold in jewellery, because it only contains 75% (by mass) gold, with the other 25% being made up of silver and copper to make the product stronger and more durable. The term 'carat' is used to indicate the ratio of precious metal to base metal in an alloy. For example, 18 carat gold is $\frac{18}{24}$ or three-quarters gold to one-quarter base metal, and 24 carat gold is $\frac{24}{24}$ or pure gold.

Magnesium is unsuitable for industrial use because of its high reactivity. Yet, when alloyed with other metals, such as aluminium, zinc and manganese, it offers both lightness and strength, and can be used for aircraft and car wheels, artificial limbs and skis.

FIGURE 3.23 (a) Jewellery made from gold alloys is harder than pure gold. (b) Magnesium alloys demonstrate considerable lightness and high tensile strength, hence their use in car wheels.

(a)

(b)

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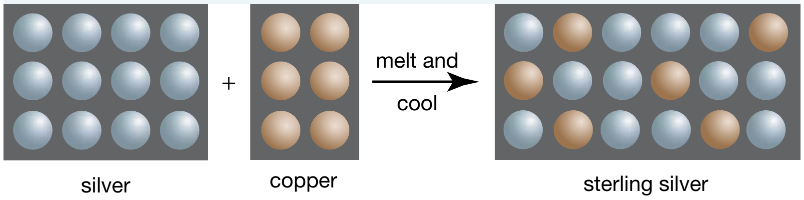
Weblink Alloys

Types of alloys

If the atoms of the metals being used to form an alloy are about the same size, they can replace each other in the metal crystals. This type of alloy is called a **substitutional alloy**. Sterling silver is a substitutional alloy made from silver and copper.

FIGURE 3.24 Sterling silver is a substitutional alloy.



Sometimes the atomic sizes of the metals in the alloy differ greatly. The smaller atoms may then fit into the spaces between the larger atoms, forming an **interstitial alloy**. Steel is an interstitial alloy made from iron and carbon.

The properties of alloys differ from those of their component metals. For example, sterling silver is harder than pure silver, and stainless steel is stronger and more resistant to corrosion than iron. Solder is composed of 70% lead (melting point 327 °C) and 30% tin (melting point 232 °C). It is used to join metals together because it melts at 180 °C, lower than either of its components. An alloy of mercury is called amalgam. Some dental fillings are amalgams of mercury (50%), silver (35%) and tin (13%) with copper and zinc. The amalgam filling starts as a pliable substance but undergoes a chemical reaction, causing it to harden after it is used to fill a cavity.

FIGURE 3.25 Steel is an interstitial alloy.

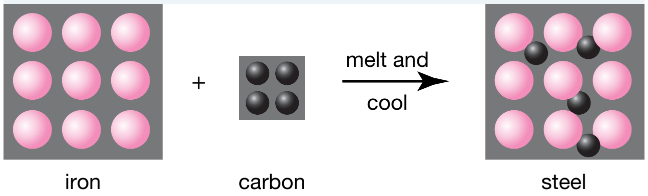


FIGURE 3.26 Dental fillings can be made from an alloy of mercury called amalgam.

3.5 EXERCISE

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1. Samurai swords were tempered to make them stronger and harder.
   (a) Describe how tempering alters the properties of a metal.
   (b) Describe two other methods that use heat to modify the properties of a metal.
2. Explain why work-hardened copper requires annealing before it may be further bent into a coiled shape.
3. When a metal paperclip is straightened out, it is difficult to bend it back to its original shape. Why?
4. Solder (melting point 180 °C) is an alloy made from 30% tin (melting point 232 °C) and 70% lead (melting point 327 °C). It is commonly used to join metal surfaces or parts.
   (a) Describe how solder could be made.
   (b) The atoms of tin and lead are similar in size. Decide which model (interstitial or substitutional) better represents the structure of solder.
   (c) What properties of solder differ from that of tin and lead? Relate this to its use.

5. Automatic fire sprinkler plugs are used in most modern buildings. When they melt, they release water onto the fire. The plugs are made from an alloy of tin, lead, bismuth and cadmium, each of which melts at a temperature well over 200 °C. What properties would you expect this alloy to exhibit, and how do these properties differ from those of the metals from which it is made? Explain your answer.

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3.6 Metallic nanomaterials

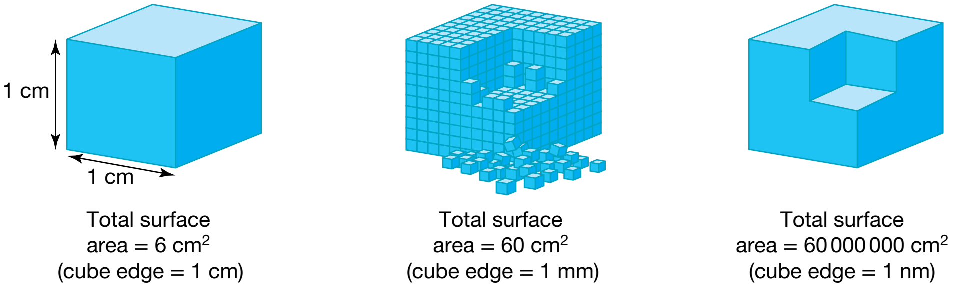
KEY CONCEPT

- Properties and uses of metallic nanomaterials and their different nanoforms, including comparison with the properties of their corresponding bulk materials

3.6.1 Properties of metallic nanomaterials

The reactivity of a metal is influenced by the size of its particles. The exceptionally small size of nanoparticles results in a relatively large surface area. This very high surface area to volume ratio of nanoparticles increases their reactivity and catalytic ability because of the greater contact with reactant particles. This means the properties of metallic nanoparticles are different from those of their corresponding bulk materials.

FIGURE 3.27 As the size of the particles decreases, the surface area increases. This has significant effects on the properties of nanomaterials.



Nanoparticles have a diameter between 1 and 100 nm. In just 1 millimetre, there are 1 million nanometres. Metallic nanomaterials are not new. Nanogold has been used in Roman pottery and in medieval stained-glass windows, as have nanocopper and nanosilver. Scientists are interested in their huge potential in biomedical science and engineering.

3.6.2 Uses of metallic nanomaterials

The antibacterial properties of nanosilver means it has many uses; socks can be permeated with nanosilver, for example, and worn for extended periods of time. Perhaps this is not an entirely appealing thought, but these antibacterial properties mean that much of the odour is eliminated. Some clothing, food packaging, surfaces, appliances, bedding and even baby bottles have also undergone the same treatment. Nanosilver's antimicrobial properties also make it useful for bandages, dressings and surfaces in hospitals. Its proven effectiveness is due to its extremely large surface area, a characteristic of nanoparticles in general. However, there is evidence of increasing bacterial resistance to nanosilver and concern about the effects of its use on the environment.

3.6.3 Comparison of nanomaterials and bulk materials

The properties discussed through most of this topic have referred to what are described as bulk materials or those at the macro level. At the macro level, solid samples can contain billions of atoms, whereas a nanoparticle has between 10 and 70 000 atoms. Metallic nanomaterials have the same metallic structure but the number of atoms in a sample is very much smaller. This results in not enough outer shell electrons to act as a cloud of delocalised electrons, which creates different properties. At the nanoscale, many of the classical laws of physics do not operate; it is here that quantum effects apply.

In the future, metallic nanotechnology has significant potential to contribute to smaller and faster electronics, biotechnology, the environment and industry. Understanding the implications of nanotechnology is essential to minimising risks to health and the environment.

3.6.4 Different nanoforms of metals and their properties

Nanomaterials look different from and react faster than bulk samples. They may be stronger and better conductors of electricity; they are good catalysts. In bulk metals, the high density of electrons on the surface causes light to be fully reflected. When the particles are much smaller, some parts of visible light are absorbed, resulting in the appearance of a colour that depends on the size of the nanoparticles. Nanosilver is yellow; nanogold is red for particles of about 20 nm diameter but can be orange, pink or even blue if the particle size is different (see figure 3.28). Bulk gold boils at 1064 °C but nanogold boils at a few hundred degrees. Nanogold is used in electronic chips, and in diagnosis and treatment of disease including cancer treatment, and drug or gene delivery. It is also an effective catalyst. Examples of nanotechnology in medicine can also be found in topic 1.

FIGURE 3.28 The colour of nanogold depends on the size of the clusters of particles; in the samples shown, the smallest particles are on the left and the largest on the right.

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Weblink Gold and nanogold in medicine

TITANIUM BICYCLES

Materials research has become an important component of competition preparation in many sports. Track and field athletes look for running shoes that provide stability, support and track grip. Tennis players seek the ultimate racquet that will give them a larger 'sweet spot' (the optimum position on the mesh from which to hit the ball), greater power and more control. Bicycle frames have until recently been built from steel, aluminium and carbon fibre composites. A new material has now emerged as the preferred frame material — titanium. (Titanium bike frames were first made in 1986 but only in recent years has the cost of these frames dropped sufficiently to make them a viable alternative for cyclists.)

Each frame material has its advantages and disadvantages, as shown in table 3.6.

FIGURE 3.29 Titanium mountain bike frames are also becoming popular.

TABLE 3.6 Materials used in bicycle frames

Material	Advantages	Disadvantages
Aluminium	• Light • Stiff	• Low resistance to repeated stresses
Steel	• Strong • Economical • Easily shaped to required specifications	• Rusts • Heavy
Carbon fibre composites	• High strength-to-mass ratio • Good shock-dampening power • Good vibration-dampening power	• Very expensive
Titanium	• Less than half the density of steel • Excellent resistance to metal fatigue • High corrosion resistance	• Expensive • Difficult to shape

Titanium is preferred by cyclists for use in bicycle frames due to its combination of toughness, elasticity and resilience. Bicycles that are built stiffly, so that they resist pedalling forces, often provide a harsh, uncomfortable ride. Although titanium frames are very stiff against high pedalling forces, they seem to transmit much less road shock than bicycles made of other materials. Steel, for example, has a significantly higher density than titanium, but shock waves travel faster in steel than in titanium. Despite the fact that titanium's high vibration-dampening properties are not yet fully understood, we do know that titanium provides three things crucial to cyclists: low weight, stiffness and a smooth ride.

Titanium is a lustrous, silvery metal with a high melting point (1667 °C). Found as an ore, titanium is quite abundant in the Earth's crust, ranking ninth of all the elements. Metals are usually extracted from their ores by heating them with carbon. The extraction of titanium, however, is quite difficult since carbon cannot be used due to the formation of titanium carbides. Titanium's unusual ability to stretch makes it hard to shape by machine. It tends to push away from even a very sharp cutting blade, giving a rather unpredictable final edge.

However, the superior bicycle that results is worth all these difficulties.

3.6 EXERCISE

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1. Consider figure 3.27.
   (a) What is the volume to surface area ratio of a cube with sides 1 mm in length?
   (b) What is the volume to surface area ratio of a cube with sides 1 nm in length?
   (c) How many times larger is the surface area of a cube of 1nm sides compared to a cube of 1 mm sides?
2. Why do nanomaterials have different properties to bulk materials?
3. Read the weblink **Gold and nanogold in medicine.** What properties of quantum dots allow them to be used in medicine?

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3.7 Review

3.7.1 Summary

Properties of metals

- When metallic atoms of the same element combine, a lattice is formed consisting of cations in a sea of delocalised electrons. The electrostatic attraction between the cations and electrons is called metallic bonding.
- The metallic lattice structure influences the properties of metals, which, in turn, determine their use.
- Metals are generally:
  - lustrous due to the mobile electrons being able to reflect light
  - good conductors of heat due to the mobile electrons
  - able to transfer the kinetic energy in hotter areas of the metal to cooler areas
  - good conductors of electricity due to the electrons being able to move when an electric field is applied
  - malleable and ductile because the attractive forces that bond the cations and mobile electrons occur in all directions, so that one layer of metal atoms can easily slip over another, without breaking the bonds
  - high-density substances because their lattices are close packed
  - hard and tough and have high tensile strength and high melting points due to the strong attraction between the electrons and cations.
- The properties of metals in different blocks of the periodic table show variations from the general properties.
  - *s*-block metals are reactive and some have lower density.
  - *p*-block metals have varied properties.
  - *d*-block metals usually form coloured compounds and are useful as catalysts.
- The ball bearing model was developed to explain different metal strengths.
  - Metals with large areas of perfect close packing (grains) have fewer dislocations (flaws in the packing) and bend easily.
  - Metals with small grains have many dislocations and do not bend easily.

Reactivity of metals

- Metals vary in their reactivity. Reactivity decreases across the periods and increases down a group of the periodic table.
- Reactions of metals with water

  Group 1 metal + water → group 1 metal hydroxide + hydrogen gas
  Group 2 metal + steam → group 2 metal hydroxide + hydrogen gas
- Reactions of metals with acids

  Group 1 metal + acid → salt + hydrogen gas
  Group 2 metal + acid → salt + hydrogen gas
- Reactions of metals with oxygen

  Group 1 metal + oxygen → metal oxide
  Group 2 metal + oxygen → metal oxide

Extraction of metal from its ore

- The method of metal extraction from its ore depends on the metal's reactivity. Iron is a reactive metal and is extracted by reaction with carbon in a blast furnace.
- Iron is an essential material for construction and manufacturing, but its production has impacts on the environment, including land degradation and pollution. These effects may be reduced by recycling, collecting waste gases and revegetation.

Modifying metals

- Metals can be modified by:
  - work hardening
  - different heat treatments
  - alloying
  - coating (used to improve usefulness and prevent corrosion).
- Work-hardened or toughened metals are more brittle because the crystal grains are smaller and have more dislocations.
- The three types of heat-modified metal are:
  - annealed metals, which are heated and then cooled slowly to produce larger crystals and hence a softer metal
  - quenched metals, which are heated until red hot and then cooled quickly in cold water to produce smaller crystals and hence a harder but more brittle metal
  - tempered metals, which are made by warming a quenched metal and allowing it to cool slowly to reduce brittleness but retain hardness.
- Metals may be coated with plastic, paint, oil or another metal to provide surface protection.
- The two types of metal coatings are:
  - noble coatings of less reactive metals, which serve as a physical barrier
  - sacrificial coatings of more reactive metals, which serve as a physical and chemical barrier, such as galvanising.
- Alloys are made by melting metals or other materials together and cooling the mixture.
  - The alloy has different properties from that of the original materials.
  - The two types of alloy are substitutional alloys, where the metal atoms are the same size, and interstitial alloys, where the metal atoms differ in size.

Metallic nanomaterials

- Because of their extremely small size and resulting large surface area, the properties of metallic nanoparticles are different from those of the corresponding bulk materials, including colour and reactivity.
- Nanosilver is a useful antibacterial agent.
- Nanogold has beneficial applications in disease detection and treatment.

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3.7.2 Key terms

alloy a substance with metallic properties that consists of two or more elements, at least one of which is a metal
alloying mixing of a metal with one or more other elements (usually metals)
annealed metal that has been heated in an inert gas to a moderate temperature, its temperature maintained for several hours, and then cooled slowly
ball bearing model representation of a metal using ball bearings to show the arrangement of cations into grains
cation a positively charged atom or group of atoms
delocalised describes electrons that are not bound to any one atom but are free to move throughout a lattice
ductile able to be drawn out into a thin wire
galvanised the coating of an iron-containing metal with zinc; the zinc forms a physical and chemical barrier to protect the iron because zinc is more reactive than iron
interstitial alloy an alloy in which the smaller atoms fit into the spaces between the larger atoms
lustrous having a shiny surface
malleable able to be hammered or pressed into shape without breaking or cracking
metallic bonding positively charged metal cations arranged in a lattice with delocalised valence electrons being able to flow around them
metallic lattice structure crystal lattice made up of an array of cations
noble coating protection of a metal from corrosion by attaching a thin metallic layer to it
quenched metal that has been plunged while hot into water to rapidly cool it
sacrificial coating protection of a metal from corrosion by attaching another more active metal to it; the more active metal reacts preferentially with corrosive elements and protects the less reactive metal
substitutional alloy an alloy in which the atoms of the different elements are about the same size, so they can replace each other in the metal crystals
tempered a quenched metal that is heated again but to a lower temperature than the initial heating
work hardened a property of metal brought about by beating the metal when cold

on Resources

Digital document Key terms glossary — Topic 3 (doc-30924)

3.7.3 Practical work and investigations

Experiment 3.2 online only

The underwater garden

Aim: To grow crystals using sodium silicate solution and various metal salts.

Digital document: doc-30870
Teacher-led video: tlvd-0616

on Resources

Digital documents Practical investigation logbook (doc-30925)
Experiment 3.1 Metallic trees (doc-30869)
Experiment 3.3 The effect of heat on metals (doc-30871)

3.7 Exercises

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3.7 Exercise 1: Multiple choice questions

1. A characteristic of metals is that their atoms
   - A. are smaller than those of non-metals.
   - B. readily form ionic bonds with non-metal atoms.
   - C. have few electrons in their outer shells, and these are readily lost in chemical reactions.
   - D. have high electronegativity.
2. How can the structure and bonding in solid sodium metal at room temperature be described?
   - A. A network lattice of closely packed sodium atoms, held together by strong ionic bonds
   - B. A lattice of Na_2 molecules
   - C. A network lattice of sodium ions, held together by a 'sea' of electrons
   - D. A lattice of Na_2 molecules held together by weak bonds
3. Which of the following properties is *not* characteristic of metals?
   - A. They are good conductors of heat.
   - B. They are capable of being polished.
   - C. They are relatively dense solids at room temperature.
   - D. They are hard and brittle.
4. What is the property of metals that enables them to conduct an electric current?
   - A. The outer electrons of metal atoms are not firmly bound to the atom.
   - B. Metal atoms are better suited as charge carriers than non-metal atoms.
   - C. Ions in the metal can move freely through the metallic lattice.
   - D. Metal atoms are not as firmly bonded to each other as are non-metallic atoms.
5. The physical properties of solid metals can best be explained by proposing that
   - A. each metal atom is held in the crystal lattice by covalent bonds.
   - B. positive metal ions are arranged in an orderly way, with delocalised electrons able to move freely throughout the crystal lattice.
   - C. positive and negative metal ions are arranged in an orderly way, with delocalised electrons able to move freely through the crystal lattice.
   - D. each metal atom is surrounded by a variable number of valence electrons, which complete a 'noble gas' electronic structure in the crystal lattice.
6. Both copper and sodium chloride have high melting points. Copper, in its solid and liquid states, is a good conductor of electricity while sodium chloride conducts electricity only when molten or dissolved in water. How can this difference in electrical conductivity be explained?
   - A. Much stronger bonding forces exist between atoms in sodium chloride than between those in copper.
   - B. Electrons in solid copper have much greater mobility than those in solid NaCl.
   - C. Chlorine is a poor electrical conductor.
   - D. Copper has only one kind of atom.
7. What is the structure of a solid formed by heating a mixture of two metallic elements at high temperature likely to be?
   - A. A metallic solid called an alloy
   - B. An ionic solid
   - C. A mixed ionic lattice and metallic lattice
   - D. Separate molecules of each element

8. Which of the following is *not* a possible mechanism for the formation of an alloy of two metals?
   - A. Atoms of one metal fit into the interstitial spaces between atoms of the other.
   - B. A new element is formed.
   - C. Atoms of one metal replace atoms of the other in the crystal lattice.
   - D. Separate crystals of one metal are dispersed throughout the other.
9. Which element can be found in nature in elemental (uncombined) form?
   - A. K
   - B. Ca
   - C. Au
   - D. Al
10. Why does aluminium, a fairly reactive metal, not react with oxygen gas in air?
    - A. It is protected by small amounts of its own impurities.
    - B. Its oxide forms a hard protective layer on the metal.
    - C. Nitrogen and carbon dioxide gas in the air prevent oxidation.
    - D. It forms a hard protective layer by a reaction with sulfur in the air.
11. Why do iron rubbish bins coated with a complete layer of zinc not rust?
    - A. The zinc acts as a sacrificial metal, reacting to prevent the iron rusting.
    - B. Iron is a more reactive metal than zinc.
    - C. The zinc combines with the iron to form a new compound that does not rust.
    - D. Particles of rust are unable to stick to the zinc surface.
12. Consider the following:
    - Metal *G* slowly reacts with cold water.
    - Metal *H* is generally shiny.
    - Metal *E* readily reacts with oxygen.
    - Metal *F* produces hydrogen when it reacts with acid.

    What is their reactivity order, from the most to the least reactive?
    - A. $H > G > F > E$
    - B. $H > E > F >$ G
    - C. $F > H > G > E$
    - D. $E > G > F > H$
13. Which is formed by the reaction of zinc with hydrochloric acid?
    - A. Zinc hydroxide + hydrogen
    - B. Zinc sulfate + chlorine
    - C. Zinc chloride + hydrogen
    - D. Zinc chloride + water
14. When a piece of sodium metal is carefully added to water, it reacts vigorously. What would a piece of caesium metal do?
    - A. React with about the same vigour
    - B. React more vigorously
    - C. React less vigorously
    - D. Not react at all with the water
15. Which of the following is a metal that reacts with dilute acids but not cold water?
    - A. Potassium
    - B. Gold
    - C. Iron
    - D. Carbon

3.7 Exercise 2: Short answer questions

1. The electron configurations of elements *A, B*, and *C* are as follows.
   - *A*: 2, 6
   - *B*: 2, 8, 1
   - *C*: 2, 8, 5

   Which of these elements is a metal?
2. Which of the following substances contain metallic bonding?
   - a. Lithium
   - b. Sulfur
   - c. Sodium bromide
   - d. Mercury(II) fluoride
   - e. Calcium
   - f. Argon
3. Woks used in cooking can be made from stainless steel or normal steel. It is believed that normal steel enhances the flavour of the food. The wok is washed without an abrasive, dried, and then stored with a thin coating of oil on its inner surfaces.

   Regular saucepans, on the other hand, are usually made of the much more expensive stainless steel and often have a copper base.
   - a. Discuss the properties of these metals with reference to their use in the kitchen.
   - b. Why do people coat the woks with oil?
   - c. What other metals are used in the kitchen?
4. a. Why was gold one of the first metals to be used?
   - b. Why was it not used for weapons?
5. Over the years, dentists have used metals for tooth fillings: first gold and then an amalgam of mercury. This amalgam is made by shaking a powdered mixture of silver, tin, copper and zinc with mercury. Within seconds, the mercury dissolves to form an amalgam that the dentist packs into the cavity. This alloy hardens in two hours.

   Match the property of each metal with those of the alloy used in a dental amalgam.

Metal	Property
Silver	Can be poured into the tooth easily
Tin	Does not tarnish or react with food or drink
Mercury	Bonds to the mercury and helps the amalgam set
Copper and zinc	Helps the amalgam stay free from tarnish
Gold	Gives the amalgam strength

6. a. What is the difference between a metal and an alloy?
   - b. Give examples of three alloys and indicate what advantages they have over the metal.
7. State, with reasons, whether each of the following substances is an element, a compound or an alloy.
   - a. Copper
   - b. Brass
   - c. Gold
   - d. 18 carat gold
   - e. Silver bracelet
   - f. Iron
   - g. Bronze
   - h. Copper oxide
   - i. Rust
   - j. Oxygen
   - k. Solder
   - l. Stainless steel
   - m. Water
8. Although mercury and mercuric compounds are toxic, mercury can safely be used in a tooth amalgam. What does this suggest about mercury and the amalgam?

9. Identify from which block (*s*, *p* or *d*) each of the following metals is found.
   a. It corrodes easily, forming a orange–brown solid compound.
   b. It has a partially filled *p* subshell and low density.
   c. It has a low melting point and can be cut easily with a knife.
   d. It is shiny and grey and, in nano form, has good antibacterial properties.
   e. It forms a compound that is white and has a subshell electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$.
10. Explain the difference between tempering and annealing a metal.
11. Explain, in terms of their bonding, why quenched metals are harder than unquenched metals.
12. Describe a situation where each of the following metal coatings may be used.
   a. Plastic
   b. Oil
   c. Noble
   d. Sacrificial
13. Describe an experiment that could be used to find the position of nickel in the activity series of metals.
14. Place a tick in the following table if the substance has the properties listed.

	Nanogold	Bulk gold	Both	Neither
Reactive				
Magnetic				
Spreads through water				
Range of colours				
Good conductor				
Very large surface area				
Insoluble				
Unreactive				
Does not conduct				
Transition metal				

15. Iron is an important metal for industry and consumers. It is used in greater quantity and variety than any other metal.
   a. List four important uses of iron.
   b. List four properties of iron that are relevant to the uses in part (a).
   c. Why is iron used so extensively?
   d. Why is iron usually alloyed with other metals?
   e. Outline the reactions in the blast furnace involving carbon.

3.7 Exercise 3: Exam practice questions

Question 1 (3 marks)

The solders *A*, *B* and *C* have the following properties.

A is a liquid above 180 °C

B melts gradually from 180 °C to 230 °C

C has a very high melting point

Which solder would you use for each of the following jobs?

a. Mending car radiators **1 mark**

b. Joining wires in an electric circuit **1 mark**

c. Filling gaps in car bodies **1 mark**

Question 2 (6 marks)

Iron is an important metal for industry and consumers. It is used in greater quantity and variety than any other metal. This is for several reasons:

- Iron is malleable and ductile when hot, so that means that different articles may be made from it.
- Iron forms a variety of alloys that may have vastly different properties.
- Although pure iron is relatively soft, when it is alloyed with chromium, its hardness and strength increase.
- Iron is the cheapest metal.

a. Suggest why iron is the cheapest of all metals. **2 marks**

b. What is the name of the alloy of iron and chromium, and what is another property of this alloy? **2 marks**

c. List four important uses of iron. **2 marks**

3.7 Exercise 4: studyON Topic Test online only

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4 Ionic compounds

4.1 Overview

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4.1.1 Introduction

To construct new substances such as smart materials, self-repairing materials and nanotubes, scientists and engineers need to understand the way atoms are put together and the limitations of the bonding models. The way that the atoms in a material are bonded together directly affects its properties and performance. This topic investigates the bonding between metals and non-metals, known as ionic bonding.

FIGURE 4.1 Giant crystals of gypsum, $CaSO_4{\cdot}2H_2O$, up to 11 metres long, found in the 'Cave of crystals' in Mexico. Note the size of the crystals compared with the human figure.

4.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- common properties of ionic compounds (brittleness, hardness, high melting point, difference in electrical conductivity in solid and liquid states), with reference to their formation, nature of ionic bonding and crystal structure including limitations of representations
- experimental determination of the factors affecting crystal formation of ionic compounds
- the uses of common ionic compounds.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of learning and assessment. Experiments and investigations, supported by a **Practical investigation logbook** and **Teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

on Resources

Digital document Key science skills (doc-30903)

Key terms glossary — Topic 4 (doc-30935)

Practical investigation logbook — Topic 4 (doc-30936)

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4.2 Structure and properties of ionic substances

KEY CONCEPT

- Common properties of ionic compounds (brittleness, hardness, high melting point, difference in electrical conductivity in solid and liquid states), with reference to their formation, nature of ionic bonding and crystal structure including limitations of representations

4.2.1 Chemical bonds

Since ancient times, humans have gathered attractive crystals to use as household decorations and jewellery. Many of these crystals consist of **ionic compounds** in which metal and non-metal atoms are joined by an **ionic bond**.

We are aware that all the varied substances around us are formed from atoms combining together. But this still leaves some questions:

- What holds atoms together once they have combined?
- Why do some atoms join or bond with others, while some remain as individual atoms?
- Why do atoms combine in specific ratios?

The answers lie in the nature of the **chemical bond**. All forces of attraction leading to chemical bonding between atoms are electrostatic in nature; that is, an attraction between protons (+) and electrons (−).

Most spontaneous changes that take place in nature occur in order to reach a more stable state. We seldom find free atoms in nature because atoms undergo changes in structure to become more stable. They do this by joining together, or bonding. Sometimes they bond with atoms of the same kind (for example, the covalent diatomic molecule of hydrogen gas, H_2) and sometimes with atoms of a different kind (for example, the ionic crystal lattice of sodium chloride, NaCl). Only the noble gases may exist as free atoms (monoatomic). This behaviour can be explained by an atom's electron configuration.

FIGURE 4.2 The mineral fluorite contains the ionic compound calcium fluoride, CaF_2, as its main component. Its beautiful crystals are often cubic in shape and appear in a variety of colours.

Atoms of other elements than group 18 (the noble gases) become more stable by gaining or losing enough electrons to achieve a complete outer shell configuration, like a noble gas. This does not mean that they have become noble gases, since they have not changed their nuclear structure (number of protons).

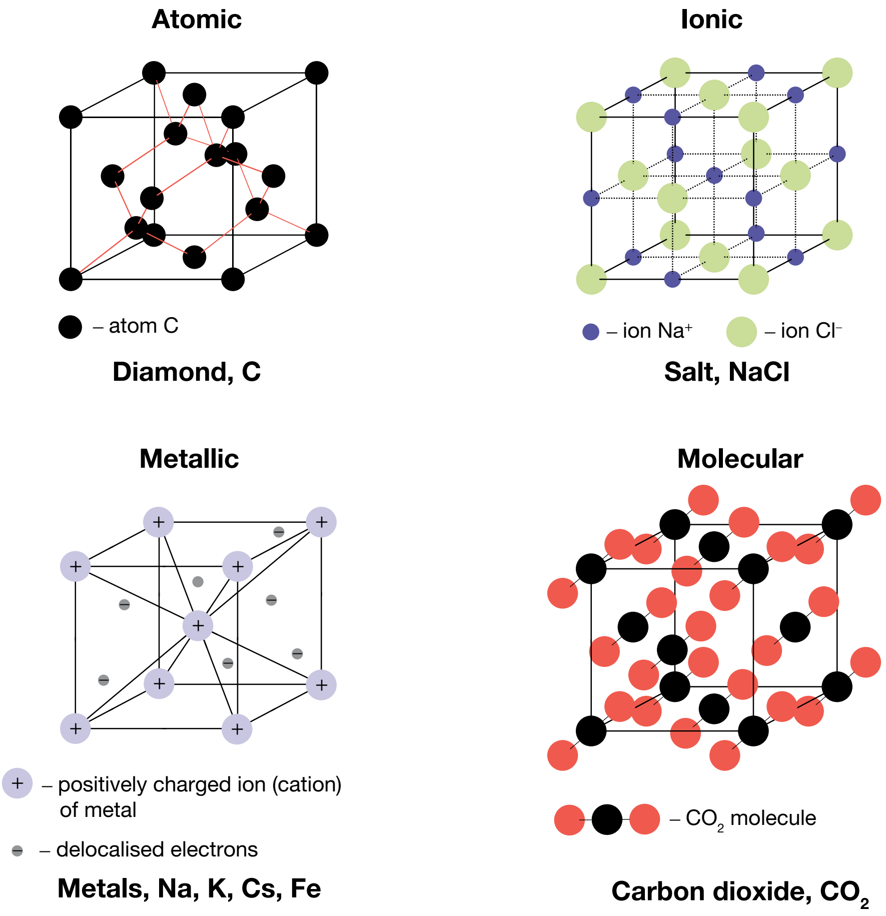
Atoms can become stable in one of three ways:

- by donating electrons to another atom
- by accepting electrons from another atom
- by sharing electrons with another atom.

When atoms combine to achieve more stable structures, three types of bonding are possible (figure 4.3).

- An ionic bond results when metallic atoms combine with non-metallic atoms to form an ionic lattice.
- A **metallic bond** results when metallic atoms combine to form a metallic lattice.
- A **covalent bond** results when non-metallic atoms combine to form either molecules or covalent lattices.

FIGURE 4.3 Many different types of crystal lattice types exist.



4.2.2 From atoms to ions

Atoms that gain or lose electrons to achieve more stable outer shell configurations are called **ions**. When an atom becomes an ion, it is no longer neutrally charged, since the number of electrons is not equal to the number of protons (see table 4.1). Note that the number of protons remains the same. An atom that has lost electrons becomes *positively* charged and is called a **cation** (e.g. Na^+). An atom that has gained electrons becomes *negatively* charged and is called an **anion** (e.g. Cl^-).

FIGURE 4.4 All elements in the periodic table are neutral and some can become charged when they are involved in a chemical bond.

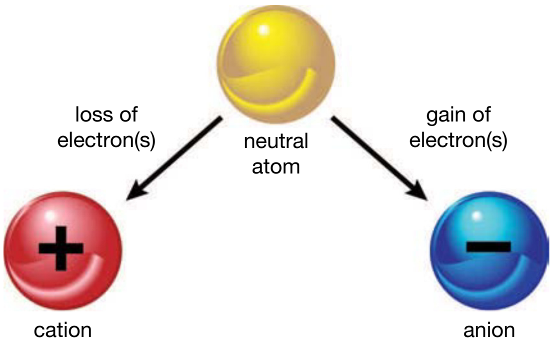


TABLE 4.1 Common atoms and their ions

Atom/ion	Symbol	Number of protons	Number of electrons
Sodium atom	Na	11	11
Sodium cation	Na^+	11	10
Chlorine atom	Cl	17	17
Chloride anion	Cl^-	17	18

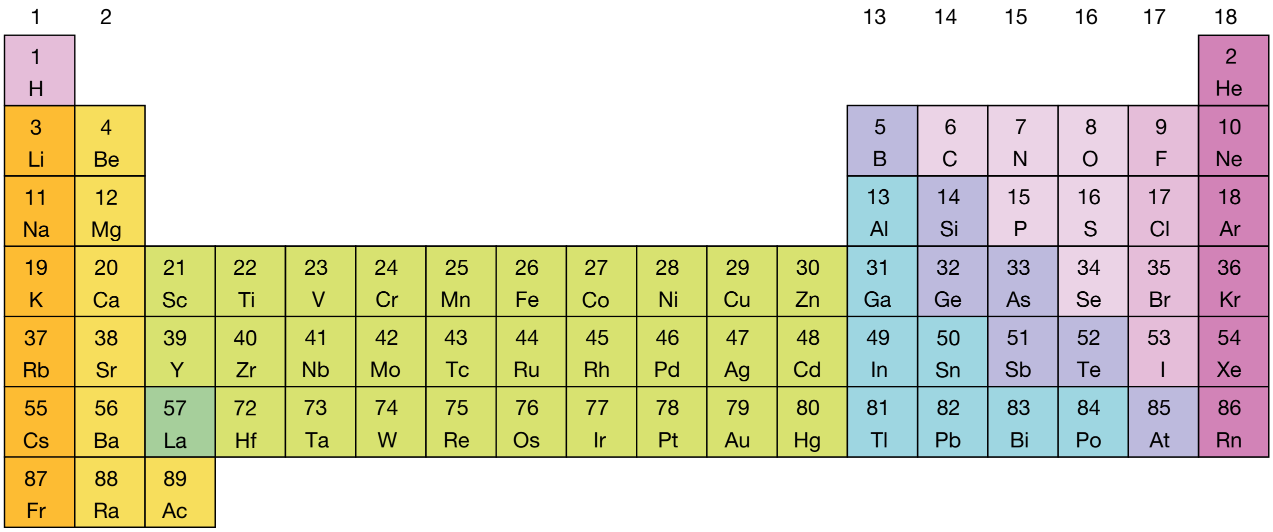
The electrons in the outer shell of an atom are sometimes called the valence electrons. The number of outer shell electrons is related to the group number in the periodic table.

When an atom gains or loses electrons, an ion is formed; in other words, the atom becomes charged.

Metallic ion formation

The metallic elements are those on the left side of the staircase in figure 4.5 (blue cells). These elements generally have low **electronegativities**. They can lose electrons to *achieve a noble gas configuration in their outer shells*.

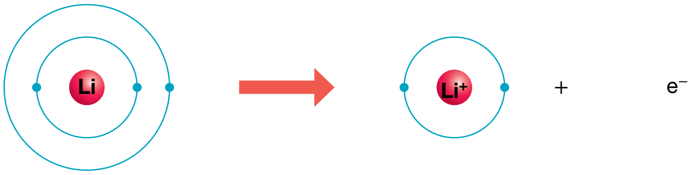
FIGURE 4.5 Periodic table (up to element 89) showing the division between metals and non-metals



1	2											13	14	15	16	17	18
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac															

For example, lithium is a very reactive group 1 metal with one outer shell electron and the electron configuration 2, 1. In order to obtain the stable configuration of a full outer shell, the lone electron is lost (figure 4.6). The electron configuration, 2, of the nearest noble gas, helium, results. Since the lithium cation has three protons but only two electrons, it has a net charge of 1+. Charges are written as superscripts above and to the right of the element symbol; thus the lithium atom is now written as Li^+. This process can be represented by electron shell diagrams or in the simple equation form in figure 4.5.

FIGURE 4.6 The lithium atom has one valence electron, which it loses to form the lithium cation, which has a charge of 1+.



lithium atom	→	lithium ion	+	one electron
Li	→	Li^+	+	e^-
2,1	→	2		

Note: when an atom's net charge is 1, it is not necessary to include the numeral 1 in the superscript notation.

The group 2 and group 13 metals contain two and three valence electrons respectively. They lose their outer shell electrons to form ions with charges of 2+ and 3+ respectively.

The electron configurations for the simple ions formed by the metals in periods 2, 3 and 4 are identical to those of the closest noble gases. Each occupied energy shell contains the maximum number of electrons. Examples of their electron configurations are as follows:

Li^+	2
Na^+	2, 8
K^+	2, 8, 8.

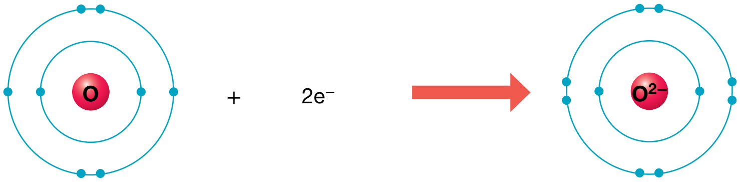
When we name a metallic ion, we use the full name of the metal followed by the word 'ion' to distinguish it from the uncharged metal. *Note:* the group 14 elements, carbon and silicon, do not form simple ions.

Non-metallic ion formation

Non-metallic elements are shown on the right side in the purple section in the periodic table presented in figure 4.5. These elements generally have high electronegativities. They gain electrons to achieve a noble gas configuration of eight electrons in their outer shells (with the exception of hydrogen).

For example, oxygen in group 16 has six outer shell electrons and has the electron configuration 2, 6. It is too difficult to remove all six electrons to achieve a full outer shell, so the oxygen atom gains two electrons instead to become a stable anion, as shown in figure 4.7 and the simple equation.

FIGURE 4.7 The oxygen atom has six valence electrons, and gains two electrons to form the oxide ion, which has a charge of 2–.



oxygen atom	+	two electrons	→	oxide ion
O	+	$2e^-$	→	O^{2-}
2, 6	+	$2e^-$	→	2, 8

An anion has more electrons than a neutral atom of the same element; here, the oxygen ion has eight protons and ten electrons, resulting in a net charge of 2–. The electron configuration of the oxygen anion is now that of a neon atom: 2, 8. The oxygen atom has become an oxide ion. (It is a convention in chemistry to indicate the ions of non-metallic elements with the suffix *-ide.*)

SAMPLE PROBLEM 1

Write the symbol, charge and name of the ions you would expect atoms of the following elements to form:

a. **Mg**

b. **S.**

Teacher-led video: SP1 (tlvd-0515)

THINK	WRITE
a. Mg is a metal found in group 2. It has the simplified electron configuration 2, 8, 2. In order to become stable, the Mg atom needs to lose two electrons. The charge of the resultant ion would, therefore, be 2+.	Mg^{2+} magnesium ion
b. S is a non-metal found in group 16. It has the simplified electron configuration 2, 8, 6. In order to become stable, the S atom needs two more electrons to complete its outer shell. The charge of the resultant ion would, therefore, be 2–.	S^{2-} sulfide ion

PRACTICE PROBLEM 1

Write the symbol, charge and name of the ions you would expect atoms of the following elements to form:

a. **Ga**

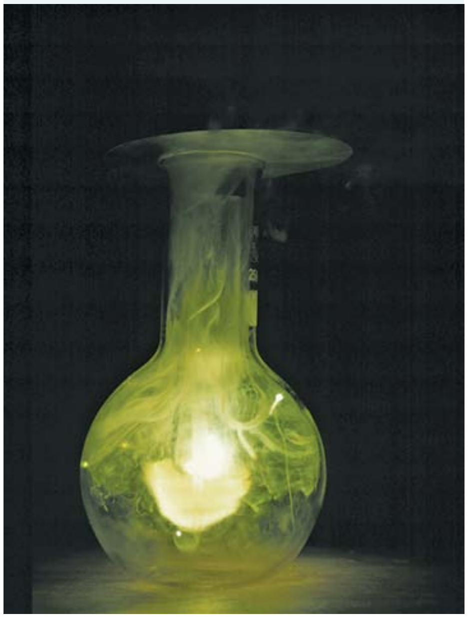
b. **P.**

4.2.3 Ionic bonding

Metals and non-metals can react together because the electrons lost by the metals can be taken up by the more electronegative non-metals. This transfer of electrons results in an ionic bond between the metal cations and the non-metal anions.

For example, when a small piece of sodium metal is added to a flask containing chlorine gas, a chemical reaction occurs, as shown in figure 4.8. The sodium ignites, and the solid ionic compound sodium chloride forms. This involves a transfer of one electron from the sodium atom to the chlorine atom. The ways in which we can represent this process are shown in the solution to sample problem 2.

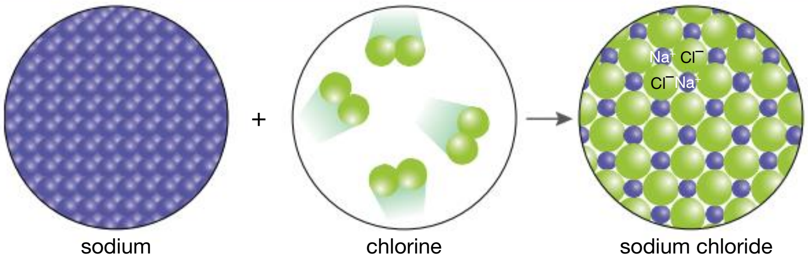
FIGURE 4.8 Molten sodium metal reacts violently with chlorine gas to form sodium chloride crystals.



Sodium ions and chloride ions are oppositely charged and, therefore, they attract each other by **electrostatic attraction**. This strong force of attraction between positive and negative ions is what produces an ionic bond.

Note that the bonded form is more stable than the unbonded form and has different properties from those of the original elements. Sodium is a very reactive metal; chlorine is a poisonous gas. Sodium chloride, however, is a non-toxic ionic compound familiar to us as table salt. The term ‘salt’, however, has a wider meaning in chemistry. The general name for an ionic compound, a substance formed from the reaction between a metal and a non-metal, is a **salt**.

FIGURE 4.9 The formation of sodium chloride from solid sodium and gaseous chlorine



on Resources

Interactivity Ionic models (int-6351)

Weblink Reaction of sodium with chlorine

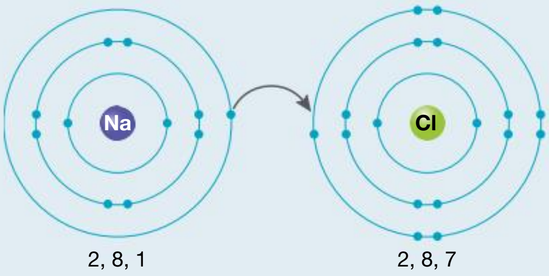
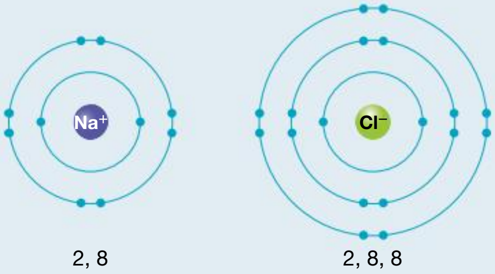
SAMPLE PROBLEM 2

Consider the reaction between sodium and chlorine and explain it using:

a. electron shell diagrams

b. a simple equation.

Teacher-led video: SP2 (tlvd-0516)

THINK	WRITE
a. Sodium and chlorine can be represented as follows:  When a sodium atom forms an ionic bond with a chlorine atom, the sodium atom loses one electron (its entire outer shell) to become a positive ion, or cation. The chlorine atom gains one electron to complete its outer shell. It is now a negative ion, or anion.	
b. The sodium and chloride ions form the substance called sodium chloride.	$Na^+ + Cl^- \rightarrow NaCl$

PRACTICE PROBLEM 2

Consider the following ionic compounds and explain the bonds between their atoms using electron shell diagrams and simple equations:

a. magnesium oxide

b. magnesium fluoride.

TIP: **It is sometimes useful to use dots and crosses when drawing shell diagrams because this makes it easier to see the origin of the transferred electron(s).**

4.2.4 Structure and properties of ionic substances

Like most ionic substances, sodium chloride is a hard, brittle crystalline solid at room temperature. It has a high melting point (800 °C). This indicates that its particles are arranged in repeating three-dimensional patterns and are strongly held together. When sodium chloride is dissolved in water or is melted, its crystal structure breaks down and in both these states it is able to conduct electricity, indicating that charged particles (ions) are free to move.

How can we explain the distinctive composition, melting point and conductivity characteristics of ionic substances? What is it about the structure of these substances that makes them behave so uniquely?

Ionic lattices

The major features of the current theory of the structure of ionic substances are as follows.

- Ionic compounds form crystals that are composed of three-dimensional arrays of positive metal ions and negative non-metal ions.
- These arrays are called ionic network lattices and are held together by the strong electrostatic attraction — the ionic bond — between the oppositely charged ions.
- The ions are arranged in a regular repeating pattern throughout the crystal. They pack together in the way that will achieve the most stable arrangement, with oppositely charged ions as close together as possible, and similarly charged ions as far apart as possible. Each cation is thus surrounded by anions and each anion by cations.
- The relative numbers of cations and anions are fixed by the requirement that the solid is neutral.
- The relative sizes and numbers of the ions present determine the actual lattice structure.

Sodium chloride is a typical ionic compound. Its appearance and ionic structure are shown in figures 4.10 and 4.11. Its crystals are cubic, reflecting the regular arrangement of the ions in the lattice. Each sodium ion is surrounded by six chloride ions and each chloride ion is surrounded by six sodium ions. In this arrangement, each ion is strongly attracted to each of its neighbours. The large attractive forces result in a very stable structure.

FIGURE 4.10 Sodium chloride lattice

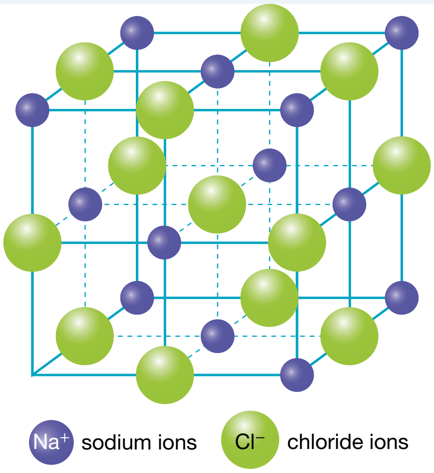


FIGURE 4.11 Each sodium ion is surrounded by six chloride ions, which gives the crystals of sodium chloride their characteristic cubic shape.

The formula of the ionic substance sodium chloride is NaCl, which means that the ratio of positive to negative ions is 1:1. The formula NaCl does not represent a distinct molecule of sodium chloride but simply means that a sample of sodium chloride contains equal numbers of sodium ions and chloride ions. The formula $MgBr_2$ signifies that a sample of magnesium bromide contains twice as many bromide ions as magnesium ions.

A chemical formula given for an ionic compound is called an empirical formula. This refers to the simplest ratio of ions in the lattice.

on Resources

Digital document	Experiment 4.1 Investigating calcite crystals (doc-30867)
Teacher-led video	Experiment 4.1 Investigating calcite crystals (tlvd-0618)
Interactivity	Pass the salt (int-0675)
Weblink	Animation of NaCl structure

4.2.5 Connecting properties of ionic compounds to structure

Ionic compounds are composed of two or more different kinds of ions that have opposite charges. The electrostatic forces of attraction holding them together are strong. As a result, ionic compounds:

- are usually crystalline solids, owing to the arrangement of ions in repeating three-dimensional patterns
- have high melting and boiling points, because a large amount of energy is needed to separate the ions
- do not conduct electricity in the solid form, because the charged particles (ions) are not able to move (to conduct electricity, a substance must contain mobile charged particles)
- are hard, since the surface of the crystal is not easily scratched due to the strong ionic bond holding the ions together
- often dissolve in water to form ions (dissociate), because water molecules are able to move between ions and free them by disrupting the rigid crystal structure
- conduct electricity when molten or in aqueous solution (dissolved in water). In the molten form, ions are able to slide past one another and can, therefore, conduct electricity. When an ionic substance is dissolved in water, the ions dissociate from the lattice and can move freely to conduct an electric current. The solution is called an **electrolyte** (figure 4.12).

FIGURE 4.12 Dissociation of sodium chloride to form a conducting solution called an electrolyte

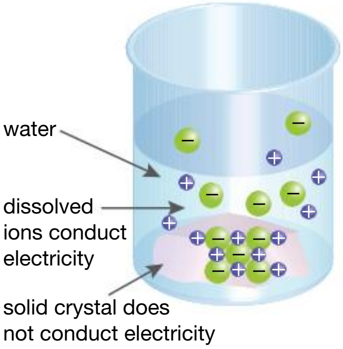
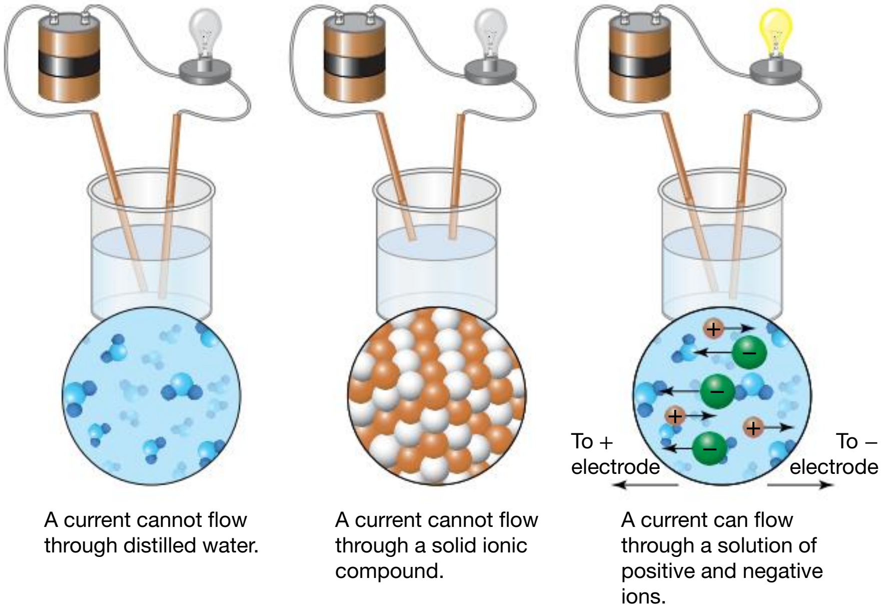


FIGURE 4.13 Electrical properties of distilled water, solid ionic compounds and ionic solutions



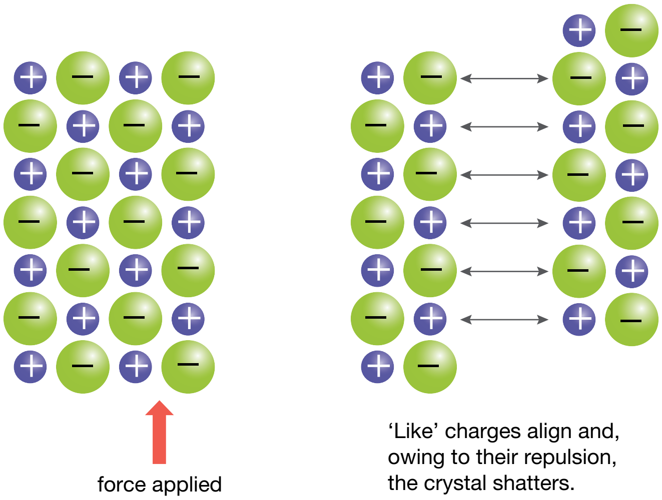
A current cannot flow through distilled water.

A current cannot flow through a solid ionic compound.

A current can flow through a solution of positive and negative ions.

- are brittle, since distortion of the crystal causes ions of like charge to come close together and the repulsion between these ions cleaves or shatters the crystal, as shown in figure 4.14. If the distortion of the crystal is caused by a sharp blow along the plane of ions, cleavage (a clean split) along the plane occurs. If the blow is along different planes of ions, the crystal lattice shatters.

FIGURE 4.14 The brittle nature of an ionic compound, which cleaves or shatters when distorted.



TIP: When explaining conductivity in ionic substances, remember that it is not the electrons that move. Only the ions are mobile in solution and when molten, not the electrons.

FIGURE 4.15 Old alkaline batteries often leak potassium hydroxide, which reacts with carbon dioxide in the air to form crystals of potassium carbonate.

on Resources

Digital document Experiment 4.2 Ionic models (doc-30868)

4.2.6 Limitations of ionic models

Electron dot diagrams

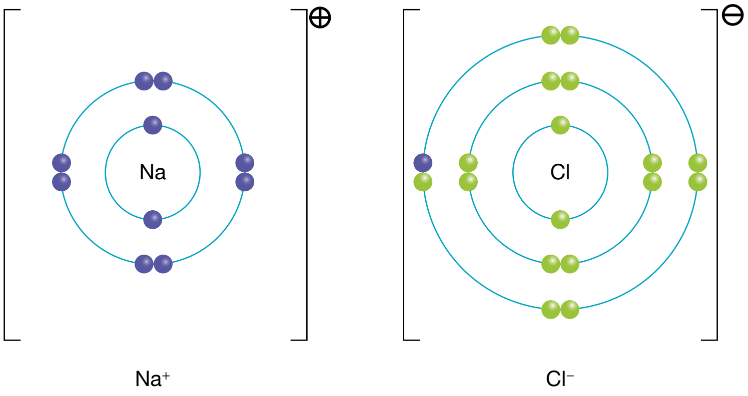
An electron dot diagram shows:

- how the ionic bonds are formed
- the ratio in which the atoms react.

However, it does not show how the ions are arranged in space.

The electron dot diagram for the sodium chloride ions shown in figure 4.16 suggests that it is made up of one pair of sodium and chloride ions. This is not the case. Sodium chloride has a giant ionic lattice structure with equal numbers of sodium and chloride ions. In a 1 g sample of pure sodium chloride, for example, will be roughly 5 × 10^{21} of each ion.

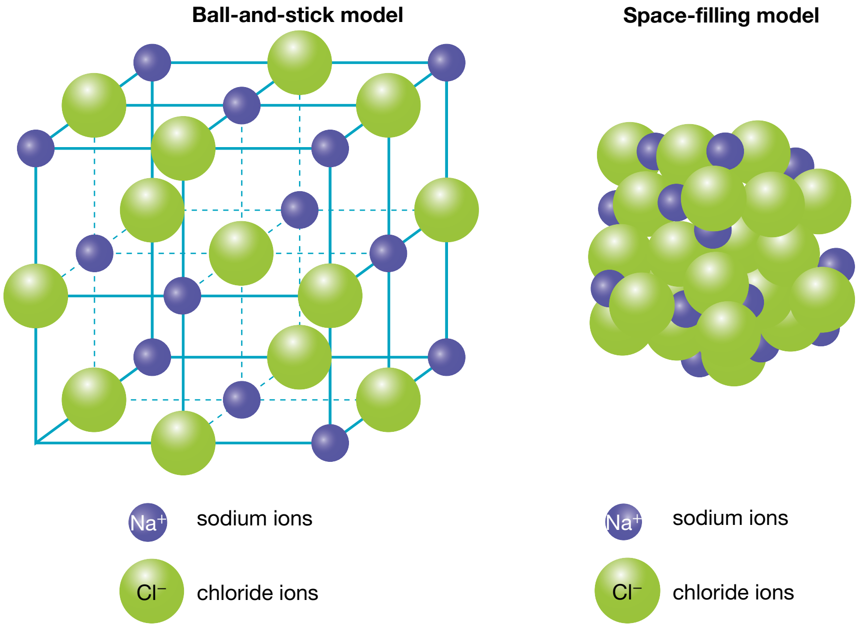
FIGURE 4.16 In electron dot diagrams, brackets are often used to indicate that the charge is spread over the whole ion.



Three-dimensional models of ionic compounds

A three-dimensional (3D) model shows how the ions are arranged in a lattice structure. These models usually have coloured balls to represent the ions and use sticks to show the ionic bonds (the ball-and-stick model), or they can be space-filling models (see figure 4.17).

FIGURE 4.17 Three dimensional representations of NaCl using the ball-and-stick model and space-filling model.



A 3D model still has limitations:

- it is not to scale
- it gives no information about the forces of attraction between the ions
- the movement of electrons to form the ions cannot be shown
- it does not show the relative sizes of the ions.

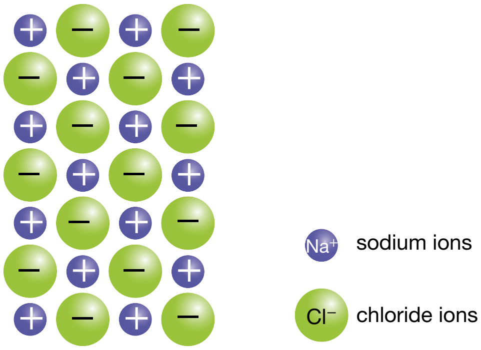
Two-dimensional models of ionic compounds

A two-dimensional (2D) representation is similar to the three-dimensional ball-and-stick model, but is an easier model to draw (see figure 4.18).

In addition to the 3D model limitations, the 2D model has further limitations:

- it only shows the arrangement of one layer of ions
- it does not show where the ions are located on the other layers or the relationship between them
- it does not show that many different arrangements of ions are possible.

FIGURE 4.18 A 2D model for the ionic lattice in sodium chloride



4.2 EXERCISE

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1. Explain why atoms such as Ca and Al form ions.
2. Use electron shell diagrams and simple equations to show how the following atoms form their corresponding ions.
   (a) The metal atoms Ca and Al
   (b) The non-metal atoms N and F
3. For each of the following atoms, predict the charge of the ion formed, and write its name and symbol.
   (a) Ba
   (b) K
   (c) P
   (d) Cl
   (e) S
4. Complete the following table, predicting the general electrovalencies for period 2 and period 3 ions in each group.

Period	Group 1	Group 2	Group 13	Group 15	Group 16	Group 17
2	1+					
3						

5. Draw electron shell structures and write simple equations to show the ionic bonding between the following.
   (a) Calcium and oxygen
   (b) Beryllium and chlorine
   (c) Lithium and fluorine
   (d) Aluminium and sulfur
   (e) Sodium and nitrogen
   (f) Magnesium and sulfur
6. Write the electron configurations of the following.
   (a) The cations potassium, calcium and aluminium
   (b) The anions fluoride, oxide and nitride
7. Name the noble gas that has the same electron configuration as the anions in fluoride, oxide and nitride.
8. Magnesium chloride, $MgCl_2$, and potassium chloride, KCl, are typical ionic compounds.
   (a) Describe how their ionic lattice structure is formed from their respective atoms.
   (b) Show why their empirical formulas have ratios of 1:2 and 1:1 respectively.
   (c) Explain what holds the ions close together in their lattice structures.
9. Calcium chloride is a crystalline substance at room temperature. Predict whether it
   (a) has a low or high melting point.
   (b) shatters when pressure is applied.
   (c) conducts electricity in the solid or liquid state.
   Justify your predictions.

10. (a) Compare the ratio of positive to negative ions in the lattice of the following.
    i. Sodium chloride
    ii. Magnesium oxide
    (b) Which of these compounds would you expect to have the higher melting point? Give a reason for your choice.
11. A solid substance has a high melting point, conducts electricity only in molten form and cleaves when struck with a sharp knife. However, it does not readily dissolve in water. Is this enough data to predict the chemical bonding in this substance? Explain in terms of the limitations of the bonding model.
12. Atom *X*, having one electron in its outer shell, combines with atom *Y*, which has six electrons in its outer shell.
    (a) Write the formula of the compound formed.
    (b) Name the type of bonding in the compound.
    (c) Predict three general properties of the compound.

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4.3 Ionic nomenclature

Background knowledge

- Naming conventions of ionic compounds
- Formulas for binary ionic compounds, ions with variable charges and polyatomic compounds

4.3.1 Naming ionic compounds

Knowing how to write formulas is a skill that all chemists require. Many known compounds exist and to memorise them all would be impossible. Chemists have developed a set of rules that allow us to predict the formulas of many compounds. Some compounds that have been known for many years also have common names, as listed in table 4.2.

TABLE 4.2 Common ionic compounds

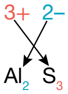
Common name	Systematic name	Formula	Used to manufacture
Salt	Sodium chloride	$NaCl$	Chlorine
Soda ash	Sodium carbonate	Na_2CO_3	Glass
Baking soda	Sodium hydrogen carbonate	$NaHCO_3$	Cake raising agent
Lime	Calcium oxide	CaO	Mortar
Limestone	Calcium carbonate	$CaCO_3$	Cement
Potash	Potassium nitrate	KNO_3	Gunpowder
Milk of magnesia	Magnesium hydroxide	$Mg(OH)_2$	Laxative or antacid
Gypsum	Calcium sulfate	$CaSO_4$	Plaster
Caustic soda	Sodium hydroxide	$NaOH$	Soap or drain cleaner

4.3.2 Formulas of binary ionic compounds

In a **binary** ionic compound, only two elements are present. Examples include sodium chloride, NaCl, calcium fluoride, CaF_2, and potassium nitride, K_3N. Note that the word 'binary' does not refer to the relative number of ions or subscripts. When writing empirical formulas for binary ionic compounds, the following rules should be followed:

- Write the symbol for the cation first, followed by the symbol for the anion.
- Determine the lowest whole number ratio of ions that provides a net charge of zero.

The 'swap and drop' method is another way to determine formulas, using the charges of the ions. For example, to obtain the formula for aluminium sulfide, the symbol for the cation is written first, followed by the symbol for the anion: $Al^{3+}S^{2-}$. The charges are then swapped and dropped so that the charge of the anion becomes the subscript of the cation and the charge of the cation becomes the subscript of the anion.



The formula is then written showing only the subscripts. Note that subscripts must always be reduced to the smallest possible whole numbers. (For example, the formula of magnesium oxide is MgO, not Mg_2O_2.)

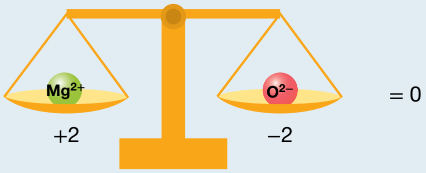
FIGURE 4.19 Sodium hydrogen carbonate, commonly known as baking soda, thermally decomposes at 80 °C to give off carbon dioxide gas and water vapour, which act as raising agents.

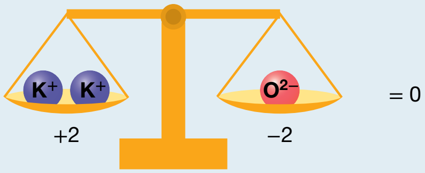
SAMPLE PROBLEM 3

Write formulas for the following compounds:

a. magnesium oxide
b. potassium oxide
c. aluminium sulfide.

Teacher-led video: SP3 (tlvd-0517)

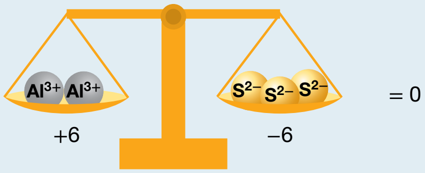
THINK	WRITE
a. The symbols for the cation and anion in the compound are Mg^{2+} and O^{2-} respectively. In order to have a net charge of zero, the number of positive charges must be balanced by an equal number of negative charges. Thus one Mg^{2+} ion is balanced by one O^{2-} ion and the ratio of ions is 1:1.  	MgO
b. The symbols for the cation and anion in the compound are K^+ and O^{2-} respectively. In order to have a net charge of zero, two positively charged potassium ions are needed to balance the two negative charges of the oxide ion. The ratio of ions is 2:1.	K_2O



The formula is written as K_2O. The subscript numeral '2' indicates two potassium ions are present for every oxide ion.

c. The symbols for the cation and anion in the compound are Al^{3+} and S^{2-} respectively. In order to have a net charge of zero, two Al^{3+} ions, with a total charge of 6+, are needed to balance three S^{2-} ions, with a total charge of 6–. The ratio of ions is 2:3.

Al_2S_3



PRACTICE PROBLEM 3

Name the compound MgF_2.

on Resources

Weblink Balanced ionic compounds quiz

4.3.3 Ions with a variable charge

As the atomic number increases, atoms become larger and have more electrons. Some transition metals have quite complicated arrangements of electrons and may be able to form more than one type of ion. Such ions have different charges. This is very significant because ions of the same element with different charges have different physical and chemical properties. For example, a solution of Cr^{6+} is orange, whereas a solution of Cr^{3+} is yellow (see figure 4.20).

FIGURE 4.20 When chromium reacts with other chemicals, often Cr^{3+} (yellow flask) changes to Cr^{6+} (orange flask). The colour change can indicate that a reaction has occurred.

Elements that have ions of variable charge can form two or more different binary ionic compounds containing the same element. Iron, for example, can form black FeO and brown Fe_2O_3, each with different physical and chemical properties. This means we must name them differently, since simply saying iron oxide is ambiguous. Therefore, we use roman numerals in brackets after the cation's name to denote its charge. For example, the black FeO contains Fe^{2+} ions, and so it is called iron(II) oxide. The brown Fe_2O_3 contains Fe^{3+} ions and is named iron(III) oxide.

Metal ions that form more than one cation include iron(II) and (III), copper(I) and (II), and mercury(I) and (II). Note that the roman numerals do not indicate the numbers of ions present in the compound, only the charge of the ion.

SAMPLE PROBLEM 4

a. **Name the compound CuO.**
b. **Give the formula for iron(II) chloride.**

Teacher-led video: SP4 (tlvd-0518)

THINK	**WRITE**
a. In order to determine the charge on the metal ion, we need to work backwards from the known charge on the anion. The charge on the oxide ion is 2–. Therefore, for a net charge of zero, the charge on the Cu ion must be 2+. The compound is, therefore, named copper(II) oxide.	copper(II) oxide
b. Iron(II) means that the Fe^{2+} ion is present in the compound. The chloride ion has a charge of 1–. For a net charge of zero, the 2+ of the Fe ion must be balanced with two chloride ions.	$FeCl_2$

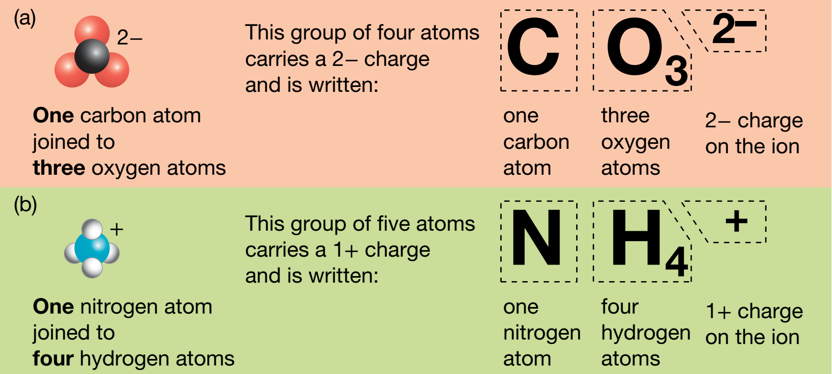
PRACTICE PROBLEM 4

a. **Name the compound of SrI_2.**
b. **Give the formula for vanadium(V) oxide.**

4.3.4 Polyatomic ions

A **polyatomic ion** is a group of tightly bound atoms that behaves as a single unit and carries an overall charge. It may be a positively or negatively charged ion. The carbonate ion, for example, is composed of one carbon atom and three oxygen atoms. The *whole group* of four atoms carries a negative two charge.

FIGURE 4.21 Two polyatomic ions: (a) the anion carbonate (CO_3^{2-}) and (b) the cation ammonium (NH_4^+).



Ionic compounds consist of arrays of cations and anions. A table of charges may be used to help write the empirical formula of an ionic compound (see table 4.3).

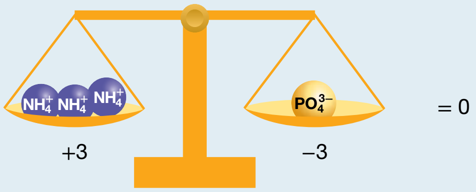
TABLE 4.3 Anions and cations

Cations					
1^+		**2^+**		**3^+**	
Lithium	Li^+	Magnesium	Mg^{2+}	Aluminium	Al^{3+}
Sodium	Na^+	Calcium	Ca^{2+}	Chromium(III)	Cr^{3+}
Potassium	K^+	Barium	Ba^{2+}	Iron(III)	Fe^{3+}
Caesium	Cs^+	Iron(II)	Fe^{2+}		
Silver	Ag^+	Nickel	Ni^{2+}		
Copper(I)	Cu^+	Chromium(II)	Cr^{2+}		
Ammonium	NH_4^+	Copper(II)	Cu^{2+}		
		Zinc	Zn^{2+}		
		Tin(II)	Sn^{2+}		
		Lead(II)	Pb^{2+}		
		Manganese(II)	Mn^{2+}		
		Mercury(II)	Hg^{2+}		
		Strontium	Sr^{2+}		
Anions					
1^-		**2^-**		**3^-**	
Hydride	H^-	Oxide	O^{2-}	Nitride	N^{3-}
Fluoride	F^-	Sulfide	S^{2-}	Phosphide	P^{3-}
Chloride	Cl^-	Sulfate	SO_4^{2-}	Phosphate	PO_4^{3-}
Bromide	Br^-	Carbonate	CO_3^{2-}		
Iodide	I^-	Sulfite	SO_3^{2-}		
Hydroxide	OH^-	Dichromate	$Cr_2O_7^{2-}$		
Nitrate	NO_3^-	Chromate	CrO_4^{2-}		
Hydrogen carbonate	HCO_3^-	Thiosulfate	$S_2O_3^{2-}$		
Hydrogen sulfate	HSO_4^-	Hydrogen phosphate	HPO_4^{2-}		
Chlorate	ClO_3^-				
Hydrogen sulfite	HSO_3^-				
Nitrite	NO_2^-				
Permanganate	MnO_4^-				
Hypochlorite	OCl^-				
Dihydrogen phosphate	$H_2PO_4^-$				
Cyanide	CN^-				

SAMPLE PROBLEM 5

Write the formula for the compound ammonium phosphate.

Teacher-led video: SP5 (tlvd-0519)

THINK	WRITE
1. The method for writing the formula for this compound is the same as that for a binary ionic compound. Write the symbol for the cation followed by the symbol for the anion.	Cation: NH_4^+ Anion: PO_4^{3-}.
2. Since the net charge must be zero, we need three NH_4^+ ions for every PO_4^{3-} ion.  Place brackets around the entire ammonium ion so that the subscript ‘3’ applies to all of it. Note that brackets need to be used whenever more than a single polyatomic ion is needed to balance a formula. This is the only time they are used.	$(NH_4)_3PO_4$

PRACTICE PROBLEM 5

Write the formula for the compound sodium dichromate.

4.3 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Write the formula for the ionic compound formed between each of the following sets of ions.
   (a) Aluminium and chlorine
   (b) Barium and oxygen
   (c) Sodium and sulfur
   (d) Magnesium and phosphorus
2. (a) Complete the following table by writing the formulas for the compounds formed when each cation is bonded to each anion.

Ions	K^+	Ca^{2+}	Al^{3+}
F^-			
O^{2-}			
N^{3-}			

   (b) Name the nine compounds formed in part (a).
3. Name the following compounds.
   (a) KCl
   (b) Ag_2S
   (c) Na_3N

4. Name the following compounds.
   (a) FeS
   (b) Fe_2S_3
   (c) $CuCl_2$
5. Name the following compounds.
   (a) SnO
   (b) Cu_2O
   (c) $PbBr_2$
6. Give the formula for each of the following compounds.
   (a) Tin(IV) fluoride
   (b) Lead(II) sulfide
   (c) Mercury(II) oxide
   (d) Iron(III) nitride
   (e) Copper(I) sulfide
   (f) Tin(II) oxide
7. Refer to table 4.3. Write the formula and name of each of the polyatomic ions that contains sulfur.
8. Using the charges provided in table 4.3, complete the following table.

Name of ionic compound	Valency of cation	Valency of anion	Empirical formula
Silver chloride			
Potassium sulfide			
Magnesium oxide			
Aluminium bromide			
Iron(III) carbonate			
Barium phosphate			
Ammonium sulfate			

9. Give the formulas for the following compounds.
   (a) Sodium sulfite
   (b) Calcium nitrite
   (c) Copper(II) hydrogen carbonate
10. Name the compound formed when each of the following pairs of ions is bonded, and write its formula.
   (a) Al^{3+} and CO_3^{2-}
   (b) Sodium and nitrate
   (c) Hg^{2+} and PO_4^{3-}
   (d) Lead(II) and sulfate

4.4 Crystal formation and uses of ionic compounds

KEY CONCEPT
- Experimental determination of the factors affecting crystal formation of ionic compounds
- The uses of common ionic compounds

4.4.1 Factors affecting crystal formation

Ionic compounds form crystals of many interesting shapes depending on the arrangement of ions in the lattice, but the different sizes of crystals are caused by the diverse conditions in which they are formed.

The largest natural crystals found have been in Chihuahua, Mexico, and these are up to 11 metres long, while the smallest fit into the nanoscale range.

Crystals of a more manageable size can be prepared in a school laboratory from saturated solutions, which are solutions in which no more solid (solute) will dissolve. To form a saturated solution, the solution is heated while the solute is added; this assists the dissolving process. As the solution cools, the water (or other solvent) evaporates and small crystals appear.

To produce a larger crystal, a regularly shaped crystal can be selected and suspended by a thread in a saturated solution and the crystal will grow. This crystal is called the seed crystal, and it provides a nucleation site, or a point where crystallisation can begin. Dust particles or scratches on a glass container can also act as nucleation sites. The rate of evaporation affects the size of the crystal formed.

Recrystallisation processes are very useful to chemists for purifying substances, because impurities cannot form part of the growing crystal structure.

FIGURE 4.22 Vanadinite, $Pb_5(VO_4)_3Cl$, crystals are hexagonal in shape and are a source of the element vanadium.

4.4.2 Uses of common ionic compounds

Ionic compounds have a wide range of uses, as can be seen in table 4.2 (refer to section 4.3.1). Many metals are obtained from ionic compounds that have been extracted from different ores. Sodium hydrogen carbonate is used in cooking and also as an environmentally friendly cleaning agent. Ammonium nitrate is used to manufacture fertilisers as well as explosives. Compounds containing nitrate, nitrite and sulfite ions are used to preserve food. Our bodies can function only because of the presence of particular ions.

Why do athletes take sports drinks?

An electrolyte is a substance that dissociates to form ions. Athletes sometimes take electrolyte drinks after exercise to rehydrate, boost energy and replace important ions in the body that are lost during sweating. Adequate water consumption is essential but much discussion surrounds the need for these drink supplements if individuals have an adequate diet. Table 4.4 provides a list of some ions that are essential for optimum functioning of our bodies.

FIGURE 4.23 As we sweat, we also lose electrolytes (water soluble ionic compounds); these may be replaced by sports drinks.

Hydrated ionic compounds

Hydrated ionic compounds contain water molecules bonded within the crystal.

A number of ionic compounds, called **hydrates**, release water, which is part of their structure, when they decompose upon heating. When the formula of a hydrated compound is written, the number of water molecules is also included. For example, the formula for iron(II) sulfate octahydrate is written as $FeSO_4{\cdot}8H_2O$. This formula indicates that eight molecules of water are bonded within the ionic crystal for every one formula unit of $FeSO_4$. This is also commonly called water of crystallisation.

TABLE 4.4 Common ions needed in the body

Name	Symbol	Functions in the body
Sodium	Na^+	Regulates fluid balance Involved in transmission of nerve impulses Controls blood pressure by controlling blood volume Involved in acid–base balance of blood
Potassium	K^+	Controls the level of body fluids Involved in transmission of nerve impulses and muscle contraction, including heartbeat Important in reactions within cells
Calcium	Ca^{2+}	Needed for building teeth and bones Involved in blood clotting Involved in transmission of nerve impulses and muscle contraction
Magnesium	Mg^{2+}	Needed for bone formation Assists in energy production Involved in functioning of muscle and nerve tissue Component of enzymes
Chloride	Cl^-	Involved in acid–base balance of blood Assists in fluid balance in the body Involved in formation of hydrochloric acid in stomach

The following prefixes are used to indicate the number of water molecules in a hydrated compound.

mono- = 1	*hexa-* = 6
di- = 2	*hepta-* = 7
tri- = 3	*octa-* = 8
tetra- = 4	*nona-* = 9
penta- = 5	*deca-* = 10

SAMPLE PROBLEM 6

a. Name the compound $ZnCl_2{\cdot}4H_2O$.

b. Give the formula for the compound calcium sulfate dihydrate.

Teacher-led video: SP6 (tlvd-0520)

THINK	WRITE
a. Name the binary ionic compound in the usual way, and because four water molecules are in the compound, use the prefix *tetra-* followed by the word 'hydrate'. Because zinc only has one possible charge (2+), it is often seen without the (II).	Zinc(II) chloride tetrahydrate
b. 1. Identify the cation and anion and write their symbols.	Cation: Ca^{2+} Anion: SO_4^{2-}
2. Balance the charges so the net charge is zero, keeping the polyatomic sulfate ion intact. Since the ions are equally charged, the formula for calcium sulfate is $CaSO_4$.	$CaSO_4$
3. determine the number of water molecules using the prefix: *di-* means 'two',	$CaSO_4{\cdot}2H_2O$

PRACTICE PROBLEM 6

a. Name the compound $CoBr_2{\cdot}6H_2O$.

b. Give the formula for the compound tin(II) chloride dihydrate.

4.4 EXERCISE

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1. The following table shows the uses and formulas of some hydrated ionic compounds. Complete either the formula for the compound or its name. The first row is completed for you.

Name of hydrated ionic compound	Common name(s); use; description	Formula
Copper(II) sulfate pentahydrate	Blue vitriol; copper plating; blue solid	$CuSO_4{\cdot}5H_2O$
Sodium carbonate decahydrate	Washing soda, soda ash; water softener; white solid	(a)
(b)	Epsom salts; explosives, matches; white solid	$MgSO_4{\cdot}7H_2O$
Magnesium chloride hexahydrate	Disinfectants, parchment paper; white solid	(c)
(d)	Dyeing fabrics, tanning leather; white solid	$BaCl_2{\cdot}2H_2O$

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4.5 Review

4.5.1 Summary

Structure and properties of ionic substances

- Ions are formed when atoms lose or gain electrons to achieve the stable electron configuration of a noble gas.
  - Metals lose electrons to gain a full outer shell and become positively charged ions called cations.
  - Non-metals gain electrons to obtain a full outer shell and become negatively charged ions called anions.
  - Whether an atom gains or loses electrons depends on its electronegativity, or electron-attracting power.
- In the periodic table, the group numbers 1, 2, 13, 15, 16 and 17 have the charge 1+, 2+, 3+, 3–, 2– and 1– respectively. Group 14 and group 18 elements do not readily form ions.
- Three types of chemical bond are possible:
  - ionic bond
  - metallic bond
  - covalent bond.
- Ionic bonding is produced by the strong electrostatic attraction that results between ions when a metallic atom transfers electrons to a non-metallic atom. The ionic compound formed:
  - is known as a salt
  - is made up of an ionic network lattice of positive cations and negative anions

  - is neutrally charged; that is, the total positive charge of the cations must equal the total negative charge of the anions
  - can be represented by a formula unit, which shows the smallest whole number ratio of cation to anion, when charges are balanced. This is also called the empirical formula.
- The structure and bonding of ionic compounds has a direct influence on their properties. Ionic network lattices are rigid structures of cations bonded to anions. Ionic compounds:
  - do *not* conduct electricity in the solid state but become electrolytes when in aqueous or molten states, when their rigid structures have broken down sufficiently to allow free movement of their ions
  - are brittle because pressure causes like charges to align, resulting in the structure shattering
  - are often soluble in water, the molecules of which disrupt the lattice and allow free mobile ions to form
  - are usually crystalline solids due to the close-packed, three-dimensional lattice structure
  - have high melting and boiling points due to the strong force of attraction between the ions.

Ionic nomenclature

- The key rules in naming ionic compounds and writing their formulas are as follows:
  - When naming a binary ionic compound, always name the metal ion in full first and then add the non-metal ion with the ending *-ide* (e.g. lithium oxide).
  - To find the correct formula for an ionic compound, determine the lowest whole number ratio of ions that gives a net charge of zero.
  - When naming ionic compounds that contain metal ions with more than one charge, such as iron(II) and (III), copper(I) and (II), lead(II) and (IV), mercury(I) and (II), and tin(II) and (IV), use roman numerals to indicate which ion is present (CuS is copper(II) sulfide). The Roman numerals indicate the charge of the ion rather than the number of ions present in the compound.
  - When more than one polyatomic ion appears in the formula of an ionic compound, the usual rules of naming apply, but use brackets to separate the ions in the formula (e.g. ammonium carbonate is $(NH_4)_2CO_3$).
  - For hydrates, use the prefixes *mono-, di-, tri-, tetra-, penta-, hexa-, hepta-, octa-, nona-* and *deca-* to indicate the number of water molecules in the compound (e.g. $CuSO_4 \cdot 5H_2O$ is copper(II) sulfate pentahydrate).

Crystal formation and uses of ionic compounds

- The size of a crystal depends on the conditions during its formation.

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4.5.2 Key terms

anion a negatively charged atom or group of atoms
binary describes compounds made up of only two elements
cation a positively charged atom or group of atoms
chemical bond arrangement of electrons between two atoms that generates a force, causing the atoms to be bound to each other
covalent bond sharing of electrons between nuclei that bonds them together in a molecule
electrolyte a liquid that can conduct electricity

electronegativity the electron-attracting power of an atom, determined by a number of factors including the size of the atom, the charge on the nucleus and the number of electrons in the atom; atoms with high electronegativities are able to attract electrons easily, whereas atoms with low electronegativities do not attract electrons as readily
electrostatic attraction strong force of attraction between positive and negative ions that produces an ionic bond
hydrates adds water
ion a charged atom
ionic bond bond between cations and anions in an ionic compound
ionic compound compound containing cations and anions in an ionic lattice structure or an ionic liquid, depending on ion size
metallic bond positively charged metal cations arranged in a lattice with delocalised valence electrons being able to flow around them
polyatomic ion charged ion composed of two or more atoms
salt ionic compound consisting of a metal ion and a non-metal ion, except oxides and hydroxides

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Digital document Key terms glossary — Topic 4 (doc-30935)

4.5.3 Practical work and investigations

Experiment 4.1

Investigating calcite crystals

Aim: To observe the nature of crystal structure using calcite crystals

Digital document: doc-30867
Teacher-led video: tlvd-0618

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Digital documents Practical investigation logbook — Topic 4 (doc-30936)
Experiment 4.2 Ionic models (doc-30868)

4.5 Exercises

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4.5 Exercise 1: Multiple choice questions

1. What happens when an atom loses an electron?
   - A. A positive ion is formed
   - B. A negative ion is formed
   - C. The atomic number changes
   - D. The atomic mass changes
2. Sodium atoms and sodium ions
   - A. are identical in their chemical properties.
   - B. have the same number of electrons.
   - C. have the same number of protons.
   - D. have the same number of protons and electrons.

3. The elements sodium, caesium and lithium form ions with the charge in their compounds of
   - A. 1+ only.
   - B. 2+ only.
   - C. 3+ only.
   - D. 1+ or 2+ only.
4. Which of the following ions does not have the same electron configuration as an oxide ion, O^{2-}?
   - A. N^{3-}
   - B. F^-
   - C. S^{2-}
   - D. Al^{3+}
5. An unknown substance was found to combine with chlorine to form a compound of formula XCl_3. How many outer shell electrons does X have?
   - A. 1
   - B. 2
   - C. 3
   - D. 4
6. An atom of X, which has two electrons in its outermost shell, is in contact with an atom of Y, which has seven valence electrons. Which of the following is *incorrect*?
   - A. X^{2+} and Y^- ions are present in the compound.
   - B. The compound has the formula XY_2.
   - C. The compound probably has high melting and boiling points.
   - D. No compound can be formed between these two elements.
7. An element has an atomic number of 20. What is its electron configuration when the element reacts to form an ionic compound?
   - A. 2, 8, 8
   - B. 2, 8, 8, 2
   - C. 2, 8, 8, 4
   - D. 2, 8, 10
8. The electron configurations of four pairs of elements are given below. Which pair of elements is most likely to form an ionic bond?
   - A. 2, 8, 1 and 2, 8, 2
   - B. 2, 8, 1 and 2, 8, 7
   - C. 2, 8, 2 and 2, 8, 4
   - D. 2, 8, 7 and 2, 8, 6
9. What does the formation of an ionic compound from a reaction between atoms of two elements involve?
   - A. Sharing of pairs of electrons between atoms
   - B. Donation of outer shell electrons to the entire crystal structure
   - C. Transfer of electrons between atoms
   - D. Ionisation of the atoms of some elements
10. What is the structure of solid calcium chloride best described as?
    - A. A lattice consisting of diatomic chlorine molecules strongly bonded to calcium atoms.
    - B. A lattice of calcium and chloride ions, strongly bonded, in the ratio 1:2.
    - C. Discrete molecules of calcium chloride with strong bonding within the molecule but weak bonding between molecules.
    - D. An infinite lattice in which calcium and chloride are linked by strong covalent bonds.
11. What is the formula of the compound ammonium phosphate?
    - A. $(NH_4)_3PO_4$
    - B. NH_3P
    - C. NH_4PO_4
    - D. $(NH_4)_2PO_3$

12. What is the compound CuO called?
    - A. Copper oxide
    - B. Copper dioxide
    - C. Copper(II) oxide
    - D. Copper(I) oxide.
13. What are the blue crystals of $CuSO_4 \cdot 5H_2O$ called?
    - A. Copper sulfate pentahydrate
    - B. Copper sulfate hexahydrate
    - C. Copper(II) sulfate pentahydrate
    - D. Copper(I) sulfate hexahydrate

4.5 Exercise 2: Short answer questions

1. Complete the following table.

Name of atom	Symbol for atom	Electron configuration of atom	Symbol for ion	Electron configuration of ion
Lithium				
Beryllium				
Nitrogen				
Oxygen				
Fluorine				
Sodium				
Magnesium				
Aluminium				
Phosphorus				
Sulfur				
Chlorine				
Potassium				
Calcium				

2. a. What is the electron configuration of the calcium atom?
   b. What is the electron configuration of the fluorine atom?
   c. Describe what happens when the calcium and fluorine atoms react chemically.
   d. Use an electron shell diagram to illustrate your answer to part (c).
   e. What holds the ions together in a crystal?
   f. What is the ratio of calcium ions to fluoride ions?
3. Cations form from metallic elements. Using an example, show how this occurs.
4. In a sample of potassium chloride (an important electrolyte in the human body), the number of K^+ ions and Cl^- ions are equal. In a sample of potassium oxide, however, the numbers of each ion are not equal. State the ratio of each ion in the potassium oxide and explain why they are not equal.
5. Write simple equations and electron shell diagrams to show what happens when the following pairs of atoms bond.
   a. K and F
   b. Al and O
   c. Be and Cl

6. Explain why ionic salts conduct electricity in the molten and aqueous states but not in the solid state.
7. Draw diagrams to illustrate the brittle nature of salt.
8. Explain why ionic substances have high melting points and are usually crystalline solids at room temperature.
9. A compound made up of *X* and *Y* atoms has the following properties.
   - It has a high melting point and boiling point.
   - It is very soluble in water but not in kerosene.
   - It conducts electricity in the liquid or aqueous state but not in the solid state.
   - The electron configuration of an atom of *X* is 2, 8, 2, while that of *Y* is 2, 6.

   a. What is the formula for the compound?
   b. What type of bonding must it have?
   c. Write a simple equation showing how the bonding is attained.
10. Using table 4.3 (in section 4.3.3), write the formulas for the following compounds.
    a. Sodium fluoride
    b. Barium nitrate
    c. Iron(III) hydroxide
    d. Sodium sulfide
    e. Aluminium oxide
    f. Calcium hydride
    g. Copper(II) sulfate
    h. Ammonium hydroxide
    i. Chromium(III) oxide
    j. Calcium nitrate
    k. Lithium chloride
    l. Potassium cyanide
    m. Sodium hydrogen phosphate
11. Give the names of the following compounds.
    a. $ZnCl_2$
    b. $Al_2(CO_3)_3$
    c. Na_2SO_4
    d. $AgNO_3$
    e. NaOH
12. Give the chemical name and formula for each of the following compounds.
    a. Soda ash, the common name of a compound containing sodium ions and carbonate ions
    b. Baking soda, commonly used in baking cakes and composed of sodium and hydrogen carbonate ions
    c. Chalk, marble and limestone, all composed of calcium ions and carbonate ions
13. Write the formulas for the following compounds.
    a. Magnesium sulfate heptahydrate
    b. Sodium carbonate decahydrate
    c. Zinc chloride hexahydrate
    d. Barium chloride dihydrate
14. Name the following compounds.
    a. $BaCl_2{\cdot}3H_2O$
    b. $LiCl{\cdot}4H_2O$
    c. $CoCl_2{\cdot}5H_2O$
15. Sodium chloride is used as a typical example of an ionic compound. Research the effects of too little or too much sodium chloride on the body.

4.5 Exercise 3: Exam practice questions

Question 1 (3 marks)

Predict the charges on the ions formed when the following atoms react with other atoms.

a. Ca **1 mark**
b. P **1 mark**
c. Al **1 mark**

Question 2 (2 marks)

Oxygen gas, which forms about 20% of the Earth's atmosphere and is essential to life, exists as O_2 molecules. Explain why ionic bonding does not exist between two atoms of oxygen.

Question 3 (4 marks)

Explain what happens to the size of the atom in the following situations.

a. It loses one or more electrons in forming an ionic bond. **2 marks**
b. It gains one or more electrons in forming an ionic bond. **2 marks**

4.5 Exercise 4: studyON Topic Test online only

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5 Quantifying atoms and compounds

5.1 Overview

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5.1.1 Introduction

Atoms are extremely tiny. Even though the air is full of oxygen and nitrogen molecules, you cannot see them. We know a lot about atoms and molecules, and this knowledge is invaluable when explaining the properties of substances. But, how do we measure atoms? The scale of atomic size means that chemists rarely deal with atoms one at a time. What is needed is a convenient way to compare and measure masses and amounts of particles in elements and compounds. The mole concept, which is central to most chemical calculations, helps us to work with the vast numbers of atoms that are present in different types of substances.

FIGURE 5.1 From 6:02 am to 6:02 pm on 23 October is International Mole Day, because the time and date represent the number for the mole, or Avogadro's constant — with the time of 6.02, in the 10th month on the 23rd day (6.02×10^{23}).

5.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- the relative isotopic masses of elements and their representation on the relative mass scale using the carbon-12 isotope as the standard, and the reason for the selection of carbon-12 as the standard
- determination of the relative atomic mass of an element using mass spectrometry (details of instrument not required)
- the mole concept, Avogadro constant, determination of the number of moles of atoms in a sample of known mass, and calculation of the molar mass of ionic compounds
- experimental determination of the empirical formula of an ionic compound.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of learning and assessment. Experiments and investigations, supported by a **Practical investigation logbook** and **Teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

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Digital documents Key science skills (doc-30903)

Key terms glossary — Topic 5 (doc-30938)

Practical investigation logbook (doc-30939)

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5.2 Relative isotopic mass and the carbon-12 scale

KEY CONCEPTS

- The relative isotopic masses of elements and their representation on the relative mass scale using the carbon-12 isotope as the standard, and the reason for the selection of carbon-12 as the standard
- Determination of the relative atomic mass of an element using mass spectrometry (details of instrument not required)

5.2.1 Review of atomic number and mass number

Recall from topic 1 that atoms are composed of a nucleus containing both protons (positive charge) and neutrons (no charge), which are significantly larger than the electrons (negative charge) that occur in shells around the nucleus. All atoms of the same element have the same number of protons in their nucleus. This is the atomic number (*Z*). Protons have approximately the same mass as neutrons, and because the mass of electrons is so small, the mass number (*A*) of an element is the sum of the protons and neutrons. Hence an element can be represented by the following:

$$\begin{array}{r}\text{mass number} \rightarrow A \\ \text{atomic number} \rightarrow Z\end{array} E$$

5.2.2 Relative isotopic mass

The mass of one hydrogen atom is 1.66×10^{-24} g. This incredibly small size is difficult to manipulate, so a scale based on a multiple of the mass of a single atom of a carbon-12 (^{12}C) isotope was devised. By international agreement the mass of an atom of carbon-12 is given as 12 **unified atomic mass units** (u). Therefore, 1 u is $\frac{1}{12}$th the mass of an atom of carbon-12. The carbon-12 isotope is very stable and abundant on Earth.

For example, an atom of hydrogen has a relative mass of 1.0 because the mass of a hydrogen atom ($1.660539040 \times 10^{-24}$ g) is 12 times less (11.999998187) than the mass of a single atom of ^{12}C ($1.992646547 \times 10^{-23}$ g). Mass calculations relative to carbon-12 are demonstrated this table.

Element	Protons	Neutrons	Protons + neutrons	Mass relative to carbon-12
$^{1}_{1}$H	1	0	1	$\frac{1}{12} \times 12 = 1$
$^{12}_{6}$C	6	6	12	$\frac{12}{12} \times 12 = 12$
$^{59}_{27}$Co*	27	32	59	$\frac{59}{12} \times 12 = 59$

* While cobalt, Co, has 22 isotopes, there is only one naturally occurring isotope, ^{59}Co.

Isotopes are atoms of the same element that contain the same number of protons but different numbers of neutrons. Individual isotopes each have a **relative isotopic mass**. A few elements have no variation in the number of neutrons and so have only one relative isotopic mass. Here the relative isotopic mass is simply the mass number of the element.

Two equivalent definitions can be used to define relative isotopic mass.

FIGURE 5.2 Isotopes that have a large difference between the number of protons and neutrons may be unstable and radioactive. Promethium-147 has 61 protons to 86 neutrons, and Thallium-204 has 81 protons to 123 neutrons.

Relative isotopic mass can be defined as

- **the mass of an atom of a single isotope relative to the mass of an atom of a single isotope of carbon-12, which has a mass of exactly 12 units**

or

- **the mass of an atom of a single isotope relative to $\frac{1}{12}$ of the mass of an atom of a single isotope of carbon-12.**

Therefore,

$$\textbf{Relative isotopic mass} = \frac{\textbf{mass of an atom of a single isotope}}{\frac{1}{12}\ \textbf{mass of a single carbon} - \textbf{12 atom}}$$

on Resources

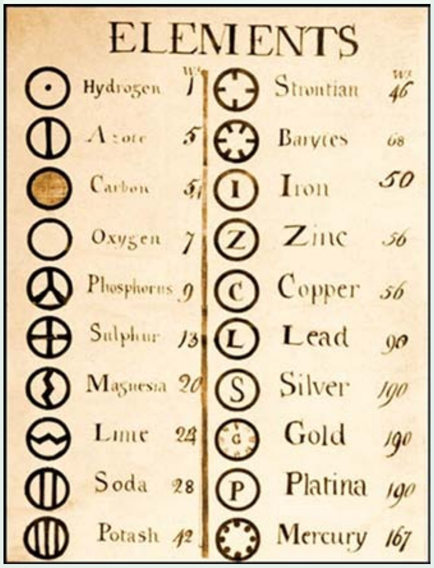
Interactivity Relative isotopic mass (int-6353)

FINDING A REFERENCE FOR THE MASS OF ATOMS

Atoms are extremely small but they still have mass. The problem is to find a way of measuring that mass. In 1803, English chemist John Dalton described matter as made up of particles that were solid, indivisible and having weight. Dalton devised his own set of symbols for the elements and prepared a table of atomic weights by experimentally determining relative weights of elements in compounds. He based his figures on hydrogen having a mass of 1. We now use the term 'mass' instead of 'weight'.

This reference standard was later changed from hydrogen to oxygen, but this led to a disagreement between chemists, who used natural oxygen as the standard, and physicists, who used the oxygen-16 isotope only. Having two slightly different lists of atomic masses caused many problems. An agreement was reached in 1961 to change the standard to carbon-12, and this settled the dispute; this change had the added advantage that carbon-12 can be measured very accurately because it is stable and abundant. The development of an instrument called the mass spectrometer allowed scientists to accurately compare the masses of all atoms.

FIGURE 5.3 Dalton's table of elemental symbols and atomic weights.



ELEMENTS

	Wt		Wt
Hydrogen	1	Strontian	46
Azote	5	Barytes	68
Carbon	5,	Iron	50
Oxygen	7	Zinc	56
Phosphorus	9	Copper	56
Sulphur	13	Lead	90
Magnesia	20	Silver	190
Lime	24	Gold	190
Soda	28	Platina	190
Potash	42	Mercury	167

5.2.3 Relative atomic mass

Most elements consist of a mixture of isotopes. The **relative atomic mass, A_r,** of an element can also be defined in two equivalent ways.

Relative atomic mass, A_r, can be defined as

- **the weighted* average of a single atom of the masses of its isotopes on a scale on which a carbon-12 atom has a mass of exactly 12 units**

or

- **the weighted* average of a single atom of the masses relative to $\frac{1}{12}$ of the mass of an atom of a single isotope of carbon-12.**

Therefore,

$$\text{Relative atomic mass} = \frac{\text{weighted average of a single atom of the masses of its isotopes}}{\frac{1}{12}\ \text{mass of a single carbon} - 12\ \text{atom}}$$

* Weighted average allows for the fact that isotopes of the same element have different natural percentage abundances. It is usual practice to neglect the tiny contribution to atomic mass from electrons and take both the mass of a proton and a neutron as being equal to 1 u.

Relative masses, M_r

Relative formula mass (M_r) uses the same symbol as relative molecular mass (M_r) and is equal to the sum of the relative atomic masses of the atoms in a formula unit of the compound. For covalent compounds, it is called the relative molecular mass; for ionic compounds, it is called the relative formula mass.

SAMPLE PROBLEM 1

Calculate the M_r of the following compounds

a. SO_2

b. $Mg(OH)_2$

c. NH_4NO_3

Teacher-led video: SP1 (tlvd-0521)

THINK	WRITE
a. Look up the relative atomic mass (A_r) of the elements that make up SO_2 using the periodic table. SO_2: 1 × S and 2 × O Given that this is a covalent molecule, the term *relative molecular mass* would be used. No units are required.	1 × S and 2 × O $M_r = 32.0 + (2 \times 16.0) = 64.0$ $M_r = 64.0$
b. Look up the relative atomic mass of the elements that make-up $Mg(OH)_2$. $Mg(OH)_2$: 1 × Mg and 2 × O and 2 × H Remember that $(OH)_2$ is the same as 2 × OH. Given that this is a ionic molecule the term *relative formula mass* would be used. No units are required.	1 × Mg and 2 × O and 2 × H $M_r = 24.3 + (2 \times 16.0) + (2 \times 1.0)$ $M_r = 58.3$
c. Look up the relative atomic mass of the elements that make-up NH_4NO_3. NH_4NO_3: 2 × N and 4 × H and 3 × O Given that this is a covalent molecule, the term *relative molecular mass* would be used.	2 × N and 4 × H and 3 × O $M_r = (2 \times 14.0) + (4 \times 1.0) + (3 \times 16.0)$ $M_r = 80.0$

PRACTICE PROBLEM 1

Calculate the M_r of the following compounds:

a. **$Ca_3(PO_4)_2$**
b. **WC**
c. **$Na_2Cr_2O_7$**

TIP: **Using significant figures.**

When adding and subtracting with decimal places, you should look to the right of the decimal point only. Round your answer to the fewest significant digits to the right of the decimal; for example:

266.38	**+**	**743.394**	**−**	**0.000357**	**=**	**1009.77**
2 decimal places		**3 decimal places**		**6 decimal places**		**2 decimal places**

When multiplying and dividing, you should retain as many significant figures as in the number with the fewest significant digits; for example:

8950.23	**×**	**39.8**	**÷**	**11.238**	**=**	**3.17×10^4**
6 sig. figs		**3 sig. figs**		**5 sig. figs**		**3 sig. figs.**

The first non-zero digit, reading from left to right in a number, is the first significant figure.

5.2.4 Mass spectrometry

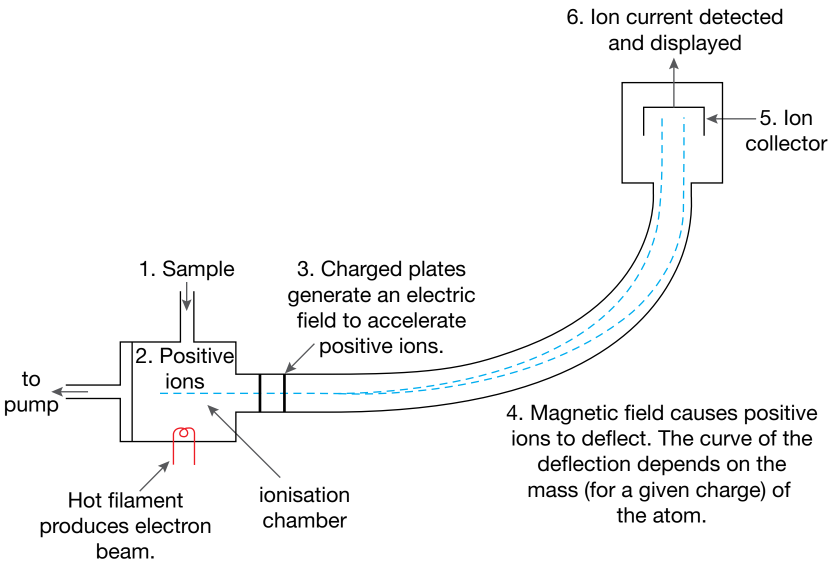
In 1913, Joseph John Thompson, an English physicist, discovered that some elements can have isotopes — atoms of the same element with a different mass. Francis Aston, another English scientist, developed Thompson's equipment into a **mass spectrometer**. Francis Aston was later awarded the Nobel Prize for identifying the 212 naturally occurring isotopes by comparing the relative masses of atoms.

Figure 5.4 shows the main features of a mass spectrometer. A sample of the element to be analysed is injected as a gas into the ionisation chamber (1), where the atoms are ionised by bombardment with electrons produced by the hot filament (2). The positive ions formed are accelerated through an electric field (3) and deflected in a magnetic field that forces the ions to travel along different paths (4). The curved paths of deflection depend on the mass-to-charge ratio of the ions. For a given charge (for example, singly charged ions), the heavier ions are harder to deflect and so travel in a wider curve. Ions corresponding to a fixed mass-to-charge ratio are picked up by the ion collector (5) and the ion current is amplified and displayed (6). That is, the path of the lightest element is deflected the most. Similarly, it deflects the most highly charged particle the furthest.

The mass spectrometer provides us with information about:

- the number of isotopes in a given sample of an element
- the relative isotopic mass of each isotope
- the percentage abundance of the isotopes.

FIGURE 5.4 A mass spectrometer. Different types of atoms of the same element can be separated on the basis of their mass:charge ratio.



Resources

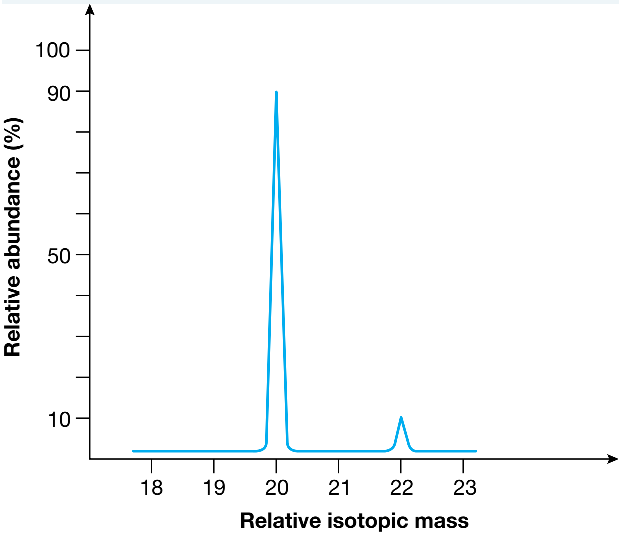
Weblink Mass spectrometry

Figure 5.5 shows the mass spectrum for neon. The two peaks in the trace represent the two isotopes of neon. Both isotopes of neon have an atomic number of 10 (that is, the number of protons is 10), but one isotope has a relative isotopic mass of 20 and the other has a relative isotopic mass of 22. The atoms of neon-20 have 10 protons and 10 neutrons, and atoms of neon-22 have 10 protons and 12 neutrons. The number of protons added to the number of neutrons in an atom is called its **mass number**.

If an ion appears on the mass spectrum, it must have a positive charge; for example, $^{20}Ne^+$ or $^{22}Ne^+$.

Using data from the mass spectrum, one could calculate the relative atomic mass for an element by using the following method of calculation.

FIGURE 5.5 The mass spectrum for neon



To calculate relative atomic mass:

$$A_r = \frac{\textbf{(relative isotopic mass of first isotope} \times \textbf{abundance)} + \\ \textbf{(relative isotopic mass of second isotope} \times \textbf{abundance)} + \ldots}{\textbf{100}}$$

For example, from the mass spectrum discussed for neon, the two isotopes, ^{20}Ne and ^{22}Ne, have relative isotopic masses of 20.0 and 22.0 respectively. Their relative abundances are 90.0% and 10.0% respectively.

$$\begin{aligned} A_r(\text{Ne}) &= \frac{(20.0 \times 90.0) + (22.0 \times 10.0)}{100} \\ &= 20.2 \text{ (no units because this is a relative value)} \end{aligned}$$

SAMPLE PROBLEM 2

Lithium consists of two isotopes. One isotope, ^{6}Li has a relative isotopic mass of 6.02 and an abundance of 7.42%. The other isotope, ^{7}Li, has a relative isotopic mass of 7.02 and an abundance of 92.58%.

a. Is the relative atomic mass closer to six or seven? Explain your answer.

b. Calculate the relative atomic mass of lithium.

Teacher-led video: SP2 (tlvd-0522)

THINK	WRITE
a. Nearly 93% of all the isotopes of lithium have a mass of seven, so when the masses are averaged it will be much nearer to seven than six.	Closer to 7
b. 1. Recall the relative atomic mass formula and substitute known values.	$A_r(\text{Li}) = \frac{(7.59 \times 6.02) + (92.41 \times 7.02)}{100}$ $= 6.944$
2. Round your answer to three significant figures because this is the fewest number of significant figures used in the question. No units are required because this is a relative value.	$A_r(\text{Li}) = 6.94$

PRACTICE PROBLEM 2

Copper has two isotopes, ^{63}Cu, which has the relative isotopic mass of 62.93, and ^{65}Cu, which has the relative isotopic mass of 64.92. The proportions of each isotope are 69.15% and 30.85% respectively. Calculate the relative atomic mass, A_r, of copper.

5.2 EXERCISE

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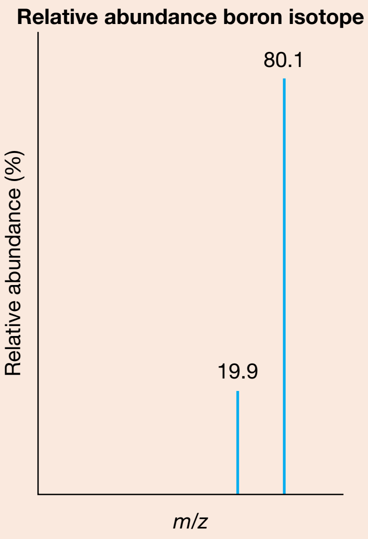
1. Calculate M_r of each of the following compounds.
   (a) CO_2
   (b) NaCl
   (c) H_2O_2
   (d) H_2SO_4
   (e) $C_6H_{12}O_6$
2. Three isotopes of magnesium and their relative abundances are $^{24}_{12}Mg$ (78.8%), $^{25}_{12}Mg$ (10.2%) and $^{26}_{12}Mg$ (11.0%). The masses of each isotope was measured to three significant figures: 24.0, 25.0 and 26.0 respectively.
   (a) Sketch on a graph the mass spectrum for magnesium.
   (b) Calculate the relative atomic mass of magnesium.
3. Gallium has two isotopes. One isotope is $^{69}_{31}Ga$ and has a relative abundance of 60.50%. The relative atomic mass of gallium is 69.70. Find the relative isotopic mass of the other isotope.
4. The green colour of fireworks can be produced using the element boron. Boron's two isotopes are $^{10}_{5}B$ and $^{11}_{5}B$. Use the mass spectrum provided in the figure to calculate the relative atomic mass of boron.



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5.3 The mole

KEY CONCEPT

- The mole concept, Avogadro constant, determination of the number of moles of atoms in a sample of known mass, and calculation of the molar mass of ionic compounds

5.3.1 The mole concept

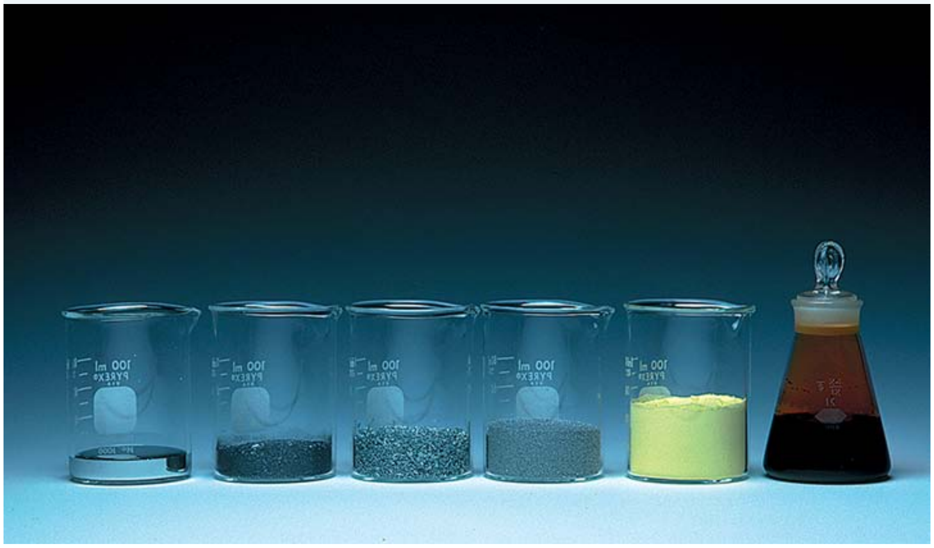
If one drop of water, H_2O, has a mass of 0.0300 g then this would contain 1.00×10^{21} molecules of water. Dealing with such large values would prove overly complicated, and ratios of masses of different substances would be difficult to see. Clearly, in chemistry, we need a measurement unit that can manage these large numbers of atoms and molecules. The **mole** concept helps us count atoms and is fundamental to chemical calculations. The term 'mole' simply represents a number. It is a unit of measurement, just as the 'kilo' represents 1000.

The Avogadro constant

The mole was recently redefined, with the new definition implemented on 20 May 2019. Prior to this date, the mole was defined as the amount of substance that contains as many particles (atoms, ions or molecules) as there are atoms in exactly 12 g of the ^{12}C isotope. The number of carbon atoms in 12 g of ^{12}C isotope has been experimentally calculated to be 602 140 760 000 000 000 000 000, or 6.02×10^{23} mol^{-1}.

However, in 2018 the General Conference on Weights and Measures determined that all SI (Système international (d'unités)) will not be defined in terms of any physical objects but rather defined by constants (due to their exact nature).

FIGURE 5.6 One mole each of mercury, zinc, silicon, aluminium, sulfur and bromine. The quantity of each of the substances shown contains 6.02 × 10^{23} particles.



The mole, abbreviation mol, symbol *n*, is the SI unit of amount of substance.

One mole contains exactly $6.022\,140\,76 \times 10^{23}$ elementary entities.

This number is the fixed numerical value of the Avogadro constant, symbol N_A or *L*, when expressed in mol^{-1}, and is called the **Avogadro constant, N_A**, (or number). Therefore, the number of atoms in 1 mole of a pure element is 6.02×10^{23} mol^{-1}.

One mole contains exactly $6.022\,140\,76 \times 10^{23}$ elementary entities.
It is typically rounded to three significant figures.
Avogadro's constant, N_A or $L = 6.02 \times 10^{23}$ mol^{-1}

TIP: Avogadro's constant can be found in Table 4 of the VCE Chemistry Data Book.

While the definition of the mole has changed, from the number of atoms in exactly 12 g of the ^{12}C isotope, to a constant of the same value, the new definition is fundamentally based on the ^{12}C isotope. You may have noticed that the definition of relative isotopic mass is also based on the ^{12}C isotope. A good reason exists for this. The relative isotopic scale allows us to deal with atoms 'one at a time'. The mole concept allows us to deal with much larger groups 'one mole at a time'. The common reference to ^{12}C means that, to find out the mass of one mole (molar mass, *M*) of any element, all you have to do is add 'g' to the relative atomic mass.

FIGURE 5.7 As well as being the basis for the mole, purified carbon is used in air, gold and water purification, decaffeination, metal extraction, poison and sewage treatment, and in air filters in gas masks.

To find the mass of one mole of an element, simply add ‘g’ to the relative atomic mass of that element.

AMEDEO AVOGADRO

Amedeo Avogadro is one of Italy’s most famous modern scientists. In this commemorative stamp, celebrating 100 years since his death, the text translates to ‘Equal volumes of gas under the same temperature and pressure conditions contain the same number of molecules’. This statement is known as Avogadro’s law. In tribute to his work on molecular theory, Avogadro’s constant, N_A or L, relates the molar mass of a substance (molecule, atom or ion) to the mass of a sample. One mole of any sample will contain exactly 6.022 140 76 × 10^{23} particles. Avogadro’s constant is a recognised SI unit.

5.3.2 Molar mass, *M*

The **molar mass** (M) of an element is defined as the mass of 1 mol of the element; that is, 6.02 × 10^{23} atoms of the element. The unit is grams per mole (g mol^{-1}). For example:

$$\begin{aligned}\text{Molar mass of carbon atoms} &= \text{mass of 1 mol of C atoms}\\ &= 12.0\ \text{g mol}^{-1}\end{aligned}$$

$\therefore$ 12.0 g of carbon contains 6.02×10^{23} atoms of carbon.

$$\begin{aligned}\text{Molar mass of oxygen atoms} &= \text{mass of 1 mol of O atoms}\\ &= 16.0\ \text{g mol}^{-1}\end{aligned}$$

$\therefore$ 16.0 g of oxygen contains 6.02×10^{23} atoms of oxygen.

$$\begin{aligned}\text{Molar mass of oxygen molecules} &= 2 \times \text{mass of 1 mol of O atoms}\\ &= 2 \times 16.0\ \text{g mol}^{-1}\\ &= 32.0\ \text{g mol}^{-1}\end{aligned}$$

$\therefore$ 32.0 g of oxygen contains 6.02×10^{23} molecules of oxygen.

32.0 g of oxygen contains 1.20×10^{24} atoms of oxygen.

5.3.3 Calculating molar mass, *M*, of covalent (molecular) and ionic compounds

The molar mass of a compound is defined as the mass of 1 mol of the compound expressed in grams per mole (g mol^{-1}).

Covalent compounds

For compounds consisting of molecules, the molar mass (M) is numerically equal to the relative molecular mass (M_r) expressed in grams per mole (g mol^{-1}). For example:

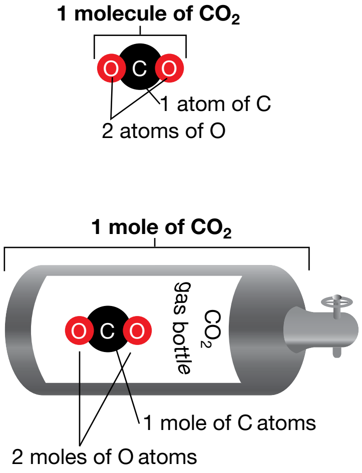
$$\text{Molar mass of water molecules} = \text{mass of } 1.00 \text{ mol of } H_2O \text{ molecules}$$
$$= 18.0 \text{ g mol}^{-1}$$

$\therefore$ 18.0 g of water contains 6.02×10^{23} molecules of water.

$$\text{Molar mass of chlorine gas} = \text{mass of } 1.00 \text{ mol of } Cl_2 \text{ molecules}$$
$$= 71.0 \text{ g mol}^{-1}$$

$\therefore$ 71.0 g of chlorine gas contains 6.02×10^{23} molecules of chlorine.

FIGURE 5.8 Whether a substance is made up of atoms, molecules or ions, the same principle applies: one mole of any substance always contains 6.02×10^{23} particles.



Ionic compounds

The M_r of an ionic compound is found by adding the A_r of each atom in the formula of the compound. For example:

$$\begin{aligned} M_r(CuSO_4) &= A_r\,(Cu) + A_r\,(Cu) + (4 \times A_r\,(O)) \\ &= 63.5 + 32.1 + (4 \times 16.0) \\ &= 159.6 \end{aligned}$$

The molar mass (M) of $CuSO_4$ is 159.6 g mol^{-1}.

$$M(CuSO_4) = 159.6 \text{ g mol}^{-1}$$

SAMPLE PROBLEM 3

What is the molar mass, *M*, of calcium nitrate, $Ca(NO_3)_2$?

Teacher-led video: SP3 (tlvd-0523)

THINK

1. The M_r of an ionic compound is found by adding the A_r of each atom in the formula of the compound.
   Brackets indicate the number of multiples; e.g. $(NO_3)_2$ is the same as $NO_3 + NO_3$.

WRITE

Atom	Number		A_r		Total A_r
Ca	1	×	40.1	=	40.1
N	2	×	14.0	=	28.0
O	6	×	16.0	=	96.0

2. Add the total A_r to determine the M_r.
   TIP: Molar mass must have units of g mol^{-1}, unlike relative masses which have no units.

$$\begin{aligned} M_r &= 40.1 + 28.0 + 96.0 \\ &= 164.1 \end{aligned}$$
$$M_r(Ca(NO_3)_2 = 164.1 \text{ g mol}^{-1}$$

PRACTICE PROBLEM 3

What is the molar mass, M, of calcium perchlorate, $Ca(ClO_4)_2$?

5.3.4 Calculating moles of atoms in a known mass

We have discussed a relationship between the number of particles and the mass of the substance. This means that, when we measure the mass of a substance equal to its molar mass (e.g. 18.0 g of H_2O), we are also indirectly counting out 6.02×10^{23} particles of that substance (e.g. water molecules). If the molecule or formula unit contains more than one kind of atom, we can also calculate the number of different atoms present in one mole of the substance. For example:

1.00 mol of carbon dioxide molecules contains:

- 1.00 mol of CO_2 molecules
- 6.02×10^{23} CO_2 molecules
- 1.00 mol of C atoms
- 6.02×10^{23} C atoms
- 2.00 mol of O atoms
- 1.20×10^{24} O atoms.

The molar mass of CO_2 is 44.0 g mol^{-1}.

These observations suggest the following:

$$\text{mass} \underset{\text{multiply by molar mass}}{\overset{\text{divide by molar mass}}{\rightleftarrows}} \text{moles}$$

To calculate moles from a known mass:

$$n = \frac{m}{M}$$

where n = number of moles, m = mass, M = molar mass.

$$\text{number of particles } (N) \underset{\text{multiply by } 6.02\times10^{23}}{\overset{\text{divide by } 6.02\times10^{23}}{\rightleftarrows}} \text{moles}$$

To calculate moles from a known number of particles:

$$n = \frac{\textbf{number of particles}}{6.02 \times 10^{23}} = \frac{N}{N_A}$$

where n = number of moles, N = number of particles, N_A = Avogadro's constant.

SAMPLE PROBLEM 4

A gas balloon contains 22 g of helium atoms. How many helium atoms are present in the balloon?

Teacher-led video: SP1 (tlvd-0524)

THINK	WRITE
1. To determine the number of helium atoms, first use the formula $n = \frac{m}{M}$ to first determine the number of mole of helium atoms.	$n(\text{He}) = \frac{m}{M}$ $= \frac{22}{4.0}$ $= 5.5\ \text{mol}$
2. To determine the number of helium atoms in the calculated number of moles of helium rearrange the formula $n = \frac{N}{N_A}$. **TIP:** Avogadro's constant, N_A, can be found in Table 4 of the VCE Chemistry Data Book.	$n = \frac{N}{N_A}$ $N(\text{He}) = n \times N_A$ $= 5.5 \times 6.02 \times 10^{23}$ $= 3.3 \times 10^{24}$ 3.3×10^{24} atoms of helium are in 22 g of helium.

PRACTICE PROBLEM 4

How many atoms of oxygen are in 66.05 g of ammonium sulfate, $(NH_4)_2SO_4$?

on Resources

Interactivity The mole concept (int-6354)

Weblinks What is mole day?
How big is a mole?

5.3 EXERCISE

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1. Calculate the molar masses of the following.
   - (a) Ozone, O_3
   - (b) Iodine, I_2
   - (c) Calcium oxide, CaO
   - (d) Hydrogen chloride, HCl
2. What is the mass of each of the following?
   - (a) 5.25×10^{24} molecules of glucose, $C_6H_{12}O_6$
   - (b) 1.83×10^{21} molecules of nitrogen dioxide, NO_2
   - (c) 3.56×10^{14} molecules of carbon dioxide, CO_2
   - (d) 4.13×10^{28} molecules of carbon disulfide, CS_2
   - (e) 3.62×10^{24} molecules of dinitrogen tetraoxide, N_2O_4
3. Which of the following substances has the greatest mass?
   - (a) 200 g of magnesium
   - (b) 5.00 mol of sulfur
   - (c) 1.2×10^{24} atoms of helium
   - (d) 3.5×10^{22} molecules of alanine, $C_3H_7O_2N$

4. Use a spreadsheet program to produce a spreadsheet that converts amounts of substances to moles. Set up your spreadsheet according to the following template, and then complete the table. Use a formula and then the 'fill' function.

Conversion of amounts of substances to moles

	A	B	C	D	E	F	G
	Substance	**Amount (g)**	**Molar mass (*M*)**	**Number of atoms in the molecule**	**Number of moles (*n*) of substance**	**Number of molecules**	**Total number of atoms**
1	Water, H_2O	3.2					
2	Methane, CH_4	2.7					
3	Ammonia, NH_3	0.056					
4	Ethanoic acid, CH_3COOH				27.3		
5	Benzene, C_6H_6				0.56		
6	Octane, C_8H_{18}				2.34		
7	Ethanol, CH_3CH_2OH					6.0×10^{24}	
8	Ozone, O_3					1.27×10^{21}	
9	Sulfuric acid, H_2SO_4						3.0×10^{26}
10	Carbon dioxide, CO_2						7.5×10^{25}

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5.4 Determination of the empirical formula

KEY CONCEPT

- Experimental determination of the empirical formula of an ionic compound

5.4.1 Steps to determine the empirical formula

The **empirical formula** of a compound gives the simplest whole number ratio of the atoms or ions present in the compound. For ionic compounds, their formulas are already in the simplest ratio of elements (because it represents the ratio of elements in the lattice structure) and so the formula for an ionic compound is its empirical formula.

Empirical formulas can only be found experimentally by determining the mass of each element present in the total mass of the compound. This *ratio of elements by mass* must be converted to a *ratio of elements by numbers*. This is done by calculating the number of moles of each element.

The steps involved in finding an empirical formula are:

1. Write down the *symbols* of the elements present.
2. Assume that the *mass* of the sample is 100 g and all percentages become grams.
3. Convert masses to *moles*.
4. Find the simplest whole number *ratio* of the atoms by dividing all numbers of moles by the smallest number of moles.
5. If necessary, multiply by a *factor* to convert all numbers to whole numbers.

Empirical formulas can also be calculated when experiments reveal the actual mass of each element that is present in a sample of a compound. In this situation, the masses involved are simply written in at step 2 in the preceding sequence of steps.

on Resources

Digital document Experiment 5.1 Empirical formula of magnesium oxide (doc-30846)

Empirical formula of hydrated compounds

Some ionic compounds crystallise from an aqueous solution to form a **hydrated** ionic compound. In these compounds, water molecules are included in the crystal lattice structure. This water is called **water of crystallisation**. Hydrated copper(II) sulfate, for example, appears as a blue crystalline solid and has the formula $CuSO_4{\cdot}5H_2O$ (see figure 5.9 (a)). This means that, for each Cu^{2+} ion and $SO_4^{\ 2-}$ ion in the crystal lattice structure, five water molecules are also included. Heating the copper(II) sulfate crystals removes the water molecules to leave a white powder known as **anhydrous** copper(II) sulfate. The formula of anhydrous copper(II) sulfate is $CuSO_4$ (figure 5.9 (b)). The mass of the water of crystallisation, and its percentage contribution to the mass of the ionic compound, can be calculated when the masses of both the hydrated and anhydrous compound are known or if the degree of hydration is shown in the formula.

FIGURE 5.9 (a) Hydrated copper(II) sulfate, $CuSO_4{\cdot}5H_2O$, is also called copper(II) sulfate pentahydrate. (b) Dehydrated (anhydrous) copper (II) sulfate, $CuSO_4$, is white in colour. When water is added, it turns the characteristic blue colour of hydrated copper sulfate, $CuSO_4{\cdot}5H_2O$.

SAMPLE PROBLEM 5

A compound of sulfur contains 2.4% hydrogen, 39.0% sulfur and 58.6% oxygen. Find the empirical formula of the compound.

Teacher-led video: SP5 (tlvd-0525)

THINK

1. Follow the steps outlined at the start of section 5.4.1. Remember to assume that the *mass* of the sample is 100 g and all percentages become grams.
   TIP: Always organise the grid of elements in the order that they appear in the question and if you get a ratio of, say, exactly 2.5 to 3, double it so both are whole numbers; e.g. 2.5 to 3 becomes 5 to 6.

WRITE

Step 1	**Symbols**	**H**	**S**	**O**
Step 2	**Masses**	2.4 g	39.0 g	58.6 g
Step 3	**Moles**	$\frac{2.4}{1.0}$ = 2.4	$\frac{39.0}{32.1}$ = 1.2	$\frac{58.6}{16.0}$ = 3.7
Step 4	**Ratio/divide by smallest**	$\frac{2.4}{1.2}$ = 2.0	$\frac{1.2}{1.2}$ = 1.0	$\frac{3.7}{1.2}$ = 3.1

2. The empirical formula can be written directly from the ratios in the table — in this case, 2 hydrogen, 1 sulfur and 3 oxygen.

H_2SO_3

PRACTICE PROBLEM 5

Washing soda crystals may be used to bleach linen. When crystallised from water, washing soda (sodium carbonate, Na_2CO_3) forms crystals of a hydrated ionic compound.

When 5.00 g of washing soda crystals were dried in a desiccator, 1.85 g of sodium carbonate remained. Calculate the empirical formula of the hydrated compound.

Molecular formula

The **molecular formula** of a compound represents the actual composition of a compound that is made up of molecules. A molecular formula is either the same as its experimentally determined empirical formula, such as in ionic compounds, or is a whole-number multiple of it. For example, acetylene is a gas that is used in welder's torches, whereas benzene is a highly flammable liquid that is a recognised carcinogen. Although both compounds have the same empirical formula, CH, the molecular formula of acetylene is C_2H_2, and the molecular formula of benzene is C_6H_6. Determination of molecular formulas is discussed in topic 8.

on Resources

Digital documents Experiment 5.2 Stoichiometry of a reaction (doc-30872)
Experiment 5.3 The empirical formula of a hydrated salt (doc-30873)
Teacher-led video Experiment 5.3 The empirical formula of a hydrated salt (tlvd-0622)

5.4 EXERCISE

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1. Aspirin is a drug used extensively for pain relief. Chemical analysis of an aspirin tablet determined that it was composed of 57.7% carbon, 37.5% oxygen and 4.8% hydrogen. Calculate the empirical formula of aspirin.
2. Hydroquinone is a liquid that is used as a photographic developer. Determine the empirical formula of hydroquinone given that it contains 65.4% carbon and 29.1% oxygen with the remainder being hydrogen.
3. A 1.124 g sample of $CaSO_4$ crystals was heated to drive off the water of crystallisation. When completely dry, a residue of 0.889 g was obtained. Determine the empirical formula of the hydrated compound.
4. A 0.942 g sample of $MgSO_4$ crystals was heated to drive off the water of crystallisation. When completely dry, a residue of 0.461 g was obtained. Determine the empirical formula of the hydrated compound.
5. Borax is a naturally occurring compound that is used in the manufacture of optical glasses. Analysis of borax shows that it is made up of three elements: sodium (22.8%), boron (21.5%) and oxygen.
   Determine the empirical formula of borax.

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5.5 Review

5.5.1 Summary

Relative isotopic mass and the carbon-12 scale

- The relative atomic mass (A_r) of elements shown in the periodic table is calculated using the weighted mean of the relative isotopic masses:

$$A_r = \frac{\begin{array}{c}(\text{relative isotopic mass of first isotope} \times \text{abundance}) + \\ (\text{relative isotopic mass of second isotope} \times \text{abundance}) + \ldots\end{array}}{100}$$

  where A_r is the relative atomic mass.
- Relative isotopic masses are determined using a mass spectrometer and are masses compared with the carbon-12 isotope.
- The carbon-12 isotope is defined as having a mass of 12 g exactly. It is used as a standard because it is a stable and abundant isotope, which means that its mass can be determined accurately.
- The mass spectrometer is an instrument that can separate isotopes of an element based on their mass-to-charge ratio.
- A mass spectrum of an element shows the number of isotopes present in an element, their relative isotopic masses and their proportions.

The mole

- The mole concept is an important chemical idea, enabling us to ‘count’ particles present in a quantity of a substance.
  - The unit that measures the amount of a substance is the mole (mol).
  - A mole of any substance is composed of Avogadro’s number (6.02×10^{23}) of representative particles.
  - The molar mass (M) of an element or compound is the mass of 1 mole of the substance (numerically equal to the relative molecular mass (M_r)) and is expressed in g mol^{-1}.
  - The number of particles in a sample of substance may be determined by the formula:

$$\text{number of particles}\,(N) = \text{number of moles of substance}\,(m) \times \text{Avogadro's number}\,(N_A)$$

Determination of the empirical formula

- Formulas for compounds take two forms:
  - An empirical formula is the simplest whole-number ratio of atoms of the elements in the compound and is determined experimentally.
  - A molecular formula is the same as, or some simple multiple of, an empirical formula and specifies the exact number of each type of atom in a molecule of the compound.

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To access key concept summaries and practice exam questions download and print the **studyON: Revision and practice exam question booklet** (doc-30940).

5.5.2 Key terms

anhydrous without water in the chemical structure

Avogadro’s constant, N_A the number of elementary particles (atoms) in exactly 12 g of carbon-12. This number is equal to 6.02×10^{23}.

empirical formula formula that shows the simplest numerical ratio in which atoms are combined

hydrated containing water molecules within the structure

isotopes forms of an element with the same number of protons but different numbers of neutrons in the nucleus, resulting in elements with different masses

mass number total number of protons and neutrons in the nucleus of a particular isotope of an element

mass spectrometer analytical instrument that determines the relative isotopic masses of the different isotopes of an element and abundance

molar mass mass of one mole of a compound expressed in grams per mol

mole 6.02×10^{23} particles of a substance

molecular formula actual number of atoms of each element in a molecule of a substance

relative atomic mass, A_r mean mass of the naturally occurring mixture of the isotopes of an element on the relative atomic mass scale, on which the masses of particles are compared with the mass of the carbon-12 isotope

relative isotopic mass mass of an individual isotope of an element on the relative atomic mass scale, on which the masses of particles are compared with the mass of the carbon-12 isotope

unified atomic mass units; u standard unit of mass based on the mass of the C-12 isotope, where mass of an atom of C-12 equals 12 u; it is approximately equal to one nucleon (proton or neutron)

water of crystallisation water molecules included in a crystal lattice structure

on Resources

Digital document Key terms glossary — Topic 5 (doc-30938)

5.5.3 Practical investigations

Experiment 5.3

The empirical formula of a hydrated salt
Aim: To determine the empirical formula of a hydrated salt
Digital document: doc-30873
Teacher-led video: tlvd-0622

on Resources

Digital documents Practical investigation logbook — Topic 5 (doc-30939)
Experiment 5.1 Determining the empirical formula of magnesium oxide (doc-30846)
Experiment 5.2 Stoichiometry of a reaction (doc-30872)

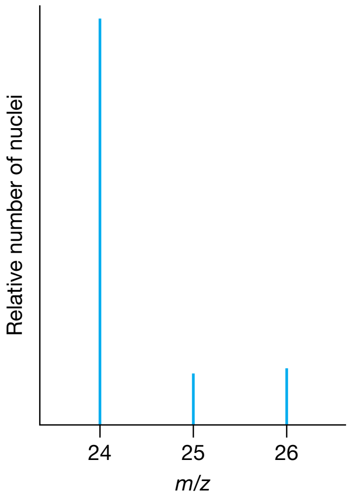
5.5 Exercises

To answer questions online and to receive **immediate feedback** and **sample responses** for every question go to your learnON title at www.jacplus.com.au.

5.5 Exercise 1: Multiple choice questions

1. What are the particles in a mass spectrometer deflected according to?
   - **A.** Their abundance
   - **B.** Their mass only
   - **C.** Their charge only
   - **D.** Their mass and charge
2. Rubidium has two stable isotopes. Rubidium-85 has an abundance of 72% and rubidium-87 has an abundance of 28%. The mass spectrum of rubidium has
   - **A.** 2 peaks: a smaller peak at relative isotopic mass 85 and a larger peak at relative isotopic mass 87.
   - **B.** 2 peaks: a smaller peak at relative isotopic mass 87 and a larger peak at relative isotopic mass 85.
   - **C.** 1 peak at about relative isotopic mass 85.5.
   - **D.** 1 peak at about relative isotopic mass 86.
3. When the element fluorine, F_2, is passed through the mass spectrometer, what particles are detected?
   A. F_2 ions **B.** F ions **C.** F^- and F^+ ions **D.** F^+ ions
4. A certain element has two isotopes of relative isotopic mass 203 and 205. Given that its A_r value is 204.4, which of the following is the most likely percentage abundance of the heavier isotope?
   A. 7% **B.** 30% **C.** 50% **D.** 70%
5. The difference between the relative molecular mass (M_r) of carbon dioxide and the molar mass of carbon dioxide is that the relative molecular mass and the molar mass of carbon dioxide are, respectively
   - **A.** 44.0 g and 1 mol.
   - **B.** 44.0 and 44.0 g mol^{-1}.
   - **C.** 44.0 g and 44.0 g mol^{-1}.
   - **D.** 44.0 and 44.0 g.

6. What is the number of moles of oxygen atoms in 143 g of sodium carbonate decahydrate, $Na_2CO_3 \cdot 10H_2O$?
   - **A.** 0.50
   - **B.** 1.5
   - **C.** 2.0
   - **D.** 6.5
7. What would the *smallest* number of molecules be contained in?
   - **A.** 1.0 g of N_2
   - **B.** 1.0 g of O_2
   - **C.** 1.0 g of NO
   - **D.** 1.0 g of NO_2
8. Where would the *largest* number of oxygen atoms be found in?
   - **A.** 300 g of water, H_2O
   - **B.** 3.2 mol of hydrated copper(II) sulfate, $CuSO_4 \cdot 5H_2O$
   - **C.** 3 kg of a fat having the molecular formula $C_{57}H_{110}O_6$
   - **D.** 7.35×10^{24} molecules of nitrogen dioxide, NO_2
9. Polymers may be made from the three monomers acrylic acid, $C_3H_4O_2$, methyl acrylite, $C_4H_6O_2$, and acrylonitrile, C_3H_3N. What is the order of increasing percentage of carbon content of these monomers?
   - **A.** Acrylic acid, methyl acrylite, acrylonitrile
   - **B.** Acrylic acid, acrylonitrile, methyl acrylite
   - **C.** Acrylonitrile, acrylic acid, methyl acrylite
   - **D.** Acrylonitrile, methyl acrylite, acrylic acid
10. Which of the following lists of substances shows only empirical formulas?
   - **A.** H_2O, CH_3COOH, HCl, $Na_2Cr_2O_7$
   - **B.** H_2SO_4, H_2O, $Al(NO_3)_2$, $Al_2(SO_4)_3$
   - **C.** HCl, $C_6H_{12}O_6$, NH_3, $Al(NO_3)_2$
   - **D.** $Al_2(SO_4)_3$, O_2, HCl, H_2SO_4
11. A substance, on analysis, was found to contain 27.3% carbon and 72.7% oxygen. Which of the following is the empirical formula of the substance?
   - **A.** CO_2
   - **B.** CO_3
   - **C.** C_2O_4
   - **D.** C_3O_7
12. The mass spectrum of magnesium is shown in the figure. Which of the following statements is true?
   - **A.** Three isotopes are shown with masses in the ratio 24:25:26 and relative abundances of about 8%, 1% and 1%.
   - **B.** The three isotopes have the same mass number but different atomic numbers.
   - **C.** Isotopes 25 and 26 have relative abundances of about 10%.
   - **D.** Isotope 24 has a relative abundance of about 10%.



5.5 Exercise 2: Short answer questions

1. Silicon-containing ores have three isotopes: 92% silicon-28, 5% silicon-29 and 3% silicon-30.
   a. Draw the mass spectrum for silicon.
   b. What is the atomic number of silicon?
   c. What are the relative isotopic masses of the three silicon isotopes?
   d. What is the relative atomic mass of silicon?
2. Silver is an unreactive metal that is used to produce jewellery. Its relative atomic mass is 107.9, and it consists of two naturally occurring isotopes: silver-107 and silver-109. Calculate the abundance of the lighter isotope.
3. In 2.0 moles of $(NH_4)_3PO_4$ (fertiliser) are
   a. ______ moles of nitrogen atoms.
   b. ______ moles of hydrogen atoms.
   c. ______ moles of oxygen atoms.
   d. ______ moles of phosphorus atoms.
   e. ______ (total number) nitrogen atoms.
   f. ______ grams of phosphate ions.
   g. ______ grams of nitrogen atoms.
4. Complete the following table.

Name	**Formula**	**Molar mass (*M*)**	**Mole (*n*)**	**Mass (*m*)**
Sodium hydroxide	NaOH			3.41 g
	CCl_4		1.40 mol	
Sodium carbonate		106 g mol^{-1}	1.00 mol	
	KCl		0.25 mol	
Ammonium phosphate				8.46 g

5. Find the number of moles *and* the number of atoms present in the following samples.
   a. 14.6 g acetylene (ethyne), C_2H_2
   b. 0.48 g propane, C_3H_8
   c. 485 g ethanol, C_2H_5OH
   d. 8.6 g carbon dioxide, CO_2
   e. 67 g iodine, I_2
6. Calculate the mass of each of the following.
   a. 160 mol Fe
   b. 0.075 mol silicon dioxide
   c. 4.23 mol NO_2
7. Sodium fluoride is thought to reduce tooth decay, especially in children. It is, therefore, added to some brands of toothpaste. If a tube of toothpaste contains 0.013 g of sodium fluoride, answer the following.
   a. How many moles of sodium fluoride does this represent?
   b. How many fluoride ions does this represent?
8. To prevent a gum disease called scurvy, the minimum daily requirement of vitamin C, $C_6H_8O_6$, is 60 mg.
   a. How many moles of vitamin C is this?
   b. How many molecules is this?
   c. If 10 g of spinach is found to contain 1.2×10^{-5} g of vitamin C, how much spinach must be eaten to attain the minimum daily requirement?

9. A typical vitamin C tablet contains 500.0 mg of ascorbic acid. (The molecular formula of vitamin C is $C_6H_8O_6$.)
   a. How many moles of vitamin C does a typical vitamin C tablet contain?
   b. How many molecules of vitamin C does a typical vitamin C tablet contain?
10. Too much cholesterol, $C_{27}H_{46}O$, is associated with heart disease, although the body produces its own cholesterol in the liver. A 250 g sample of cholesterol was removed from the arteries of a patient.
   a. How many moles of cholesterol are present in this sample?
   b. How many moles of carbon atoms does this represent?
   c. How many grams of carbon atoms are present in the sample?
   d. How many grams of oxygen atoms are present in the sample?
11. Diamond is a naturally occurring form of pure carbon. The mass of a diamond is measured in a unit called a carat, where 1.00 carat = 0.200 g. How many atoms of carbon are in the following?
   a. A 1.00-carat diamond
   b. A 3.15-carat diamond
12. Coffee contains the stimulant caffeine. Analysis shows it consists of 49.48% carbon, 28.85% nitrogen, 16.48% oxygen and 5.19% hydrogen, by mass. Calculate the empirical formula of caffeine.
13. The amino acid cysteine contains the elements carbon, hydrogen, nitrogen, oxygen and sulfur. Analysis of a 1.210 g sample of cysteine shows it to contain 0.0704 g of hydrogen, 1.80×10^{22} carbon atoms, 0.0100 mol of nitrogen and equal masses of sulfur and oxygen. Determine the empirical formula of cysteine.
14. Methanol is an alternative fuel to petrol that has been used in experimental cars. Determine the empirical formula of methanol, given that it is composed of 49.9% oxygen, 37.5% carbon and 12.6% hydrogen.

5.5 Exercise 3: Exam practice questions

Question 1 (6 marks)

Octane, C_8H_{18}, is a major component of petrol.

a. What is the M_r of octane? **1 mark**
b. What is the mass of 3.20 mol of octane? **1 mark**
c. How many molecules are in 2.5 g of octane? **1 mark**
d. How many individual atoms are in 5.0 g of octane? **1 mark**
e. What is the percentage of hydrogen in octane? **1 mark**
f. What mass of carbon would be present in 150 g of octane? **1 mark**

Question 2 (2 marks)

The odour of rancid butter is caused by butyric acid, which contains 54.5% carbon and 9.1% hydrogen, the rest being oxygen. Given that the M_r of butyric acid is 88.0, calculate the empirical formulas of the acid.

Question 3 (3 marks)

An oxide of copper is heated in a stream of hydrogen until only the copper remains, according to the equation:

$$Cu_xO\,(s) + H_2\,(g) \rightarrow x\text{Cu}\,(s) + H_2O\,(l)$$

The data for the experiment is given in the following table. Calculate the empirical formula of the oxide of copper.

Item	Mass (g)
Crucible	27.002
Crucible plus contents before heating	27.128
Crucible plus contents after heating	27.114

Question 4 (2 marks)
4.6 g of anhydrous zinc sulfate with an *M* of 161.5 g mol^{-1} was obtained by evaporating the water from 8.2 g of the crystalline hydrated salt with empirical formula $ZnSO_4 \cdot xH_2O$. Calculate the value of *x*.

Question 5 (2 marks)
Insects of a particular species can identify their mate by using special chemicals called pheromones that transmit chemical messages. The pheromone that serves as a sex attractant for gypsy moths is called disparlure and contains the elements C, H and O. Analysis of disparlure shows that 0.282 g contains 16.00×10^{-3} g of O atoms and 0.228 g of C atoms. The M_r of disparlure is 282 g mol^{-1}. Determine the empirical formula of disparlure.

Question 6 (6 marks)
Use examples to explain the difference between the terms 'mass number', 'relative atomic mass' and 'molar mass'.

5.5 Exercise 4: studyON Topic Test online only

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AREA OF STUDY 1 How can knowledge of elements explain the properties of matter?

OUTCOME 1

Relate the position of elements in the periodic table to their properties, investigate the structures and properties of metals and ionic compounds, and calculate mole quantities.

PRACTICE EXAMINATION

STRUCTURE OF PRACTICE EXAMINATION		
Section	**Number of questions**	**Number of marks**
A	20	20
B	8	30
	Total	**50**

Duration: 50 minutes

Information:

- This practice examination consists of two parts, you must answer all question sections.
- Pens, pencils, highlighters, erasers, rulers and a scientific calculator are permitted.
- You may use the VCE Chemistry Data Book for this task.

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Weblink VCE Chemistry Data Book

SECTION A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

1. Which of the following statements about the two particles ${}^{59}_{27}X$ and ${}^{59}_{28}Z^{+}$ is correct?
   A. They are a pair of isotopes of the same element.
   B. They have the same number of neutrons.
   C. Particle *Z* is significantly heavier than particle *X*.
   D. They contain the same number of electrons.
2. When the light emitted by hydrogen in the sun is analysed, the emission spectrum shows many distinct lines. Which one of the following statements about these distinct lines is true?
   A. Each line corresponds to a particular electron energy level in the hydrogen atom.
   B. Each line is produced as electrons gain the energy required to reach an excited state.
   C. Each line represents the energy released as electrons transition to lower energy levels.
   D. Each line represents a different element present in the sun.

3. Silver nanoparticles, considered to have antibacterial properties, are typically 10 nm in diameter. A red blood cell has an average diameter of 7 μm. How many silver nanoparticles, lined up side by side with no spaces between them, could fit across the diameter of the blood cell?
   A. 700
   B. 70
   C. 1.43
   D. 0.7
4. What is the maximum number of electrons that can occupy the 4*p* subshell of an atom?
   A. 4
   B. 6
   C. 8
   D. 10
5. What could the electron configuration $1s^22s^22p^63s^13p^6$ represent?
   A. A ground state atom of argon, Ar
   B. A ground state chloride ion, Cl^-
   C. An excited atom of chlorine, Cl
   D. An excited potassium ion, K^+
6. In which group of the periodic table will the element that has uncharged atoms with an electronic configuration $1s^22s^22p^63s^23p^64s^23d^3$ be located?
   A. Group 2
   B. Group 5
   C. Group 11
   D. Group 13
7. Which of the following statements about trends in the periodic table are true?
   I. Atomic radius increases going across a period in the periodic table.
   II. Both ionisation energy and non-metallic character increase across a period in the periodic table.
   III. Metallic character increases going down a group and decreases going across the periodic table.
   IV. Metallic character decreases going down a group and increases going across the periodic table.
   A. Only statements (II) and (III) are correct.
   B. Only statements (I) and (II) are correct.
   C. Only statements (I) and (III) are correct.
   D. Only statements (I) and (IV) are correct.
8. Where would an element with a relatively low first ionisation energy and a relatively high reactivity be found?
   A. *s* block
   B. *p* block
   C. *d* block
   D. *f* block
9. When element Q from group 16 of the periodic table reacts with element J from group 1, what compound is formed?
   A. Ionic with the empirical formula Q_2J
   B. A metal alloy with the empirical formula JQ
   C. Ionic with the empirical formula J_2Q
   D. Ionic with the empirical formula JQ
10. How do metallic atoms differ from non-metals?
   A. They have a smaller atomic radius than non-metals in the same period.
   B. They readily form ionic bonds when bonding with other metals.
   C. They have a smaller number of valence electrons.
   D. They have a higher effective nuclear charge.

11. Which of the following statements contrasting transition metals and the main group metals of groups 1, 2 and 13 is NOT true?

A. Transition metals typically form coloured compounds, whereas main group metals do not.

B. Transition metals are generally less reactive than main group metals.

C. Main group metals tend to have higher density than transition metals.

D. Main group metals form ions of a fixed charge, whereas transition metal ions can vary in their charge.

12. Consider the following qualities of metals *L, M,* and *Q*:

- Metal *L* tarnishes quickly in air and bubbles in cold water, producing hydrogen gas.
- The lustrous quality and inertness of metal *M* makes it typical for use in jewellery.
- Although metal *Q* does not react with water, it will react in dilute acid.

In order of reactivity from the most reactive to the least reactive, the metals are

A. $M > Q > L$.

B. $Q > L > M$.

C. $L > M > Q$.

D. $L > Q > M$.

13. Which of the following lists contains only ionic compounds?

A. CH_4, HCl, NaCl

B. H_2O, $CuSO_4{\cdot}5H_2O$, $MgSO_4{\cdot}7H_2O$

C. BaF_2, AlF_3, OF_2

D. $BaCl_2{\cdot}2H_2O$, FeS, $AgNO_3$

14. A student tested an ionic compound for electrical conductivity and found that it did not conduct as a solid, but did when dissolved in distilled water. What is the best reason to explain this?

A. When dissolved, the delocalised electrons are free to move.

B. Dissolving allows the conducting metal cations to separate from the non-conducting non-metal anions.

C. Cations and anions held in the solid ionic lattice can move freely once the lattice is dissolved.

D. The ionic compounds gain electrical conductivity from the surrounding water molecules.

15. Element *X* exists in only two stable isotopes, ^{72}X and ^{74}X. If the lighter isotope makes up 70% of all *X* atoms, what will the relative atomic mass of *X* be?

A. Less than 72

B. Greater than 74

C. Between 72 and 73

D. Between 73 and 74

16. What is the molar mass of $CuSO_4{\cdot}5H_2O$?

A. 159.6 g mol^{-1}

B. 249.6 g mol^{-1}

C. 177.6 g mol^{-1}

D. 127 g mol^{-1}

17. Magnetite is mineral found in iron ore. Its formula is Fe_3O_4. What is the percentage composition by mass of iron in magnetite?

A. 24.1%

B. 42.9%

C. 72.3%

D. 77.7%

18. One mole of iron(III) nitrate nonahydrate, $Fe(NO_3)_3{\cdot}9H_2O$ contains

A. 3 mol of O atoms.

B. 9 mol of O atoms.

C. 10 mol of O atoms.

D. 18 mol of O atoms.

19. A 27.77 g sample of a mystery element contains 0.1460 mol atoms of the element. What is the element most likely to be?

- **A.** Beryllium, Be
- **B.** Osmium, Os
- **C.** Niobium, Nb
- **D.** Rhenium, Re

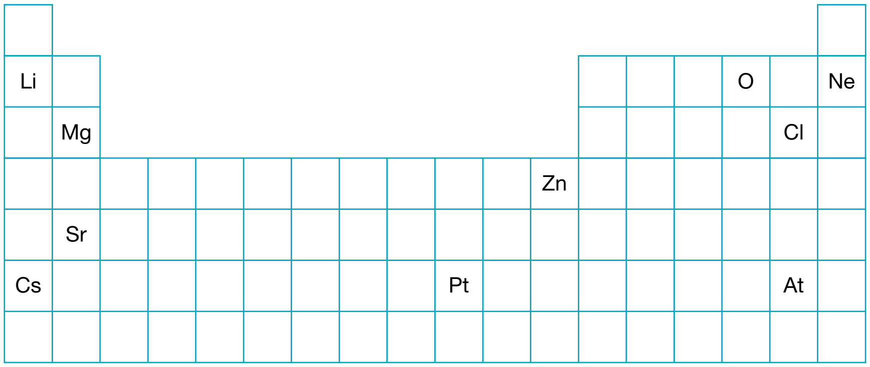
20. Which of the following contains the greatest number of hydrogen atoms?

- **A.** 10 g of hydrogen gas, H_2
- **B.** 6.02×10^{23} molecules of glucose, $C_6H_{12}O_6$
- **C.** 6.02×10^{24} methane, CH_4, molecules
- **D.** 5 mol of ammonia, NH_3

SECTION B — Short answer questions

Question 1 (5 marks)

Use only the elements marked in the following periodic table to match the descriptions in the questions. Each element can be selected only once.



a. What is the most electronegative element? **1 mark**

b. What is the halogen with the most metallic character? **1 mark**

c. Which element is filling its 5*s* subshell? **1 mark**

d. Which *d*-block element could form a substitutional alloy with copper? **1 mark**

e. Which element is closest in chemical properties to fluorine? **1 mark**

Question 2 (3 marks)

Write the subshell electronic configuration for the following.

a. Germanium, Ge **1 mark**

b. Phosphide ion, P^{3-} **1 mark**

c. Vanadium(II) ion, V^{2+} **1 mark**

Question 3 (3 marks)

Complete the following table by filling in the names or formulae for these ionic compounds.

Name	Formula
Potassium carbonate	
	$CaSO_4 \cdot 2H_2O$
	Cu_3PO_4

Question 4 (3 marks)

Electronegativity is a property that shows distinct trends in the periodic table.

a. Define what is meant by the term 'electronegativity'. **1 mark**

b. Electronegativity is found to increase going across a period in the periodic table. Explain this trend. **1 mark**

c. Elements with a large atomic radius tend to have lower electronegativities than smaller ones. Explain why this occurs. **1 mark**

Question 5 (5 marks)

Complete the following.

a. Draw a diagram representing a metal lattice, and describe the components of the lattice. **1 mark**

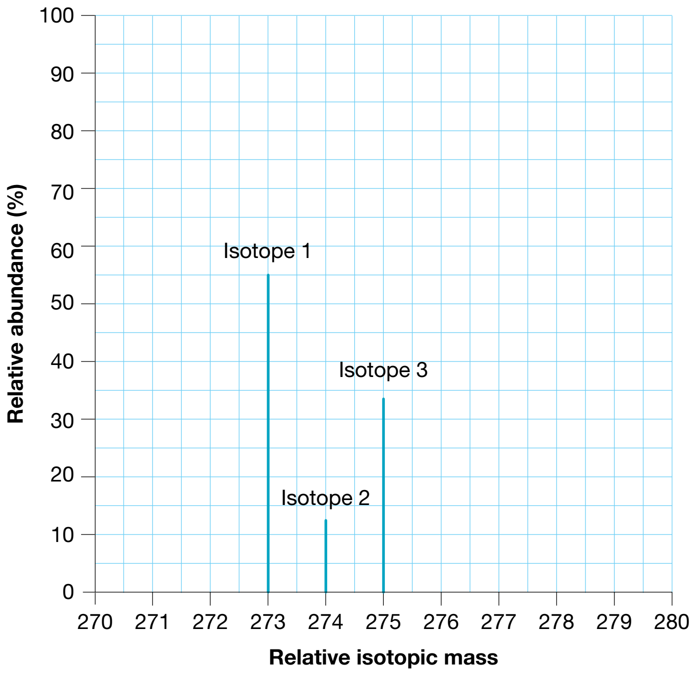
b. Explain why solid metals are malleable whereas ionic solids are not. **2 marks**

c. Explain why rapid cooling of molten (melted) metal produces a brittle substance. **1 mark**

d. Both metals and ionic compounds usually have high melting points. Explain why this is the case. **1 mark**

Question 6 (3 marks)

The mass spectrum of a rare element is shown in the following figure.



a. Use the data in the mass spectrum to complete the following table. **2 marks**

	Isotope 1	**Isotope 2**	**Isotope 3**
Relative isotopic mass			
Relative abundance (%)			

b. Using the data, calculate the relative atomic mass, A_r, of the element. **1 mark**

Question 7 (5 marks)

Ethanoic acid, CH_3COOH, is a naturally occurring compound found in vinegar.

a. In 20.6 grams of ethanoic acid, calculate

- **i.** the amount in moles of ethanoic acid. **1 mark**
- **ii.** the amount of carbon atoms in moles. **1 mark**
- **iii.** the number of hydrogen atoms. **2 marks**

b. What mass in mg of ethanoic acid would be present in 5.0×10^{-4} moles of the compound? **1 mark**

Question 8 (3 marks)

A student carried out an experiment to form an oxide of titanium. In the experiment, a sample of titanium metal was heated strongly in a crucible until no further visible change or change in mass was observed. The data obtained during the experiment is in the following table.

Mass of titanium before heating (g)	8.216
Mass of titanium oxide after heating and cooling (g)	13.692

Use the data to calculate the following.

a. The mass of oxygen added to the titanium metal during the experiment **1 mark**

b. The empirical formula of the titanium oxide **2 marks**

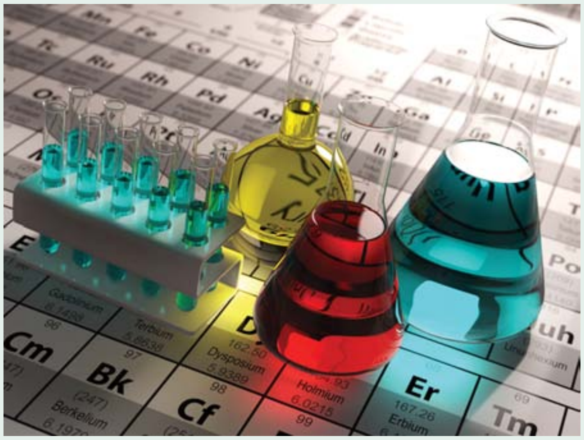
ASSESSMENT TASK — DATA ANALYSIS

In this task, you will focus on the trends of the periodic table and their relationship with the atomic model. You will be asked to predict periodic trends and graph these. Additionally, you will apply your practical knowledge to devise experimentation for placing unknown elements in the periodic table.

- This practice SAC comprises of three questions; you must complete ALL question sections.
- Pens, pencils, highlighters, erasers, rulers and a scientific calculator are permitted.
- You may use the VCE Chemistry Data Book for this task.

Total time: 55 minutes (5 minutes reading, 50 minutes writing)

Total marks: 50 marks



TRENDS IN THE PERIODIC TABLE

Question 1 (23 marks)

a. **(i)** Draw a graph of the number of electrons versus ionisation energy of the approximate values for the first six ionisation energies of aluminium. The first ionisation energy of aluminium is 584 kJ mol^{-1}. **5 marks**

(ii) Explain your reasoning in drawing your graph. **3 marks**

b. Complete the following table by writing the first three ionisation energy equations for aluminium and the resultant electronic configuration. **6 marks**

Ionisation energy	Equation	Electronic configuration
First		
Second		
Third		

c. Does Al^{2+} or Al^{3+} have the largest first ionisation energy? Use the Bohr model to explain your answer. **3 marks**

d. There is a decrease in period 3 ionisation energies from magnesium to aluminium of 736 to 577 kJ mol^{-1} respectively. This is contradictory to the general trend, where ionisation energy increases across a period. Explain why this decrease occurs, referring to the electronic structure of these elements. **2 marks**

e. What factors affect the size of the ionisation energy? **4 marks**

Question 2 (12 marks)

a. Draw, label and explain graphs to predict the general trend for the *second* period for the following.

(i) Electronegativity **2 marks**

(ii) Ionisation energy **2 marks**

(iii) Atomic radii **2 marks**

b. Draw, label and explain graphs to predict the general trend on *descending* group II for the following.

(i) Electronegativity **2 marks**

(ii) Ionisation energy **2 marks**

(iii) Atomic radii **2 marks**

Question 3 (15 marks)
You are given three unknown elements, X, Y and Z, two of which you are told are in the same group. You have been given access to all the equipment of the school laboratory, but do not have any internet access.

What experimentation would you do to predict where X, Y and Z should be placed in the periodic table and how would you determine which two elements are in the same group?

Write down all experimentations you would perform and why you have chosen these.

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Digital document U1AOS1 School-assessed coursework (doc-30962)

6 Materials from molecules

6.1 Overview

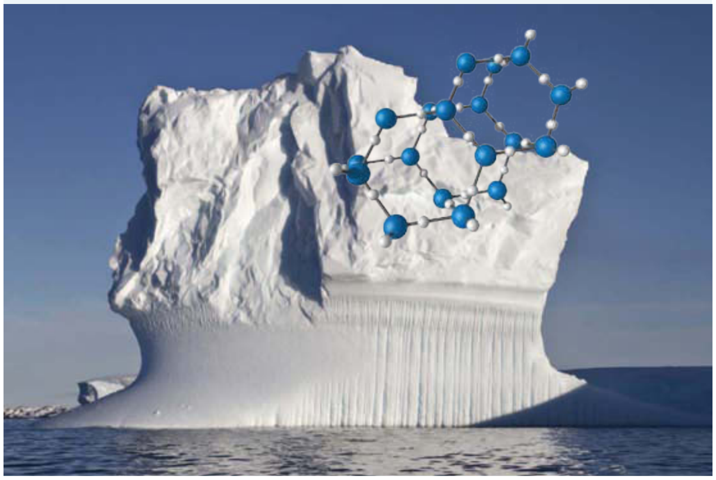
Numerous **videos** and **interactivities** are available just where you need them, at the point of learning, in your digital formats, learnON and eBookPLUS at www.jacplus.com.au.

6.1.1 Introduction

Water is present in vast amounts on the Earth. It is critical to life and plays an important part in moderating our climate. Within the normal temperature range of our planet, it exists in all three states: liquid, solid and as water vapour in the atmosphere.

Why is it that ice melts easily and other substances do not? Why does ice float on water? Why do ponds freeze over in winter, forming a layer of ice on their surfaces, instead of freezing solid? To answer these questions it is necessary to examine how hydrogen and oxygen atoms join together to make water molecules, and how water molecules then interact with each other.

FIGURE 6.1 Why does solid ice float on liquid water?



The properties of many other compounds made from non-metallic elements can be explained in the same way — by examining how the atoms join together to make the molecules, and then how these molecules interact with each other.

This topic explores how to predict the manner in which atoms of non-metals join together to make molecules and the methods that chemists use to represent the molecules that are formed. The important ways that molecules interact with one another will lead to an explanation for many of the observed properties of substances that bond this way — as covalent molecules.

These investigations build upon prior knowledge of electron sharing, chemical naming and formula writing. Of particular importance will be knowledge of electronic structure and noble gas configurations.

6.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- representations of molecular substances (electron dot formulas, structural formulas, valence structures, ball-and-stick models, space-filling models), including limitations of these representations
- shapes of molecules and an explanation of their polar or non-polar character with reference to the electronegativities of their atoms and electron-pair repulsion theory
- explanation of properties of molecular substances (including low melting point and boiling point, softness, and non-conduction of electricity), with reference to their structure, intramolecular bonding and intermolecular forces

- the relative strengths of bonds (covalent bonding, dispersion forces, dipole–dipole attraction and hydrogen bonding) and evidence and factors that determine bond strength, including explanations for the floating of ice and expansion of water at higher temperatures.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of learning and assessment. Experiments and investigations, supported by a **Practical investigation logbook** and **Teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

on Resources

Digital documents Key science skills (doc-30903)
Key terms glossary — Topic 6 (doc-30947)
Practical investigation logbook (doc-30948)

studyon

To access key concept summaries and practice exam questions download and print the **studyON: Revision and practice exam question booklet** (doc-30949).

6.2 Representing molecules

KEY CONCEPT

- Representations of molecular substances (electron dot formulas, structural formulas, valence structures, ball-and-stick models, space-filling models), including limitations of these representations

6.2.1 What is covalent bonding?

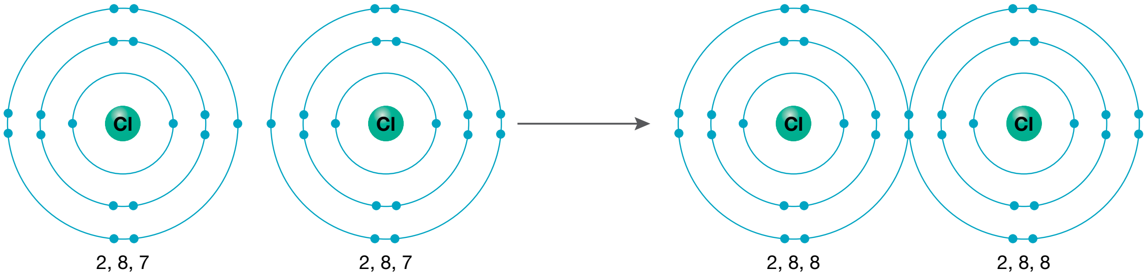
We have observed that atoms of elements with a stable outer shell (or an outer shell that contains eight electrons), such as the noble gases helium, neon and argon, are particularly stable. Other atoms tend to achieve noble gas configuration — that is, eight electrons in the outer shell — by losing or gaining electrons to form ionic bonds, or *sharing* electrons to form **covalent bonds**. The hydrogen atom needs two electrons in its outer shell for stability. Most of the elements in the second and third periods of the periodic table need eight outershell electrons for stability.

Non-metallic atoms generally have high **electronegativities**. This means they can attract electrons easily. Except for hydrogen, non-metals have four or more outershell electrons. When two non-metallic atoms react together, both of them need to gain electrons in order to obtain eight electrons in their outer shells. Since both cannot become negative ions, they share pairs of electrons to achieve eight electrons in the outer shell. This sharing of electron *pairs* between two non-metallic atoms produces the covalent bond. The atoms bond in order to reach a more stable state.

To better understand this process, we can use electron shell diagrams. Consider the two chlorine atoms shown in figure 6.2. Chlorine, as well as all members of group 17, exists in nature as **diatomic molecules** (two atoms covalently bonded together). Each chlorine atom has seven valence electrons. In order to attain a noble gas configuration (an outer shell of eight electrons), each chlorine atom needs one electron.

By sharing the single, unpaired electrons, both atoms can obtain eight in their outer shells, as shown in figure 6.2. A covalent bond results from the simultaneous attraction of the two positive chlorine nuclei to the same shared pair of electrons. This bond holds the two atoms together. A molecule of chlorine forms and is represented by the symbol Cl_2.

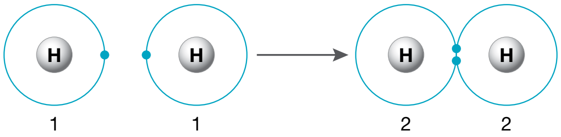
FIGURE 6.2 Two atoms of chlorine sharing a pair of electrons to create a covalent bond and the chlorine molecule Cl_2



Hydrogen — a special case

The hydrogen atom is an exception. This is because each of two hydrogen atoms contains one electron in its first (and outer) shell. They both need another electron to attain the stable electron configuration, 2, of the nearest noble gas, helium. This is achieved by each atom gaining a share in the electron of the other and forming a covalent bond, as shown in figure 6.3. In this way, the diatomic hydrogen molecule, H_2, is formed.

FIGURE 6.3 The formation of a hydrogen molecule



The molecule

A molecule is a group of non-metallic atoms held together by covalent bonds. The atoms are combined in a fixed ratio and are electrically neutral.

A **covalent molecular element** is made up of identical atoms held together in discrete groups by covalent bonds. Such elements may be found in various arrangements, including diatomic molecules. These molecules are described as **discrete**, because each molecule is separate and distinct from the others.

The noble gases, helium, neon, argon and krypton, all exist as discrete atoms (see figure 6.4). Since they have stable outer shells, they are unreactive and few of their compounds are known.

Diatomic molecules are composed of two atoms of the same element (see figure 6.5). For example, the covalent molecular element hydrogen exists as a gas made up of discrete diatomic molecules. As shown in figure 6.5, each molecule is composed of two hydrogen atoms. This is represented by the symbol H_2.

Hydrogen is one of several non-metallic elements that are made up of diatomic molecules. Other molecular elements include oxygen, O_2; nitrogen, N_2; phosphorus, P_4; fluorine, F_2; and sulfur, S_8.

In addition to these elements, a large number of compounds also exist as molecules. In a molecular compound, atoms of *different* elements share electrons with each other. These compounds are often called **covalent molecular compounds** because they contain covalent bonds. Water, carbon dioxide and methane are examples of these compounds (see figure 6.6).

FIGURE 6.4 Discrete atoms of helium gas

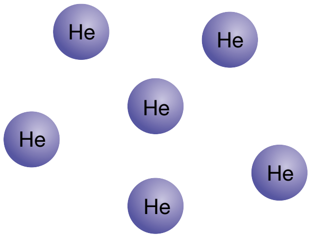


FIGURE 6.5 The molecules of hydrogen gas are diatomic.

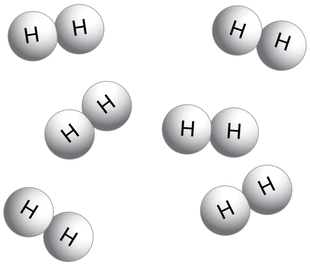
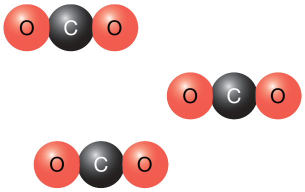


FIGURE 6.6 Carbon dioxide is a covalent molecular compound formed from atoms of carbon and oxygen.



A theory for the structure of covalent molecular substances

The major features of the current theory of the structure of covalent molecular substances are as follows:

- The basic units of covalent molecular substances are groups of atoms called **molecules**. All molecules within a pure substance are identical.
- Adjacent atoms within a molecule share electrons in order to achieve a stable outer shell.

- Electrical attraction between the nuclei of adjacent atoms and the shared electrons causes the atoms in a molecule to be held together. This force of attraction is called covalent bonding.
- The overall charge on each molecule is zero and so adjacent molecules are held together by intermolecular forces, which are weak in comparison to covalent bonds.

The octet rule

As we have already seen, the driving force behind all bonding is to obtain a more stable electron arrangement for the atoms involved. This arrangement in most cases is that of the nearest noble gas. Because all the noble gases have an outer shell of eight electrons, this tendency is often referred to as the **octet rule**. In other words, if an atom does not already have eight electrons in its outer shell, it will bond with other atoms in such a way as to achieve this.

The octet rule should not be confused with a 'full outer shell'. Many of the molecules that you will come across involve period two atoms and, because these atoms have outer shells with a maximum capacity of eight, effectively no difference exists between these terms. However, once period three or higher atoms are involved, outer shells can contain more than eight electrons. In cases such as these, it is inaccurate to state that they are bonding with other atoms *'to obtain a full outer shell'*. The alternative, that they bond to *'achieve eight in their outer shell'*, or they bond to *'achieve a noble gas configuration'*, is more accurate.

As with many generalisations such as this, exceptions always exist. The most common of these is hydrogen. Because hydrogen's outer shell is stable with only two electrons, the octet rule obviously cannot apply, but the 'stable outer shell' can. In other words, hydrogen atoms will always bond in such a way as to obtain *two* electrons in their outer shell.

on Resources

Weblinks Ionic and covalent bonding

Covalent bonding

6.2.2 Electron dot diagrams for atoms and molecular elements

In order to simplify drawings of atoms, **electron dot diagrams** (also called electron dot formulas or Lewis diagrams) may be used to represent the outershell electrons. These can be drawn for atoms and molecules and are a useful way to understand the sharing of electrons in covalent bonds.

Electron dot diagrams for atoms

In an electron dot diagram, the atom's nucleus and all innershell electrons are replaced by its element symbol. The outershell electrons are represented by dots or small crosses around the symbol in a square arrangement. These dots are arranged in pairs if more than four outer electrons are present. (*Note*: the innershell electrons do not participate in the bonding.). For example, nitrogen has five outershell electrons and can be represented by the following electron dot diagram. 

The two dots arranged as a pair represent **non-bonding electrons** or lone pairs. The three that appear as single dots (unpaired) are *available for sharing* and are called **bonding electrons**.

Elements of the same group in the periodic table have electron dot diagrams with the same number of dots or crosses. The halogens (group 17), for example, all have seven dots around their element symbols representing seven outershell electrons. Electrons in the outermost shell of an element are called **valence electrons**.

Electron dot diagrams for molecular elements

Earlier in this topic, figure 6.2 showed how, by sharing two electrons, two chlorine atoms can combine to each achieve a noble gas configuration (eight electrons in their outer shells). Electron dot diagrams can be used to represent this process in a simpler and clearer way. As can be seen in figure 6.7, dots (·) and crosses (x) are often used to represent the electrons that come from each atom. This does not mean that there are different types of electrons — all electrons are the same. These diagrams are just a convenient way of tracking the origins of the electrons.

In the case of chlorine, each atom shares a pair of electrons to form a single covalent bond. Three pairs of unused electrons remain on each chlorine atom, and these form lone pairs. The electron dot diagram for the formation of a chlorine molecule would, therefore, look as shown in figure 6.7.

The formation of a hydrogen molecule (refer to figure 6.3) may also be simplified using the electron dot method, as shown in figure 6.8. Notice that no unused electrons are present and, therefore, no lone pairs.

FIGURE 6.7 The formation of a chlorine molecule, shown using electron dot representation

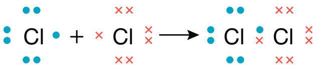


FIGURE 6.8 The formation of a hydrogen molecule, shown using the electron dot representation

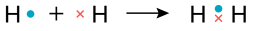


TABLE 6.1 Electron dot diagrams for non-metals in the second period of the periodic table

Group	Element	Valence electrons	Electron dot diagram	Lone pairs	Bonding electrons
14	C	4		0	4
15	N	5		1	3
16	O	6		2	2
17	F	7		3	1
18	Ne	8		4	0

Resources

Video eLesson Dot diagrams (eles-2474)

SAMPLE PROBLEM 1

Draw an electron dot diagram for a molecule of fluorine, F_2.

Teacher-led video: SP1(tlvd-0526)

THINK

Only outershell electrons are to be shown.

Fluorine is in group 17 and has 7 outershell electrons.

Draw the two fluorine atoms with three pairs and a single electron in a square pattern.

Draw atoms close together so that sharing of single electron is obvious.

TIP: Check that each atom has eight electrons.

WRITE

PRACTICE PROBLEM 1

Draw an electron dot diagram for a molecule of iodine, I_2.

Multiple covalent bonds

Many molecular substances are held together by **multiple bonds**; that is, bonds formed when two atoms share two or more pairs of electrons. If two atoms share two pairs of electrons, the covalent bond is called a **double bond**. Double bonds are found in molecules such as carbon dioxide, CO_2, and oxygen, O_2. When two atoms share three pairs of electrons, a **triple bond** is formed, as in the case of nitrogen, N_2.

SAMPLE PROBLEM 2

Draw an electron dot diagram for a nitrogen molecule, N_2.

Teacher-led video: SP2 (tlvd-0527)

THINK

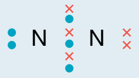
Nitrogen has five electrons in its outer shell. It needs three more. It will share three of its electrons with three from another N atom. This leaves two unused on each atom.

Draw the diagrams with three electrons on one side and two on the other.

Draw atoms close together so that the sharing of the single electron is obvious.

TIP: Check that each atom has eight electrons.

WRITE



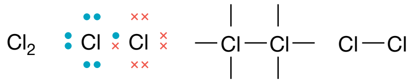
PRACTICE PROBLEM 2

Draw an electron dot diagram for an oxygen molecule, O_2.

6.2.3 Structural formulas and valence structures

Electron dot diagrams themselves may be further simplified by substituting a dash (–) for each pair of electrons. These dashes may represent bonding electron pairs (covalent bonds) or lone pairs (non-bonding electrons), depending on where they are located. Such a representation is called a **valence structure**. If the lone pairs are omitted and only bonds shown, the representation is called a **structural formula**. The structural formula thus shows the way the atoms in a molecule are connected. Although it is not necessary for a structural formula to indicate the shape of a molecule, it is possible to use them for this. This is discussed later in this topic.

FIGURE 6.9 Different ways of representing the chlorine molecule; from left to right: molecular formula, electron dot diagram, valence structure and structural formula



Electron dot diagrams, structural formulas and valence structures are, therefore, different ways that molecules may be represented 'on paper'. Often a combination of these is used. For example, structural formulas are sometimes used with important lone pairs shown by a dash or by dots. The method chosen will often depend on the context in which the diagram is to be used.

6.2.4 Electron dot diagrams for molecular compounds

The procedure for drawing electron dot diagrams for molecular elements is easily extended to molecular compounds. However, these can quickly become complicated as molecular size and the number of different elements increases. For this reason, the electron dot method tends to be used for smaller and simpler molecules. To draw the electron dot formula for a molecular compound, the following steps are often useful:

1. Place the atom with the most bonding electrons (called the **central atom**) in the centre so that the other atoms can be arranged around it.

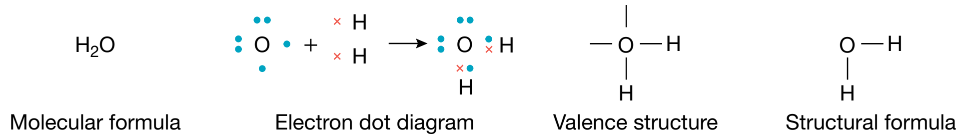
2. Draw dot diagrams for each of the atoms in the molecule.
   Remember that lone pairs do not participate in the bonding.
3. Attach all atoms in such a way that each atom (except hydrogen) is surrounded by eight electrons.
   Remember that if you have trouble achieving eight in an outer shell, you might have to consider multiple bonding.

Water

The formula for water is H_2O. In each water molecule, an oxygen atom shares electrons with two hydrogen atoms so that they all achieve complete outer shells.

Notice that the central oxygen atom in a water molecule has two pairs of unbonded electrons. These are called lone pairs. Although they do not participate in the covalent bond, they affect the shape of the molecule, and are shown in the valence structure by the two 'unbonded' dashes.

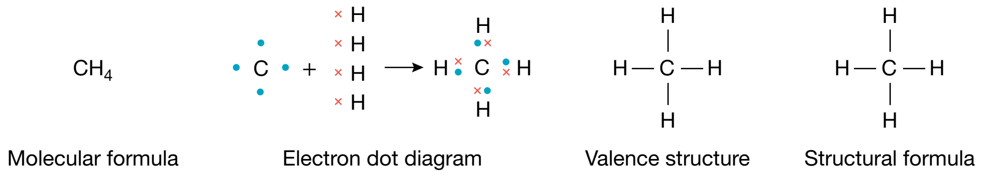
FIGURE 6.10 Different ways of representing the water molecule



Methane

The formula for methane is CH_4. In each methane molecule, carbon forms a single bond with each of the four hydrogen atoms. All atoms have achieved complete outer shells.

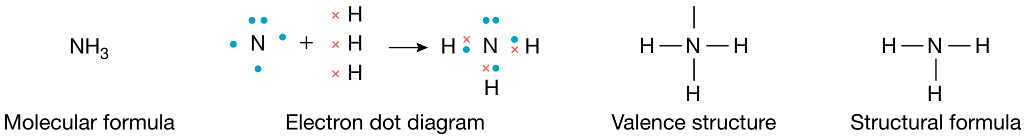
FIGURE 6.11 Different ways of representing the methane molecule



Ammonia

The formula for ammonia is NH_3. In each ammonia molecule, nitrogen forms a single bond with each of the three hydrogen atoms in addition to having one lone pair. This arrangement achieves a stable outer shell for each atom.

FIGURE 6.12 Different ways of representing the ammonia molecule

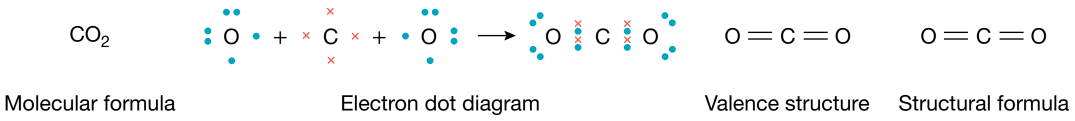


Carbon dioxide

The formula for carbon dioxide is CO_2. In each carbon dioxide molecule, carbon forms a double bond with each of the oxygen atoms.

The central carbon atom in a carbon dioxide molecule now has eight valence electrons, and so does each oxygen atom.

FIGURE 6.13 Different ways of representing the carbon dioxide molecule

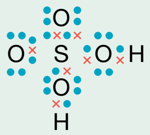


EXCEPTIONS TO THE OCTET RULE

The octet rule is a very useful *rule of thumb* for chemists when working out bonding within molecules and molecular shape.

However, many molecules do not obey this rule. In some instances, molecules bond with fewer than eight electrons in their outer shell, and in many cases have more. Obviously, the latter cannot occur for period two elements due to the capacity of the second shell. However, for periods three and beyond, the outer shell has room for more than eight electrons. Although many compounds involving period three and beyond still obey the octet rule, some don't. Some of these are quite common; for example, sulfuric acid, the structure of which is shown in figure 6.14.

FIGURE 6.14 Electron dot diagram of sulfuric acid showing 12 electrons on the sulfur atom



If you count the electrons in the outer shell of the sulfur atom, you will notice that there are 12!

Some other molecules that also exhibit this feature are PCl_5, SF_6, IF_5, SO_3, H_3PO_4 and $HClO_4$.

Look up the structures for these molecules (or even try to work them out first)! Also look for any other examples, and whether you can find some examples of molecules that have fewer than eight outershell electrons around their central atom.

This exercise is a good illustration that many of the tools and techniques used by chemists have their limitations. The bonding models discussed in this topic are a further example of this.

This does not mean that such ideas aren't any good. The main thing to remember is that they serve a purpose and, so long as their limitations are recognised, no reason exists to not keep using them.

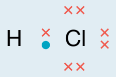
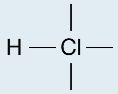
SAMPLE PROBLEM 3

For hydrogen chloride, draw the following.

a. Electron dot diagram

b. Valence structure

Teacher-led video: SP3 (tlvd-0528)

THINK	WRITE
a. Hydrogen has one electron in its shell, and so needs one more. The chlorine atom has seven electrons in its outer shell, so it needs one more. Draw the diagrams together, showing overlap for sharing. **TIP:** Check that hydrogen has two electrons in its outer shell and eight electrons for chlorine.	
b. Draw valence structure by replacing each pair of electrons with a dash.	

PRACTICE PROBLEM 3

For hydrogen fluoride, draw the following.

a. Electron dot diagram

b. Valence structure

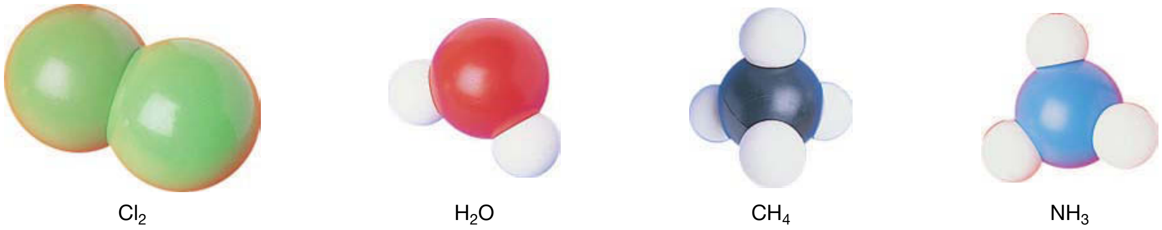
6.2.5 Physical representations

The representations discussed so far have one very obvious disadvantage: they are two-dimensional. Physical models allow us to visualise molecules in three dimensions. As we shall see later in this topic, the three-dimensional shape of a molecule can have a significant effect on its properties. As introduced in topic 4 with reference to ionic compounds, two main types of three-dimensional model are used by chemists to represent molecules. These are **ball-and-stick** models and **space-filling** models. These can be used to inform **computer-generated** models.

Space-filling models

Space-filling models use spheres (usually made of plastic) to represent atoms. These are colour coded to represent different types of atoms. Typical examples are black (carbon), white (hydrogen), red (oxygen), blue (nitrogen) and green (chlorine). The hydrogen spheres are usually smaller than the others to represent the much smaller relative size of the hydrogen atom. These are attached to each other so as to resemble the shape of the molecule as closely as possible.

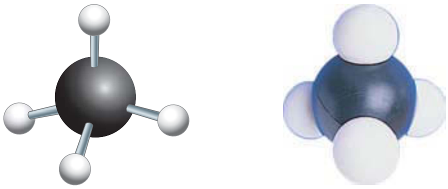
FIGURE 6.15 Space-filling models for (left to right) chlorine, Cl_2, water, H_2O, methane, CH_4, and ammonia, NH_3



Ball-and-stick models

Ball-and-stick models are similar to space-filling models except that the spheres are separated and joined by sticks. Each stick represents a covalent bond.

FIGURE 6.16 Methane, CH_4, represented by a ball-and-stick model (left) and a space-filling model (right)



Ball-and-stick models have the advantage that they can show multiple bonds (see table 6.2).

6.2.6 Limitations of these representations

Why are there different ways to represent molecules? The answer is that chemists use different ways for different purposes. Each representation has its own strengths and limitations. The method chosen will often depend on the context in which it is to be used. If the bonding between atoms is being emphasised, a ball-and-stick model would be better than a space-filling one. While both electron dot and valence structures will also highlight bonding between atoms, the former may lose clarity among a maze of dots representing the electrons. If shape is important, both ball-and-stick and space-filling may be adequate, with both being superior to electron dot and valence structures. In many situations, especially in organic chemistry, a structural formula is an excellent way to represent molecules.

Table 6.2 lists some of the limitations of each of these methods. You may be able to think of more.

TABLE 6.2 Methods of molecular representation with some of their limitations

Method	Limitations
Electron dot diagrams (also called Lewis diagrams)	• Two-dimensional • Does not indicate shape • Tedious and slow to produce • Can look messy
Valence structure	• Two dimensional • Does not necessarily indicate shape* • Can be tedious and slow to write for larger molecules • Important lone pairs may be difficult to distinguish from non-important ones
Structural formula	• Two dimensional • Does not necessarily indicate shape* • Can be tedious and slow to write for larger molecules
Space-filling	• Does not show the bonding between atoms • Relative sizes of atoms involved are only approximate
Ball-and-stick	• Can be physically difficult to produce for larger molecules • Kits to produce are usually expensive

* These methods can be adapted to give an indication of shape by using shape diagrams (see subtopic 6.3).

6.2 EXERCISE

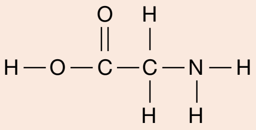
To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

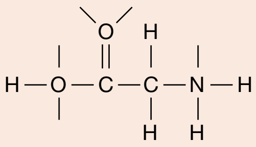
1. Draw an electron dot diagram for an atom of the following.
   (a) Tellurium (Te)
   (b) Arsenic (As)
2. Complete the following table.

Group	Element	Valence electrons	Electron dot diagram	Lone pairs	Bonding electrons
	S				
	Si				
	P				
	Cl				
	Br				
	Se				
	Ar				

3. Which of the following substances contain discrete diatomic molecules?
   (a) Magnesium oxide
   (b) Iodine
   (c) Aluminium
   (d) Neon
   (e) Phosphorus
   (f) Carbon dioxide
4. Draw an electron dot formula for a molecule of bromine, Br_2.
5. Although it is not a preferred form of phosphorus, the synthesis of diphosphorus, P_2, has been reported. Draw the electron dot formula for a molecule of diphosphorus.

6. Draw the electron dot diagram and the valence structure for hydrogen iodide, HI.
7. The structural formula and valence structure of the amino acid glycine is shown.
   (a) Draw this structure as an electron dot diagram.
   (b) Give one disadvantage of showing this molecule as an electron dot diagram.





8. Use electron dot diagrams and structural formulas to illustrate the bonds between each of the following pairs of atoms.
   (a) Hydrogen and bromine
   (b) Oxygen and fluorine
9. Draw the electron dot diagram and the structural formula for nitrogen trifluoride, NF_3, in which all three fluorine atoms are bonded to the nitrogen atom.
10. Draw electron dot diagrams and structural formulas for the following molecules.
    (a) CS_2
    (b) C_2F_4
    (c) C_3H_8
    (d) HCN

studyon

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6.3 Predicting molecular shape

KEY CONCEPT

- Shapes of molecules and an explanation of their polar or non-polar character with reference to the electronegativities of their atoms and electron-pair repulsion theory

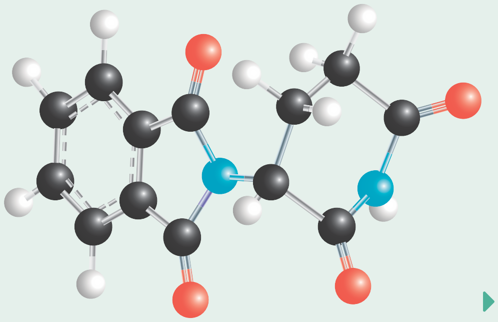
6.3.1 The valence shell electron pair repulsion (VSEPR) theory

The shape of a molecule can have an important effect on its properties and how it behaves. For this reason, it is important to know not only how atoms are joined together to make a molecule, but also in what orientation they are with respect to one another. This is especially important for biological molecules and drugs where a relatively small change in shape can have dramatic differences in effect.

THALIDOMIDE

Thalidomide is a drug that was first produced in the 1950s. It became popular as an over-the-counter drug sold to combat morning sickness in pregnant women. However, it soon became apparent that it was producing a devastating side effect: children were being born with severely deformed limbs. Over 10 000 children were born before the drug was withdrawn. Today the use of the drug is strictly controlled, and prescribed for treating leprosy and certain types of cancers.

Thalidomide has the chemical formula $C_{13}H_{10}N_2O_4$, shown in the figure.

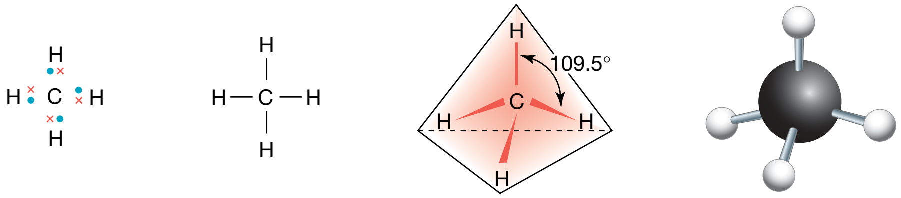


You notice the usual colour convention in the figure: hydrogen atoms are white, oxygen atoms are red, nitrogen atoms are blue and carbon atoms are black.

Thalidomide exists in two forms that are nearly identical. The only difference is a small change in orientation around one of the carbons. One form *did* cure morning sickness. However, the other form caused the birth defects. To make matters worse, the body is able to convert the harmless form into the harmful one, making it almost impossible to prevent the drug's serious side effects.

All discrete molecules have a definite three-dimensional shape. Electron dot diagrams and structural formulas fail to represent the three-dimensional shapes of molecules. For example, the electron dot diagram and structural formula of methane, CH_4, show the molecule in only two dimensions. In reality, methane molecules exist in three dimensions. The hydrogen atoms are arranged in a tetrahedral shape around the central carbon atom, as shown by the ball-and-stick diagram included in figure 6.17.

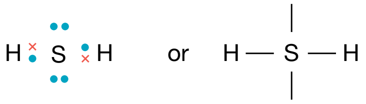
FIGURE 6.17 The electron dot diagram and structural formula of methane do not show its three-dimensional structure, which is tetrahedral. This structure is better represented using a ball-and-stick model.



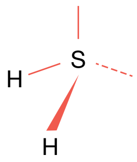
The **valence shell electron pair repulsion (VSEPR) theory** provides a relatively simple basis for understanding and predicting molecular geometry. The theory requires only that the number of outershell electrons for each atom in the molecule be known. The electron pairs in the molecule repel each other and take up positions as far from one another as possible. When determining the shapes of molecules, the electron pairs of a multiple bond count as only one pair of electrons for prediction purposes. Lone pairs should also be considered, especially when they are on a central atom. Examples of the five most common molecular shapes are given in table 6.3. Note that, in the shape diagrams, a dotted line indicates that the bond is directed into the plane of the paper and the wedge indicates that the bond is directed out of the plane of the paper. The solid lines represent bonds that are in the plane of the paper. The use of ball-and-stick-modelling is an excellent way to understand this theory.

For example, consider hydrogen sulfide, H_2S, and silane, SiH_4.

To predict the shape of H_2S, first draw either its electron dot diagram or valence structure.



Next, count all the electron pairs (both bonding and non-bonding) around the central atom (in this case, sulfur) and arrange them as far from each other as possible. For four electron pairs in three dimensions (this orientation is a tetrahedral one). The shape diagram may now be drawn.



A similar procedure yields the shape of the SiH_4 molecule. Once again there are four pairs of electrons around the central atom and they therefore adopt a tetrahedral orientation.

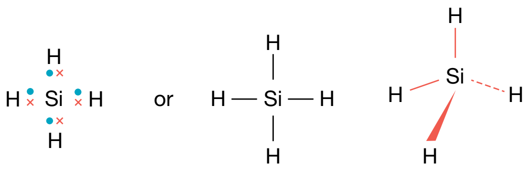
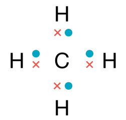
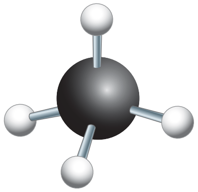
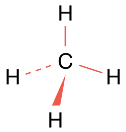
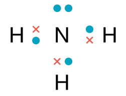
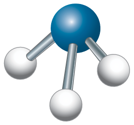
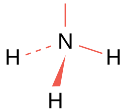
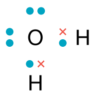
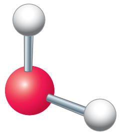
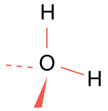
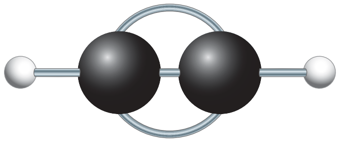
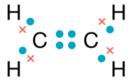
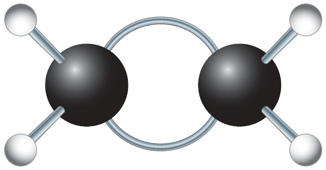
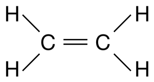


TABLE 6.3 Shapes of some common molecules

Compound	Electron dot diagram	Number of lone pairs around central atom	Number of bonding electron groups (pairs) around central atom	Shape	Ball-and-stick model	Shape diagram
Methane $CH_4(g)$		0	4	Tetrahedral (around C atom)		
Ammonia $NH_3(g)$		1	3	Pyramidal (around N atom)		
Water $H_2O(l)$		2	2	V-shaped (around O atom)		
Ethyne $C_2H_2(g)$		0 for each C atom	2 for each C atom	Linear (around each C atom)		$H-C\equiv C-H$
Ethene $C_2H_4(g)$		0 for each C atom	3 for each C atom	Planar (around each C atom)		

on Resources

Video eLessons Molecular shapes (eles-2475)

VSEPR — Sulfur dioxide (eles-2488)

TIP: To recognise the shape of a molecule:

- **determine the arrangement of electron pairs around the central atom (for example, four pairs will be tetrahedral) and ignore the lone pairs**
- **remember that the distribution of the atoms will help you describe the shape.**

SAMPLE PROBLEM 4

Using its electron dot diagram or its valence structure, draw the shape diagram for a molecule of H_2Se.

Teacher-led video: SP4 (tlvd-0529)

THINK

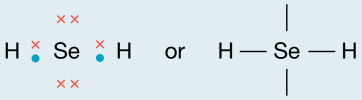
1. Draw the electron dot diagram. H must only show one electron. Se is in group 16 so it must show six electrons in the outer shell.

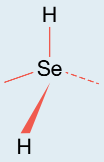
   Arrange H atoms around the Se with an overlap to show sharing. (Remember that electron sharing may be clearer using a valence structure.)

2. Draw the shape diagram. Count the pairs of electrons around central atom to identify four pairs. Two lone pairs of electrons and two pairs of bonding electrons around the central atom correspond to a V-shaped arrangement. Assign H atoms to the bonding electrons.

   TIP: VSEPR theory places electrons as far from each other as possible.

WRITE





PRACTICE PROBLEM 4

Using its electron dot diagram or its valence structure, draw the shape diagram for a molecule of PH_3.

SAMPLE PROBLEM 5

Using its electron dot diagram or its valence structure, draw the shape diagram for a molecule of C_2H_4.

Teacher-led video: SP5 (tlvd-0530)

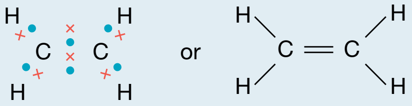
THINK

1. H atoms are always arranged around the outside of the central atom, so the two C atoms must be in the centre.

   Draw each H showing one electron and each C showing four electrons. Arrange H atoms around the C with an overlap to show sharing.

   TIP: If sharing is difficult to show with an electron dot diagram, try forming multiple bonds. Electron sharing in this molecule works with a double bond!

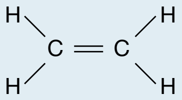
WRITE



2. Draw the molecule showing four electrons between each C as a double bond.

Three areas of electrons are around each C (double bonds count as one) so the shape is planar around each carbon, with bonds 120° apart.

This planar structure makes the whole molecule flat.



PRACTICE PROBLEM 5

Using its electron dot diagram or its valence structure, draw the shape diagram for a molecule of C_2H_2.

The shape of a molecule is not necessarily the same as the distribution of electron pairs around the central atom. The shape considers the three-dimensional orientation of the atoms in space.

VSEPR theory predicts the shape of molecules.

- **A molecule will assume the shape that minimises electron pair repulsions.**
- **Electron pairs around the central atom will be as far apart as possible.**

In general:

- **molecules with no lone pairs around the central atom are tetrahedral, unless a double or triple bond creates a linear or planar structure (depending upon the distribution of atoms around the central atom)**
- **molecules with a single lone pair around the central atom are pyramidal**
- **molecules with a two lone pairs around the central atom are V-shaped.**

on Resources

Digital document Experiment 6.1 Building molecular models (doc-30848)

Teacher-led video Experiment 6.1 Building molecular models (tlvd-0623)

6.3.2 Polar and non-polar molecules

All bonding involves the attractions set up between positive charges and negative charges. It is electrical in nature and electrons play a critical role. Even though molecules are electrically neutral in an overall sense, the possibility exists of regions within a molecule where small localised imbalances may occur. **Polarity** is the term that chemists use to describe this situation. A slight deficiency of electrons in a region (and, therefore, a slight positive charge) is indicated by the symbol $\delta+$ (taken from the lowercase Greek letter for delta). Likewise, a slight excess of electrons is indicated by $\delta-$.

Electronegativity

Covalently bonded atoms usually exhibit unequal attractions for shared electrons. Scientists have found that different atoms have different electron-attracting abilities. The relative attraction that an atom has for shared electrons in a covalent bond is known as its electronegativity. A scale of electronegativities was developed by Linus Pauling in which the most electronegative atom, fluorine, is assigned a value of 4.0.

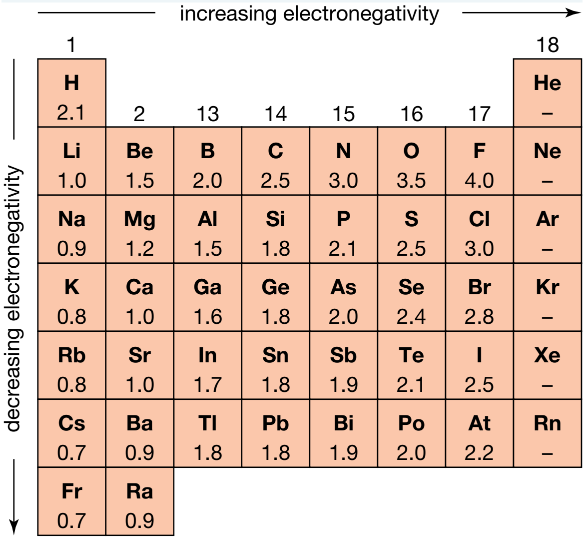
Fluorine attracts electrons almost twice as well as hydrogen, which is given an electronegativity value of 2.1. No values are assigned for the noble gases. Electronegativity can be likened to a ‘tug of war’, where the two teams are like the atoms on each end of the bond trying to pull the electrons towards them.

The Pauling scale of electronegativities

The following electronegativity trends may be seen in the periodic table.

- Electronegativities increase from left to right within a period.
- Electronegativities decrease from top to bottom within a group.
- Metals generally have lower electronegativities than non-metals.

FIGURE 6.18 The Pauling scale of electronegativities shows electronegativity trends in the periodic table.



1	2	13	14	15	16	17	18
H 2.1							**He** –
Li 1.0	**Be** 1.5	**B** 2.0	**C** 2.5	**N** 3.0	**O** 3.5	**F** 4.0	**Ne** –
Na 0.9	**Mg** 1.2	**Al** 1.5	**Si** 1.8	**P** 2.1	**S** 2.5	**Cl** 3.0	**Ar** –
K 0.8	**Ca** 1.0	**Ga** 1.6	**Ge** 1.8	**As** 2.0	**Se** 2.4	**Br** 2.8	**Kr** –
Rb 0.8	**Sr** 1.0	**In** 1.7	**Sn** 1.8	**Sb** 1.9	**Te** 2.1	**I** 2.5	**Xe** –
Cs 0.7	**Ba** 0.9	**Tl** 1.8	**Pb** 1.8	**Bi** 1.9	**Po** 2.0	**At** 2.2	**Rn** –
Fr 0.7	**Ra** 0.9						

LINUS PAULING: CHEMIST AND PEACE ACTIVIST

Linus Carl Pauling (1901–1994) was an American scientist famous for his work on chemical bonding and biochemistry, and his activism for peace. Voted by *New Scientist* as one of the 20th greatest scientists of all time, he was also regarded as one of the two greatest scientists of the 20th century (the other being Einstein). He is widely regarded as the greatest chemist since the founder of the discipline, Antoine Lavoisier, and was the recipient of two Nobel Prizes.

Until the 1930s, chemistry had only been able to describe the properties of substances; it hadn’t been able to explain them. Pauling applied the newly developed quantum theory to describe the bonding between atoms and how this explained a substance’s properties. In 1931, he published the first of seven papers on the subject; a paper now regarded as one of the most significant works in chemistry. In 1939, following further research, Pauling collated his work to up to that date in the landmark book *The nature of the chemical bond*. This book is still used by chemists across the world today.

During the mid-1930s, Pauling became interested in the field of biomolecules and began to apply his knowledge to the structure of the large molecules that exist in living things. He proposed that molecules such as proteins and DNA have a helical structure and laid the groundwork for the ultimate discovery of DNA’s structure by Watson, Crick and Franklin in 1953. He was one of the pioneers of model building. He used data from X-ray diffraction to build scale models of the molecules he was studying.

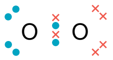
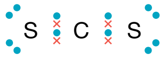
During World War II, Pauling worked for the United States Government on a number of projects associated with the war. In 1948, he was awarded the Presidential Medal for Merit for his efforts. However, his unease in the postwar period led to him to speak out with several other scientists (including Einstein) against the proliferation of nuclear weapons. He became a tireless campaigner against the testing of nuclear weapons.

Pauling was the winner of two Nobel Prizes. The first, for chemistry in 1953, was for his work on the nature of the chemical bond and his second, for peace in 1962, was for his campaigning against the testing of nuclear weapons. He is the only person to have been awarded two unshared Nobel Prizes, and one of only two people to have won them in different fields. Pauling was a popular personality and was as widely known in the general population as he was in the scientific community. Chemistry students around the world recognise him as the developer of the scale of electronegativites that bears his name.

Non-polar and polar covalent bonds

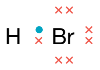
Covalent bonds in which the bonding electron pair is shared equally and is uniformly distributed between the nuclei of two bonded atoms are called **non-polar covalent bonds**. Such bonds can result only when two atoms of the same element or two atoms of equal electronegativity simultaneously attract a shared pair of electrons.

TABLE 6.4 Non-polar covalent bonding in molecules

Molecule	Electron dot diagram	Structural formula
Hydrogen		$H-H$
Chlorine		$Cl-Cl$
Oxygen		$O=O$
Carbon disulfide		$S=C=S$

Covalent bonds in which the bonding electrons are unequally shared, and therefore unsymmetrically distributed between the nuclei of two bonded atoms, are called **polar covalent bonds**. Such bonds occur between atoms of different electronegativities. The shared pair of electrons moves closer to the more electronegative atom. This means that the atom that has greater control of the electron pair becomes slightly negatively charged ($\delta-$), while the atom that lost some control of the electron pair becomes slightly positively charged ($\delta+$). Polar covalent bonds can be said to have a charge separation, or a **bond dipole**.

TABLE 6.5 Polar covalent bonding in molecules

Molecule	Electron dot diagram	Structural formula and polarity
Hydrogen chloride		$\overset{\delta+}{H}-\overset{\delta-}{Cl}$
Hydrogen bromide		$\overset{\delta+}{H}-\overset{\delta-}{Br}$

A polar covalent bond is not purely covalent, since there is not an *equal* sharing of electrons. It has some characteristics of ionic bonding, although the transfer of electrons from one atom to the other is not complete because it is in a purely ionic bond. In fact, if the difference in electronegativity between two atoms is two or greater on the Pauling scale, an ionic bond does form. The relationship between electronegativity and bond type is shown in table 6.6.

TABLE 6.6 Relationship between electronegativity and bond type

Difference in electronegativity between bonding atoms	Type of bond formed
Zero; e.g. H (2.1) and H (2.1)	Non-polar covalent
Medium; e.g. H (2.1) and Cl (3.0)	Polar covalent
Large; e.g. Na (0.9) and F (4.0)	Ionic

SAMPLE PROBLEM 6

State whether the following bonds are polar or non-polar.

a. Oxygen to oxygen bonds
b. Hydrogen to phosphorus bonds
c. Oxygen to chlorine bonds

Teacher-led video: SP6 (tlvd-0531)

THINK	WRITE
A bond will be polar if a difference in electronegativity exists between the two atoms on each end.	
a. No difference exists for O to O.	Oxygen to oxygen bonds are non-polar.
b. A difference does exist for H to P.	Hydrogen to phosphorus bonds are polar.
c. A difference does exist for O to Cl.	Oxygen to chlorine bonds are polar.

PRACTICE PROBLEM 6

State whether the following bonds are polar or non-polar.

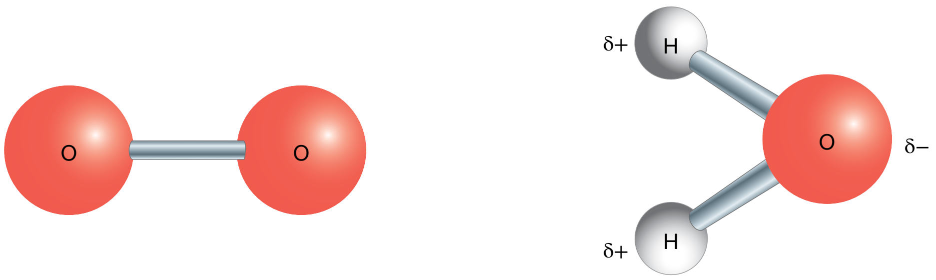
a. Chlorine to chlorine bonds
b. Phosphorus to chlorine bonds

Non-polar and polar molecules

Molecules containing only non-polar bonds are **non-polar molecules**, such as O_2.

In a **polar molecule**, one end of the molecule is slightly negative and one end is slightly positive. Such a molecule is sometimes called a dipolar molecule, or molecular dipole. All polar molecules contain polar bonds. However, some molecules with polar bonds are non-polar. This occurs when the individual dipoles are arranged in such a way that they cancel each other out. So having polar bonds in a molecule does not necessarily mean that the molecule is polar overall. The polarity or non-polarity of the molecule depends, rather, on the *direction* of the bond dipoles in the molecule, which is determined by the molecule's *shape*. Figure 6.19 shows examples of both a non-polar and polar molecule.

FIGURE 6.19 In the non-polar oxygen molecule, O_2, the oxygen atoms have equal electronegativity and share electrons equally (left). Water molecules are polar because the oxygen atom has greater electronegativity than the hydrogen atoms, so it becomes the more negatively charged end of the molecular dipole (right).



Determining the polarity of a molecule

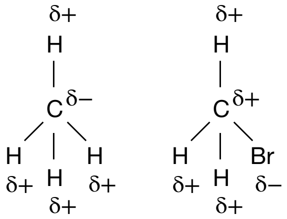
To determine the polarity of a molecule the following steps should be followed:

1. Draw an electron dot diagram of the molecule.
2. Apply the VSEPR rules to draw a shape diagram of the molecule.
3. Use electronegativities to determine bond dipoles.
4. Use the shape diagram along with bond dipoles to determine whether the molecule is polar or non-polar. If bond dipoles cancel each other out, the molecule is non-polar. If bond dipoles do not cancel each other out, one side of the molecule attracts electrons more than another side and the molecule is polar.

The following guidelines are often useful.

- Linear, planar and tetrahedral molecules with equal polar bonds that cancel each other out are non-polar. If the bond dipoles are not equal or do not cancel each other out, the molecule is polar. For example, CH_4 is non-polar whereas CH_3Br is polar because the C—Br bond dipole is not cancelled by the C—H bond dipoles.
- V-shaped or pyramidal molecules are polar because their polar bonds do not cancel.

FIGURE 6.20 CH_4 is non-polar because the bond dipoles cancel each other out. CH_3Br, however, is polar.



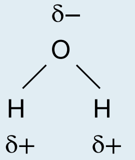
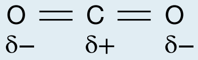
SAMPLE PROBLEM 7

Predict the polarity of the following molecules.

a. Water

b. Carbon dioxide

Teacher-led video: SP7 (tlvd-0532)

THINK	WRITE
Recall that a molecule being polar requires two things: it must have polar bonds and its shape must be such that the dipoles from these bonds do not cancel out (that is, it must be asymmetrical). Therefore, check electronegativities first and then draw the shape diagrams.	
a. Water, H_2O, is a V-shaped molecule due to the effect of the two lone pairs on O. The two O—H bond dipoles are equal but since they are not at an angle of 180° to one another (that is, aligned in opposite directions) they do not cancel each other out. A water molecule is, therefore, polar.	 Water is polar.
b. Carbon dioxide, CO_2, is a linear molecule with two polar bonds. However, the bond polarities cancel each other out because they are in opposite directions. Electrons are not attracted preferentially to one side or the other of the molecule. Carbon dioxide is, therefore, a non-polar molecule.	 $\underset{\delta-}{O}=\underset{\delta+}{C}=\underset{\delta-}{O}$ Carbon dioxide is non-polar.

PRACTICE PROBLEM 7

Predict the polarity of the following molecules.

a. Iodine

b. Hydrogen bromide

6.3 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Using its electron dot diagram or its valence structure, draw the shape diagram for a molecule of AsH_3.
2. State whether the following bonds are polar or non-polar. Justify your answer.
   (a) Sulfur to oxygen bonds
   (b) Bromine to bromine bonds
3. Explain whether the following molecules are polar or non-polar.
   (a) Chlorine (Cl_2)
   (b) Ammonia
4. Use the electronegativity trends in the periodic table (shown in figure 6.18 in section 6.3.2) to predict whether the following bonds are polar or non-polar. Show the bond dipoles where they are present using $\delta+$ and $\delta-$.
   (a) H—F
   (b) O—H
   (c) C—H
   (d) N—H
   (e) C—C
5. Classify the following bonds as ionic, polar covalent or non-polar covalent. Justify your answers.
   (a) HI
   (b) KCl
   (c) F_2
6. Arrange the following bonds in order from non-polar covalent to ionic. F—N, N—N, Cl—N.
7. Explain why compounds with a large electronegativity difference between their atoms are likely to show ionic bonding, whereas atoms with only a small difference will display covalent bonding.
8. A molecule has a tetrahedral shape. Does this mean that it is non-polar? Explain.
9. Sketch the shapes of the following molecules and use your knowledge of electronegativity to draw bond dipoles for each to predict whether it is polar or non-polar.
   (a) CH_4
   (b) CH_3Cl
   (c) CH_2Cl_2
   (d) $CHCl_3$
   (e) CCl_4
10. ICl_3 and ICl_5 are two compounds that form between iodine and chlorine and do not obey the octet rule. These two compounds have ten (ICl_3) and twelve (ICl_5) electrons in the outer shells of the central atom.
   (a) Draw electron dot diagrams and valence structures for each of these compounds.
   (b) Why is it possible to have more than eight electrons in the outer shell for iodine?

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6.4 Explaining the properties of molecular substances

KEY CONCEPT

- Explanation of properties of molecular substances (including low melting point and boiling point, softness, and non-conduction of electricity), with reference to their structure, intramolecular bonding and intermolecular forces

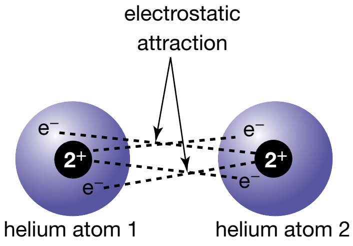
6.4.1 Inter-molecular attractions

The existence of molecular crystals indicates that molecules can be held together in an orderly array. How do these molecules 'stick' together?

In addition to covalent bonds *within* molecules, called **intramolecular** bonding, attractions exist *between* molecules, called **intermolecular** forces of attraction, which hold molecules to each other. These attractive forces are weaker than either covalent or ionic bonds, but they determine whether a molecular compound exists in the solid, liquid or gaseous states. The temperature at which a molecular substance melts or boils, therefore, depends on the strength of the intermolecular forces. The stronger the intermolecular forces, the higher the melting or boiling point.

Three types of intermolecular forces are possible. In some molecular substances, more than one force may be operating between the molecules. The weakest attractions between molecules are **dispersion forces** followed by **dipole–dipole interactions**. The third and strongest intermolecular force is called **hydrogen bonding**.

FIGURE 6.21 Below ~4 K (−269 °C) He is a liquid as atoms are very weakly attracted to each other by the electrostatic attraction of one nucleus to the electrons of other atoms.



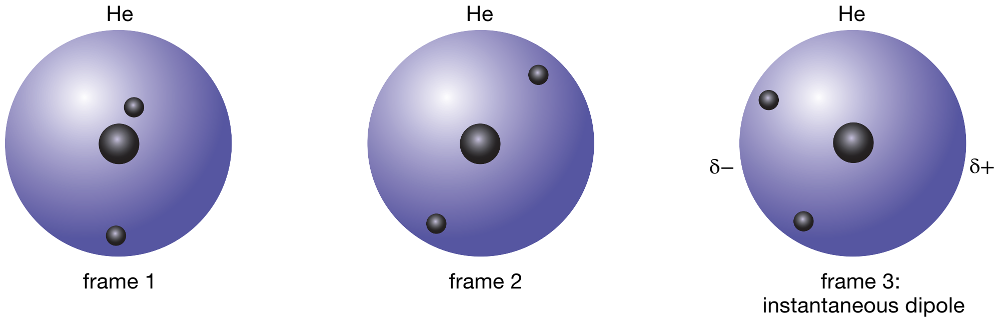
Dispersion forces

In a sample of molecules, the nuclei of atoms in one molecule are able to attract the electrons of atoms in neighbouring molecules (in addition to attracting their own electrons in forming a covalent bond). All electrons are attracted by all neighbouring nuclei.

Dispersion forces are the weakest of the intermolecular forces. To be condensed to a liquid, the noble gas helium must be cooled to –269 °C so that the atoms are travelling slowly enough for the dispersion forces to have a better chance of pulling these two atoms together. The helium nucleus in one atom attracts the electrons in a neighbouring atom. If cooled to –272 °C at pressures of 25 atm, liquid helium solidifies.

Dispersion forces can also occur between instantaneous dipoles. Electrons are moving constantly. At any point in time, a chance exists that the electrons may be found on one side of an atom or molecule, and it can develop an 'instantaneous dipole' (frame 3 in figure 6.22). This temporary dipole can cause a shift in the distribution of electrons in neighbouring atoms or molecules, resulting in 'induced dipoles' (figure 6.23). Very small forces of attraction result between the particles.

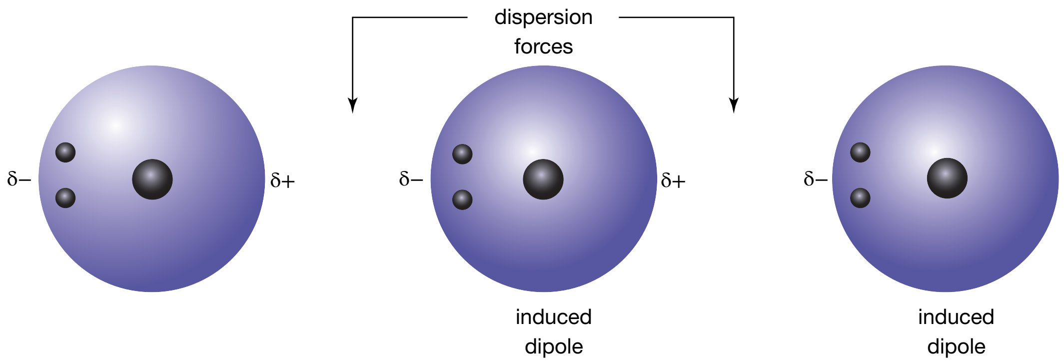
FIGURE 6.22 The electrons in this helium atom are moving all the time; an instantaneous dipole is created.



Two factors influence the strength of dispersion forces:

1. The number of electrons in the molecules. In general, the more electrons the molecules of a substance have, the stronger the dispersion forces between them.
2. The shapes of the molecules. Shape affects how closely the molecules may approach each other in the solid and liquid states. The closer the molecules can get, the stronger the attraction is. All covalent molecular substances have dispersion forces between their molecules.

FIGURE 6.23 The atom on the left has caused a temporary induced dipole in the middle atom, which in turn has caused another temporary induced dipole in the atom on the right.



Consider the noble gases He and Ne, or small molecules such as F_2, H_2 and CH_4. These atoms or molecules have relatively few electrons and so have very weak dispersion forces. Such substances have low boiling points and exist as gases at room temperature. Larger compounds, such as octane, C_8H_{18} (a component of petrol), exist as liquids. However, candle wax, $C_{25}H_{52}$, is a solid with a low melting point.

This trend can be illustrated using the halogens, as shown in table 6.7. As we go down the group, the molecules become larger and a progression occurs from a gaseous to liquid and ultimately solid state. Boiling and melting points of halogens show a corresponding increase.

TABLE 6.7 Properties of group 17 elements

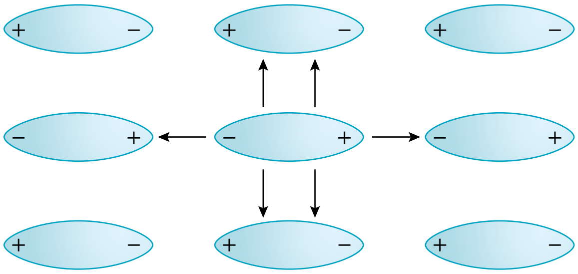
Name	Formula	State	Melting point (°C)	Boiling point (°C)
Fluorine	F_2	Gas	–220.3	–188
Chlorine	Cl_2	Gas	–101	–34
Bromine	Br_2	Liquid	–7.3	59
Iodine	I_2	Solid	113	184

Dipole–dipole interactions

If the molecules in a sample are polar, the presence of molecular dipoles causes simultaneous intermolecular attraction. The positive side of one molecule attracts the negative side of another molecule, which attracts the next, and so on to the limits of the sample.

Polar molecules have two forces of attraction operating between their molecules. If two molecules are of similar size — that is, have a similar number of electrons — the dispersion forces acting on them are similar. However, if one of the molecules is polar, it will also be affected by dipole–dipole attractions. This results in a stronger overall intermolecular force and a higher boiling point. Consider Ar (boiling point −186 °C) and HCl (boiling point −83.7 °C). Both substances have 18 electrons, yet HCl has a much higher boiling point. This is because HCl is a polar molecule and so has an extra intermolecular force of dipole–dipole interactions. Such a compound is generally more likely to exist as a liquid or solid at room temperature than a non-polar compound of similar size.

FIGURE 6.24 In dipole–dipole interactions, the central polar molecule is attracted to other polar molecules around it. They, in turn, are attracted by their neighbours.



Hydrogen bonding

Hydrogen bonding is a special case of dipole–dipole attraction. Hydrogen is an atom containing one proton in the nucleus and one electron revolving around it.

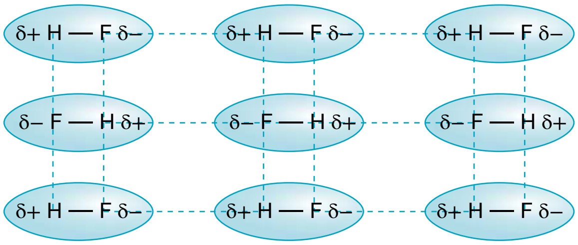
When hydrogen bonds to a more electronegative atom such as nitrogen, oxygen or fluorine, its electrons move slightly toward that atom. This causes the hydrogen nucleus to be exposed or unshielded. The molecule that forms is a dipole.

The negative end of one of these dipoles is attracted to the positive end of another.

Because the hydrogen (positive end) is unshielded, the other dipole can approach far more closely. The closer the dipoles get, the stronger the bond that forms. The bond between the dipoles is called a hydrogen bond. These bonds are represented in figure 6.25 by dotted lines, indicating that they are a weaker bond type than a covalent bond. Hydrogen bonds form only *between* molecules and only when hydrogen has been bonded to fluorine, oxygen or nitrogen.

Hydrogen bonds are stronger than other dipole–dipole bonds and result in materials with higher melting and boiling points than would otherwise be expected.

FIGURE 6.25 The negative ends of the dipoles are attracted to the positive ends, causing a hydrogen bond to form.



6.4.2 Properties of molecular substances in relation to structure

Although the covalent bonding holding the atoms together in a molecule is strong, the forces between molecules are usually weak. The physical properties of a compound depend on the type of bonding it displays. Ionic compounds are crystalline solids, but molecular substances may exist as solids, liquids or gases. A great variety of physical properties occurs among molecular substances due largely to the differences in strength of the intermolecular attractions. Even so, a few generalisations can be made.

- Molecular substances do not conduct electricity in the solid or molten form because the molecules are electrically neutral.
- If dissolved in water without reacting with it, molecular substances do not conduct electricity, again because the molecules are electrically neutral.
- Some molecules dissolve in water to produce ions. These molecules are said to ionise when they are dissolved in water, and can conduct electricity owing to the movement of the ions produced.
- Molecular substances vary in their solubility in water and other solvents. Generally, polar compounds are soluble in polar solvents such as water but insoluble in non-polar solvents such as tetrachloromethane. Non-polar solutes tend to be more soluble in non-polar solvents than polar solvents. This is sometimes described as the ‘like-dissolves-like’ rule.
- Molecular compounds have low melting and boiling points since the forces between the molecules are weak, and relatively little energy is required to break them. Many molecular substances are gases or liquids at room temperature.
- Most molecular substances are soft and easily scratched. Once again this is due to the weak forces of attraction between the molecules that mean molecules can be easily removed from the surface.

FIGURE 6.26 The softness of wax, together with its insolubility in water, make it ideal to apply to surfboards.

Note that in all of these cases (except for the third point above), the molecules themselves remain unchanged. Although they may be separated from other molecules by processes such as boiling, dissolving and scratching, the molecules themselves remain intact. This is because the *intra*molecular force of covalent bonding is very strong.

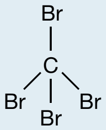
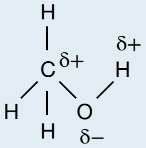
SAMPLE PROBLEM 8

Name the type(s) of intermolecular forces that exist between molecules in the following substances.

a. **CBr_4**

b. **CH_3OH**

Teacher-led video: SP8 (tlvd-0533)

THINK	WRITE
a. CBr_4 is a non-polar tetrahedral molecule.  	Only dispersion forces exist between these molecules.
b. CH_3OH has a tetrahedral shape around the carbon atom and a V-shape around the second central atom, oxygen, resulting in a polar molecule. It also has a highly electronegative atom (oxygen) bonded to a hydrogen atom.  	Dispersion forces and hydrogen bonding exist between these molecules.

PRACTICE PROBLEM 8

Explain the type(s) of intermolecular forces that exist between molecules in the following substances.

a. **CI_4**

b. **CH_3NH_2**

on Resources

Weblink Intermolecular forces

6.4 EXERCISE

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1. Explain the type(s) of intermolecular forces that exist between molecules in the following substances.
   - (a) CCl_4
   - (b) $CH_3CH_2NH_2$
   - (c) Cl_2
   - (d) CH_3OH
   - (e) CH_3Cl

2. Explain why water, H_2O, has a higher boiling point than hydrogen sulphide, H_2S, despite it being a smaller molecule with fewer electrons.
3. The boiling points of HCl, HBr and HI are, respectively, – 85 °C, – 67 °C and – 35 °C. Explain the difference in these boiling points with reference to the forces that exist between the molecules of each substance.
4. Br_2 and ICl have the same number of electrons, yet Br_2 melts at –7.2 °C, and ICl melts at 27.2 °C. Explain this in terms of intermolecular forces.
5. N_2 has 14 electrons and Cl_2 has 34 electrons. Predict which of these substances would have the higher boiling point. Justify your answer in terms of intermolecular forces.
6. Two noble gases, helium and argon, have boiling points of –269 °C and –186 °C respectively. Explain the large difference in their boiling points.
7. Kr (boiling point –152 °C) and HBr (boiling point –67 °C) have the same number of electrons. Explain what factors could affect intermolecular forces to cause the difference in boiling points between Kr and HBr.
8. HCl has more electrons than HF so we would expect it to have the higher boiling point. However, this is not the case; the boiling points of HCl and HF are –83.7 °C and 19.4 °C respectively. Explain what factors could account for this reversal in trend.
9. Explain why F_2, O_2 and N_2 are all gases at room temperature.
10. Predict whether I_2 dissolves more readily in non-polar tetrachloromethane or in polar water. Explain your answer.
11. Would you expect candle wax, $C_{25}H_{52}$, to be
    (a) soft or hard?
    (b) soluble in water?
    (c) an electrical conductor?
    Explain your predictions using your knowledge of structure and bonding.
12. Only certain substances display hydrogen bonding.
    (a) Is hydrogen bonding an intramolecular force or an intermolecular force?
    (b) The causes of hydrogen bonding are within a molecule but its effect is felt between molecules. Explain this statement.
    (c) State the requirements for hydrogen bonding.

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6.5 The relative strengths of bond types in molecular substances

KEY CONCEPT

- The relative strengths of bonds (covalent bonding, dispersion forces, dipole–dipole attraction and hydrogen bonding) and evidence and factors that determine bond strength, including explanations for the floating of ice and expansion of water at higher temperatures

6.5.1 Factors that determine bond strength

Table 6.8 lists some properties of the hydrogen halides (compounds between hydrogen and an element from group 17). The patterns and trends within this table illustrate some of the factors that determine bond strengths.

TABLE 6.8 Some physical properties of the hydrogen halides

Hydride	Formula	Melting temperature (°C)	Boiling temperature (°C)	Bond dissociation energy* (kJ mol^{-1})	Bond length (picometres) (pm)
Hydrogen fluoride	HF	+19.5	–83	568	92
Hydrogen chloride	HCl	–85	–115	432	127
Hydrogen bromide	HBr	–67	–87	366	141
Hydrogen iodide	HI	–35	–51	298	161

* Bond dissociation energy measures how difficult it is to break the bonds within a molecule and separate it into its component atoms. The higher the value, the stronger the bonds.

When we examine the bond dissociation energies, a trend appears. Going down the table (which corresponds to going down group 17), the values decrease, meaning the covalent bond between the hydrogen and the halide atom is becoming weaker. This is because the electrons involved from the halide atom are one electron shell further out each time. This makes the length of the bond longer (as shown in the last column in table 6.8). The bond is weaker because the two electrons in it are more spread out and also further away from the nucleus. This trend is obvious in nearly all other molecules — the longer the length of the bond, the weaker it is.

Another trend is that double bonds are stronger than single bonds, and triple bonds are stronger yet again. Table 6.9 illustrates this for carbon to carbon bonds.

TABLE 6.9 Bond lengths and strengths for carbon-carbon bonds

Bond type	Bond dissociation energy (kJ mol^{-1})	Bond length (pm)
C—C	348	154
C=C	614	134
C≡C	839	120

Returning to table 6.8, trends are also apparent in melting and boiling temperatures, although not quite as perfect. Descending down a group, both these increase, except for hydrogen fluoride. This means that the forces between the molecules are getting stronger. This is because the molecules are getting larger and contain more electrons, increasing the strength of the dispersion forces that hold the molecules together.

But why is hydrogen fluoride an exception? It is because hydrogen fluoride has an additional type of bonding between its molecules: hydrogen bonding. This makes the forces between hydrogen fluoride molecules stronger and the molecules harder to separate from each other when melting and boiling take place. The influence of hydrogen bonding also occurs in water, and it is explored below and in greater detail in topic 11.

SAMPLE PROBLEM 9

Consider the following hydrides from group 16: H_2O, H_2S, H_2Se.

a. **Which will have the strongest bonding within its molecules?**
b. **Which will have the strongest dispersion forces between its molecules?**
c. **Which will have the overall strongest bonding between its molecules?**
d. **Which will have the highest boiling point?**

Teacher-led video: SP9 (tlvd-0534)

THINK	WRITE
a. Recall that bond length gets longer down a group, and that the longer the bond, the weaker it is. H_2O has the shortest bonds and, therefore, the strongest intramolecular bonding.	H_2O has the strongest bonding within its molecules.
b. Recall that dispersion forces depend on the number of electrons in the molecule — the more electrons, the stronger the dispersion forces. H_2Se is the largest molecule and has the greatest number of electrons. It will, therefore, have the strongest dispersion forces between its molecules.	H_2Se has the strongest dispersion forces between its molecules.
c. Recognise that hydrogen bonding exists in H_2O but in none of the others. This is stronger than the dispersion forces that are (always) present. H_2O, therefore, has the strongest overall intermolecular bonding.	H_2O has the strongest overall bonding between its molecules.
d. Recall that the stronger the forces between molecules, the higher the boiling point. H_2O, therefore, has the highest boiling point.	H_2O has the highest boiling point.

PRACTICE PROBLEM 9

Consider the following hydrides from group 17: HF, HCl, HBr, HI.

a. Which will have the strongest bonding within its molecules?
b. Which will have the strongest dispersion forces between its molecules?
c. Which will have the overall strongest bonding between its molecules?
d. Which will have the highest boiling point?

Relative strengths of bond types

Molecular substances may experience up to four different types of bonding. *All* molecules are held together by covalent bonding. Because this acts *within* a molecule, it is classified as an *intra*molecular force. It is a very strong force and considerably greater than any of the other forces that may be experienced. *All* molecules also experience attraction to other molecules from dispersion forces. These are the very weak forces that originate due to the movement of electrons within the molecule. Other forces that may be experienced *between* molecules are dipole–dipole attractions and hydrogen bonding. These two are due to the polarity that exists within *some* molecules. All forces *between* molecules may be classified as *inter*molecular forces.

The decreasing order of strength between the four types of bonding is as follows, starting with strongest.

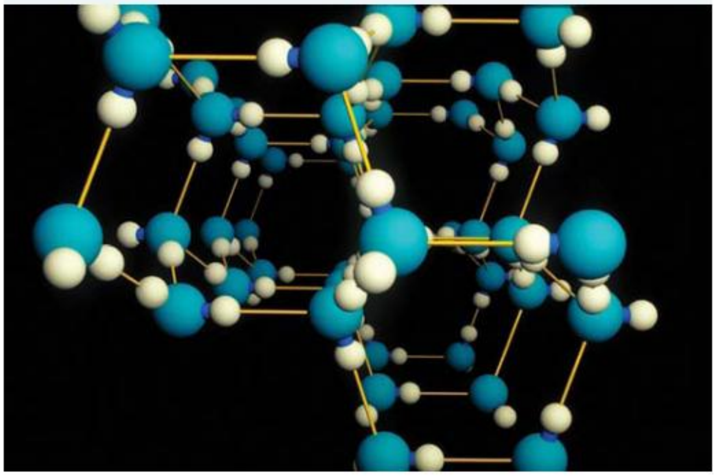
Covalent bonding >>>>> hydrogen bonding >> dipole–dipole attraction > dispersion forces

6.5.2 The effect of hydrogen bonding on the properties of water

Hydrogen bonding significantly affects the physical properties of water. The existence of hydrogen bonding between water molecules results in:

- the relatively high melting and boiling points of water compared with other substances. More heat is required to enable the molecules to gain sufficient kinetic energy to break free of the hydrogen bonds, which are stronger than dispersion forces alone. For example, H_2O is a liquid at room temperature with a boiling point of 100 °C, whereas H_2S, which has more electrons, is a gas at room temperature and boils at –61 °C. The behaviour of H_2O as it changes state between liquid water and solid ice is also governed by the hydrogen bond.

FIGURE 6.27 The hexagonal structure of ice. The molecules in liquid water are hydrogen bonded but free to move about. In ice, molecules are held further apart and in a more orderly arrangement, owing to the extensive hydrogen bonding.



- the expansion of water upon freezing. Generally, as substances are heated, they expand. However, water behaves a little differently. The density of water is greatest at 4 °C. (Density refers to how much mass is in a given volume; for example, oil is less dense than water so it floats on top of water.) Water does expand when heated from 4 °C but, unusually, it also expands as the temperature decreases from 4 °C. As the temperature continues to decrease, water at 4 °C sinks and eventually the temperature at the surface becomes 0 °C and freezes. Ice forms in an open, hexagonal crystalline lattice that places the water molecules further apart than occurs in the liquid state. Because water expands on freezing, it is less dense as a solid. Ice, therefore, floats on water.

If ice were denser than water, some bodies of water would freeze solid during winter and the aquatic life would die. Icebergs form when sections of the frozen icecap break off and float in the sea. When ice forms on the surface of water, it acts as an insulator, preventing the water below from freezing. This means that aquatic life can survive, even in sub-zero conditions.

FIGURE 6.28 The density of water is greatest at 4 °C

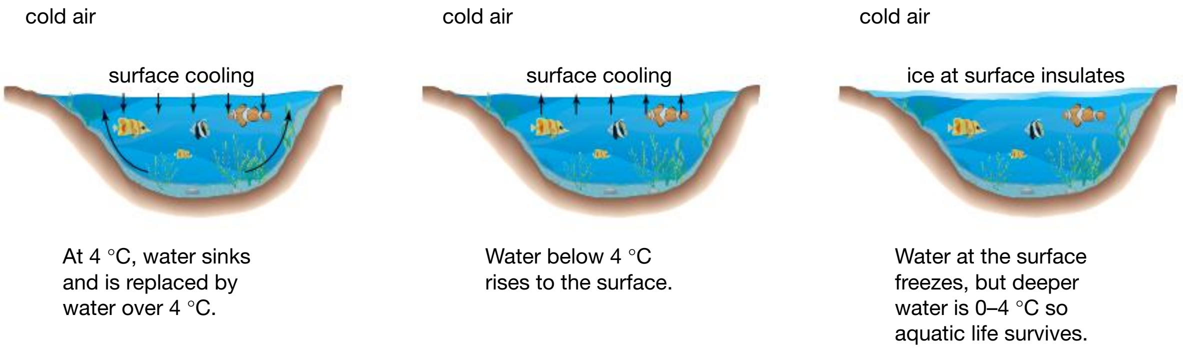
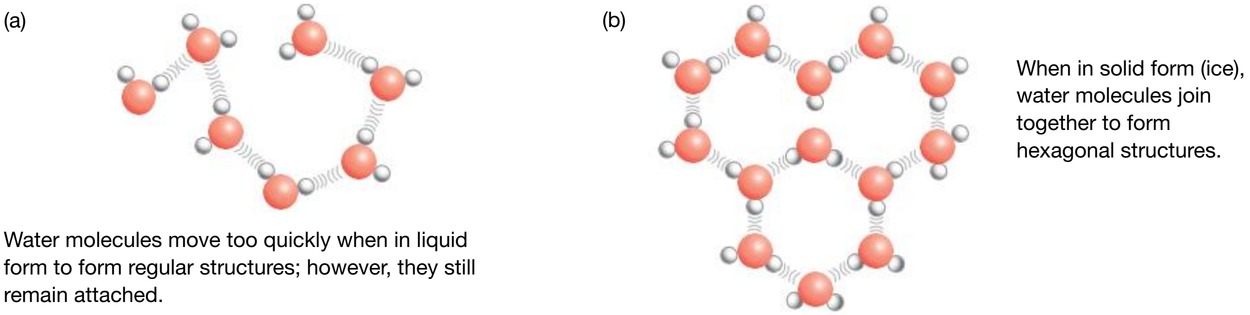


FIGURE 6.29 Hydrogen bonding in water when (a) water molecules are in liquid form and (b) water molecules in solid form (ice)



6.5 EXERCISE

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1. Place the following list of bonding types in order, from weakest to strongest.
   covalent bonding, hydrogen bonding, dispersion forces, dipole–dipole attractions
2. Consider the following list of compounds: CF_4, CCl_4, CBr_4.
   (a) Does hydrogen bonding exist in any of these substances?
   (b) Which has the strongest bonding within its molecules?
   (c) Which has the strongest dispersion forces between its molecules?
   (d) Which has the highest boiling point?
3. Consider the following list of compounds: CH_3F, CH_3Cl, CH_3Br.
   (a) Does hydrogen bonding exist in any of these substances?
   (b) Which substance has the weakest intramolecular bond(s)?
   (c) Which substance has the strongest dispersion forces between its molecules?
4. How would you expect the boiling points of argon (Ar) and fluorine (F_2) to compare? Explain.
5. Write the structural formulas for the following and name the shape of each molecule.
   (a) O_2
   (b) $CHCl_3$
   (c) Br_2
   (d) OF_2
   (e) CCl_4
   (f) C_2H_2
6. Draw and name the shape diagrams of the following compounds.
   (a) CF_4
   (b) PH_3
   (c) HCl
7. Predict whether the following molecules are polar or non-polar.
   (a) HI
   (b) SiH_4
   (c) CS_2
   (d) SF_2
   (e) CH_3Cl
   (f) C_2H_4
8. CH_3F has a boiling point of –78 °C, whereas CH_3OH has a boiling point of 65 °C. Explain this difference.
9. Which of the following exhibits hydrogen bonding? Why?
   (a) H_2O
   (b) C_2H_6
   (c) CH_3OH
   (d) CH_3CH_2OH
10. Explain why a full, sealed glass bottle of water will crack when left in a freezer.

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6.6 Review

6.6.1 Summary

Representing molecules

- Covalent molecular substances are substances held together by covalent bonds and include elements and compounds. These substances are made up of molecules.
- A molecule is a discrete (separate) group of non-metallic atoms held together by covalent bonds. The atoms are combined in a fixed ratio and are electrically neutral.
- A covalent molecular element is made up of molecules of the *same* element. The molecules may contain varying numbers of atoms. These include diatomic molecules, which are composed of two atoms that share electrons (e.g. Cl_2).
- A covalent molecular compound is made up of atoms of *different* elements (e.g. CH_4).
- Atoms tend to lose, gain or share electrons in order to achieve a stable noble gas configuration of eight electrons in the valence shell. The exception to this rule is the hydrogen atom, which needs only two electrons to complete its outer shell.
- Covalent bonds are produced when non-metallic atoms react together, *sharing* pairs of electrons so that each may get eight outershell electrons. These bonds are the result of the force of attraction between shared electrons and the nuclei of the non-metal atoms in the bond.
- Covalent bonds may be single bonds (formed by one shared pair of electrons) or multiple bonds.
  - In a double bond, two electron pairs are shared by two atoms.
  - In a triple bond, three electron pairs are shared by two atoms.
- Molecules can be represented in several ways to better understand their structures. Ball-and-stick, space-filling and computer-generated models are used to represent the three-dimensional structure of molecules. Two-dimensional models include:
  - electron dot diagrams
  - valence structures, which replace each pair of electrons (bonding and non-bonding) with a dash (–)
  - structural formulas, where dashes are used for bonding electron pairs and non-bonding pairs are often omitted
  - combinations of these representations.
- Though electrons are identical, when representing a covalent bond using an electron dot diagram they may be thought of as either bonding or non-bonding electrons.
  - Bonding electrons are the single electrons available for sharing, and they determine the number of bonds an atom can form with other non-metals.
  - Non-bonding electrons or lone pairs are paired electrons that are not available for sharing.
- The atom with the most bonding electrons is known as the central atom.

Predicting molecular shape

- Molecule shapes can be determined using electron dot diagrams and VSEPR theory, which allows us to determine the best shape for minimum repulsion between the electron pairs around the central atom.
- In general, the molecule shape is:
  - linear if the central atom has *one* bonding pair or one pair on either side (e.g. H_2, CO_2)
  - V-shaped if the central atom has *two* bonding pairs and *two* lone pairs (e.g. H_2O)
  - planar if the central atom has *three* bonding pairs and no lone pairs (e.g. C_2H_4).
  - pyramidal if the central atom has *three* bonding pairs and *one* lone pair (e.g. NH_3)
  - tetrahedral if the central atom has *four* bonding pairs and *no* lone pairs (e.g. CH_4).
- Non-metallic atoms have high electronegativities, which means that they have a strong attraction for the shared electrons in a covalent bond. Differing electronegativities cause electrons to be unequally shared, and can affect the polarity of the bond.
  - Bonds across which no difference exists in electronegativity will share electrons equally and are non-polar.

  - Bonds across which a difference in electronegativity does exist will share electrons unequally and are polar.
- Molecules can be polar or non-polar. This depends on whether their bonds are polar and on their shape.
  - If all bonds are non-polar, the molecule will be non-polar.
  - If the molecule contains polar bonds but the effects of all of these cancel out due to the molecule's shape, the molecule will be non-polar.
  - If the molecule contains polar bonds but the effects of all of these do not cancel out due to the molecule's shape (i.e. the molecule is asymmetric), the molecule will be polar.

Explaining the properties of molecular substances

- In addition to the intramolecular forces that work within molecules to hold them together, weak bonds exist between molecules and are called intermolecular forces.
  - Dispersion forces are found in all atoms or discrete molecules and increase with the size of the atom or molecule and the corresponding number of electrons.
  - Dipole–dipole interactions are found only in polar discrete molecules.
- Hydrogen bonding is a special case of dipole–dipole interaction. It is found only in molecules where a hydrogen atom is directly bonded to a more electronegative fluorine, oxygen or nitrogen atom. These bonds are stronger than other dipole–dipole bonds and result in higher melting and boiling points.
- Discrete covalent molecular substances have similar properties. Molecular substances:
  - have low melting and boiling points due to weak intermolecular forces
  - are usually liquid or gaseous at room temperature due to weak intermolecular forces
  - do not conduct electricity because the molecules are electrically neutral
  - are soluble in water if polar, and soluble in non-polar solvents if non-polar (the 'like-dissolves-like' rule)
  - are soft.

The relative strengths of bond types in molecular substances

- The forces existing within and between the molecules of a covalently bonded substance have different strengths. In order of decreasing strength, these are:

  Covalent bonding >>>>> hydrogen bonding >> dipole–dipole attraction > dispersion forces.

- Hydrogen bonding between water molecules explains some of its unusual properties, including its expansion on freezing and the floating of ice on water. It also explains the relatively high melting and boiling points for water.

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6.6.2 Key terms

ball-and-stick model	representation of a molecule in which the atoms are shown as balls and the bonds as sticks
bond dipole	separation of charge in a polar covalent bond
bonding electrons	the pairs of electrons involved in forming a covalent bond
central atom	the atom in a molecule with the most bonding electrons
computer-generated model	representation of a molecule that is produced by a computer
covalent bond	sharing of electrons between nuclei that bonds them together in a molecule
covalent molecular compound	a molecular compound in which atoms of different elements share electrons with each other

covalent molecular element	element made up of identical atoms held together by covalent bonds
diatomic molecule	substance containing two atoms only
dipole–dipole interactions	weak bonding caused by the positive end of one dipole attracting the negative end of another dipole
discrete	separate, distinct; not in an infinite array or lattice
dispersion force	the bond between adjacent molecules formed by instantaneous dipoles; this weak non-directional bonding is also known as van der Waals force
double bond	strong bond between two atoms formed by two pairs of electrons that are shared by the two nuclei
electron dot diagram	representation where the atom's nucleus and all innershell electrons are replaced by its element symbol and the outershell electrons are represented by dots around the symbol in a square arrangement
electronegativity	the electron-attracting power of an atom
hydrogen bonding	the bond between a hydrogen atom covalently bonded to an atom of F, O or N and another molecule that also contains an atom of H, F, O or N
intermolecular bonding	bonding that occurs between molecules
intramolecular bonding	internal bonds within a molecule
molecule	group of atoms bonded together covalently
multiple bond	bond formed when two atoms share two or more pairs of electrons
non-bonding electrons	electrons that are not involved in bonding
non-polar covalent bond	bond formed between atoms with the same electronegativity
non-polar molecule	molecule that does not have permanent dipoles or is symmetrical
octet rule	a generalisation that works for many (but not all) atoms, stating that atoms will donate or share electrons in order to achieve eight electrons in their outer shells
polar covalent bond	bond formed when two atoms that have different electronegativities share electrons
polar molecule	a molecule which, due to its polar bonds *and* its asymmetric shape, has an overall imbalance in the distribution of its electrons
polarity	localised imbalances in electric charges within a molecule resulting in a negatively charged end and positively charged end
space-filing model	three-dimensional representation of a molecule that shows the relative sizes of atoms within the molecule and the distances between them
structural formula	a diagrammatic representation of a molecule showing every bond
triple bond	strong bond between two atoms formed by three pairs of electrons that are shared by the two nuclei
valence electrons	outershell electrons
valence shell electron pair repulsion (VSEPR) theory	a model with the main point that the structure around a given atom in a molecule is determined principally by minimising electron repulsions
valence structure	a diagrammatic representation of the outershell electrons in a molecule; similar to an electron dot diagram but replaces each pair of electrons with a dash (–)

on Resources

Digital document Key terms glossary — Topic 6 (doc-30947)

6.6.3 Practical investigations

Experiment 6.1

Building molecular models

Aim: To construct molecular models of H_2, Cl_2, HCl, H_2O, NH_3, CH_4, CH_2F_2, H_2S, C_2H_6, C_2H_5Cl and CH_3CH_2OH and then use these to determine molecular shapes

Digital document: doc-30848

Teacher-led video: tlvd-0623

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Digital document Practical investigation logbook (doc-30948)

6.6 Exercises

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

6.6 Exercise 1: Multiple choice questions

1. The best model to illustrate the shape of a molecule is which of the following?
   - **A.** Electron dot
   - **B.** Space-filling
   - **C.** Valence structure
   - **D.** Ball-and-stick
2. A molecule has four bonding pairs of electrons around its central atom. What would the shape of this molecule be described as?
   - **A.** Tetrahedral
   - **B.** Pyramidal
   - **C.** V-shaped
   - **D.** Linear
3. A molecule has three bonding pairs of electrons and a lone pair around its central atom. What would the shape of this molecule be described as?
   - **A.** Tetrahedral
   - **B.** Pyramidal
   - **C.** V-shaped
   - **D.** Linear
4. What does the sharing of electrons in bond formation *always* involve?
   - **A.** The formation of positive and negative ions
   - **B.** The formation of polar molecules
   - **C.** Shared electrons being attracted more by one atom than another
   - **D.** The bonded atoms having greater stability than the unbonded atoms
5. Which one of the following electron dot diagrams best represents the bonding in the nitrogen molecule?
   - **A.** 
   - **B.** 
   - **C.** 
   - **D.** 
6. Which of the following is a characteristic 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. What is the structure of C_2H_2?
   - **A.** Linear
   - **B.** V-shaped
   - **C.** Planar
   - **D.** Pyramidal

8. Which of the following best describes the molecular shape, polarity of bond and molecular polarity of the H_2S molecule?
   - A. Linear, polar covalent, non-polar
   - B. Linear, polar covalent, polar
   - C. V-shaped, ionic, polar
   - D. V-shaped, polar covalent, polar
9. Which of the following substances might one expect to exhibit the weakest intermolecular forces?
   - A. HCl
   - B. He
   - C. NH_3
   - D. H_2O
10. At room temperature, methane is a gas while tetrachloromethane is a liquid. Choose the statement that best explains this.
   - A. There is appreciable hydrogen bonding in tetrachloromethane but not in methane.
   - B. The tetrachloromethane molecule is polar, while the methane molecule is non-polar.
   - C. Tetrachloromethane has an appreciably higher molecular mass than methane.
   - D. The bonds in tetrachloromethane are polar covalent.
11. What is the correct order of decreasing attractive strength, from highest to lowest, for weak intermolecular forces?
   - A. Hydrogen bonding, dipole–dipole interaction, dispersion forces
   - B. Hydrogen bonding, dispersion forces, dipole–dipole interaction
   - C. Dipole–dipole interaction, hydrogen bonding, dispersion forces
   - D. Dispersion forces, dipole–dipole interaction, hydrogen bonding.
12. Which one of the following substances forms hydrogen bonds?
   - A. HCOOH
   - B. CH_3CN
   - C. SiH_4
   - D. CH_3OCl
13. Which of the following *best* explains the relatively low melting point of covalent molecular substances?
   - A. Covalent molecular materials depend on weak electrostatic forces holding the ions together.
   - B. The intermolecular forces between the molecules are weaker than ionic or covalent bonds.
   - C. The metals involved create uneven bonding with the non-metals.
   - D. The similar electronegativity of the atoms causes repulsions between the molecules.
14. Which of the following is *not* a covalent compound?
   - A. CaO
   - B. SO_3
   - C. NH_3
   - D. CH_3Cl
15. How can the floating of ice on water be explained?
   - A. Each water molecule forms four hydrogen bonds with neighbouring molecules, resulting in the formation of crystals with the molecules closer together than in the liquid.
   - B. Each water molecule forms four hydrogen bonds with neighbouring molecules, resulting in the formation of crystals that have the molecules further apart than in the liquid.
   - C. The water molecules form crystals in which they are held together by dispersion forces, with the water molecules closer together in these crystals than in the liquid.
   - D. The water molecules form crystals in which they are held together by dispersion forces, with the water molecules further apart in these crystals than in the liquid.

6.6 Exercise 2: Short answer questions

1. List three typical properties of covalent molecular substances.
2. a. How many electrons are in the outer shell of all of the following elements: Ne, Ar, Kr, Xe, Rn?
   - b. To which group number do these elements belong?
   - c. What is the name of this group?
   - d. He is also a member of this group. Why does He have a different number of outershell electrons from all the other members?

3. a. What is the octet rule?
   b. Is the octet rule always obeyed?
4. Give the group number of the element that each of the following electron dot diagrams represents.
   a. 

   b. 

   c. 

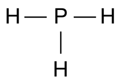
   d. 

5. Use electron dot diagrams to draw the structural formulas of the following.
   a. H_2S
   b. NCl_3
   c. CH_4
   d. Br_2
   e. HOCl
   f. H_2CO
6. Write the structural formulas for the following and name the shape of each molecule.
   a. NH_3
   b. CH_3OH
   c. HCN
   d. PCl_3
   e. N_2
7. Draw and name the shape diagrams of the following compounds.
   a. H_2S
   b. NF_3
   c. C_2Br_4
8. Predict whether the following bonds are non-polar covalent, polar covalent or ionic.
   a. H—H
   b. H—S
   c. C—S
   d. Cl—F
   e. Na—N
9. a. Draw a polar molecule and explain what this means.
   b. Why are molecular elements always non-polar?
   c. Why are molecular compounds sometimes polar and sometimes non-polar?
10. State whether the following molecules are polar or non-polar: NH_3, CH_3OH, HCN, PCl_3, N_2, H_2S, NF_3, C_2Br_4.
11. Explain why the ammonia molecule is polar but carbon dioxide molecules are not.
12. a. The boiling point of iodine is 184 °C, whereas fluorine has a boiling point of –188 °C. Explain this difference.
    b. What states do fluorine and iodine assume at room temperature? Explain.
13. Which of the following sets of compounds has the stronger intermolecular forces? Explain your answer.
    a. CO_2 or OCS
    b. HF or HBr
14. Which of the following exhibits hydrogen bonding?
    a. CF_3H
    b. CH_3OCH_3
    c. NH_3
    d. HF
15. Methanol and ethanol have the formulas CH_3OH and CH_3CH_2OH. Predict which has the higher boiling point and give clear reasons why.

6.6 Exercise 3: Exam practice questions

Question 1 (8 marks)

A molecule of phosphine, PH_3, has the following structural formula.



$$\begin{array}{c} H-P-H \\ | \\ H \end{array}$$

a. Does this molecule have any lone pairs on the central P atom? **1 mark**
b. Draw the shape diagram for phosphine, making sure you make an attempt to demonstrate its three-dimensional structure and shape. **2 marks**
c. What is the shape of the phosphine molecule? **1 mark**
d. Explain how you determined the shape shown in (b). **2 marks**
e. A molecule may have a tetrahedral arrangement of electron pairs around its central atom, but its shape may be described as pyramidal (rather than tetrahedral). Explain. **2 marks**

Question 2 (6 marks)

The amount of CO_2 in the atmosphere is increasing.

a. Use dot diagrams to determine the structure and shape of a CO_2 molecule. **1 mark**
b. What is the shape of the CO_2 molecule? **1 mark**
c. How many lone pairs are in a molecule of carbon dioxide? **1 mark**
d. Are the bonds in the molecule polar or non-polar? **1 mark**
e. Is the molecule polar or non-polar? **1 mark**
f. What type of interactions would you expect between the molecules in solid carbon dioxide? **1 mark**

Question 3 (5 marks)

Ammonia, NH_3, is a gas at room temperature. It has a characteristic pungent odour.

a. Show how the shape of the molecule is determined. **2 marks**
b. Discuss the polarity of the molecule. **1 mark**
c. Explain in terms of the intermolecular bonding why ammonia is a gas at room temperature. **2 marks**

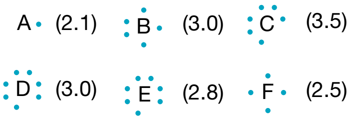
Question 4 (9 marks)

Water has some unique properties and we cannot live without it.

a. Use a labelled diagram to explain why the water molecule is V-shaped. **3 marks**
b. Explain the intramolecular and intermolecular forces in water. **2 marks**
c. Why is the density of ice less than the density of water? **2 marks**
d. Describe an experiment to show the difference in density between water and ice. **2 marks**

Question 5 (10 marks)

The electron dot diagrams for six elements are shown with their corresponding electronegativities in brackets. The usual symbols have been replaced by the letters A to F.



a. Describe the position of D and E relative to each other on the periodic table. **1 mark**
b. Name a possible element that C might represent. **1 mark**
c. Element A forms covalent bonds with B, C, D, E and F. Which bonds will be the most polar? **1 mark**
d. Draw electron dot diagrams for the following. **3 marks**
   i. A bonding with B
   ii. A bonding with D
   iii. A bonding with F
e. Which of the molecule(s) from part (d) will be **3 marks**
   i. polar?
   ii. non-polar?
f. What is the shape of the molecule formed between A and B? **1 mark**

6.6 Exercise 4: studyON Topic Test

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7 Carbon lattices and carbon nanomaterials

7.1 Overview

Numerous **videos** and **interactivities** are available just where you need them, at the point of learning, in your digital formats, learnON and eBookPLUS at www.jacplus.com.au.

7.1.1 Introduction

The lead in the 'lead' pencil that you use to record your notes is not actually made of lead; it is made of a form of carbon called graphite that is mixed with clay. When you write, the pencil makes contact with the paper and many layers of carbon atoms are deposited. Although lead was used by the Romans to make marks on papyrus, it was not until the 1500s that graphite was used for writing. At that time, sticks of graphite were wrapped in string so that they were easy to hold. It was mistakenly thought that the shiny graphite was a form of lead. Graphite is also a source of a useful nanomaterial graphene, which has some remarkable properties.

FIGURE 7.1 While diamonds are composed of the same atoms as graphite (carbon), their arrangement causes their vastly different physical properties.

You might also find it extraordinary to discover that the dull, black substance of graphite has exactly the same chemical composition as spectacular, light refracting diamonds. Both are composed of carbon atoms, and it is their different arrangement that creates their different properties. In fact, graphite and diamond can even change backwards and forwards into each other.

This topic examines a type of covalent bonding that forms lattices instead of discrete molecules, the effect this has on a substance's properties and, hence, why they are put to the uses that they are. This topic also explores a relatively new field of carbon chemistry: the study of carbon-based nanomaterials such as fullerenes and graphene.

The basics of covalent bonding and spatial arrangement introduced in topic 6 will be important, and understanding the metric system and its prefixes will be critical to appreciating the scale of nanotechnology. An exposure to the concept of nanomaterials will also prove to be advantageous.

7.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- the structure and bonding of diamond and graphite that explain their properties (including heat and electrical conductivity and hardness) and their suitability for diverse applications
- the structures, properties and applications of carbon nanomaterials, including graphene and fullerenes.

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Digital documents Key science skills (doc-30903)
Key terms glossary — Topic 7 (doc-30950)

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7.2 Covalent lattice bonding

KEY CONCEPT

- The structure and bonding of diamond and graphite that explain their properties (including heat and electrical conductivity and hardness) and their suitability for diverse applications

7.2.1 Carbon

What do you think is the most important element? Carbon should probably be at the top of the list. It has been known from ancient times and yet there are still many exciting discoveries of new materials based on carbon. Carbon, with its four bonding electrons, is present in all forms of life and is known to form millions of compounds. We saw in topic 6 that this element forms simple, small molecular compounds such as methane, CH_4, and carbon dioxide, CO_2. We breathe out carbon dioxide, and it is an essential reactant in photosynthesis, the process by which plants make food. Carbon dioxide produced as a result of human activity is increasing and so contributing to climate change. Topic 4 introduced ionic metal carbonate compounds that are components of many rocks.

Topic 8 focuses on **organic chemistry**, which is a branch of chemistry that is devoted entirely to compounds of carbon from the small to very large molecules. Topic 9 introduces the incredible versatile materials that are generally called plastics, but scientists call polymers, which are large molecules with carbon backbones. Organic compounds are obtained from fossil fuels that have been formed from the decayed remains of plants and formed over millions of years. Fossil fuels also provide us with energy for households, transportation and manufacturing. Carbon in its different forms is a major part of our lives.

How can both the hardest material, diamond, and one of the softest materials, graphite, consist of the same element? How can one material be an insulator and another a conductor? It is all about the arrangement of the atoms. Diamond and graphite are naturally occurring **allotropes** of carbon; 'allotrope' is the term used to describe different forms of the same element. In this topic, carbon's ability to form both network lattices and layer lattices is discussed as well as new developments in nanotechnology that are also based on carbon.

7.2.2 Covalent network lattices

Some non-metals form giant structures in which no individual molecules exist. They consist of countless numbers of atoms covalently bonded to each other, forming a *three-dimensional* **network lattice**. Common examples of covalent network lattices include diamond, silicon carbide, silicon dioxide (quartz) and tungsten carbide.

Diamond — the hardest known substance

Diamond is the hardest substance known because of the bonding pattern of its carbon atoms. Their hardness and brilliance are the consequence of the ordered and rigid internal structure in which each carbon atom bonds with four single covalent bonds to four other carbon atoms. Millions of carbon atoms can combine in a three-dimensional covalent network lattice to form a diamond crystal.

Hardness is measured on the **Mohs scale** (see table 7.1). The Mohs scale gives the order of hardness, but this is not linear; the difference between corundum and diamond is only one on this scale but diamond is many times harder than corundum.

A diamond is really one 'giant' molecule. It is made up of carbon atoms bonded together by strong covalent bonds. Each carbon atom has four covalent bonds around it and is surrounded by four other carbon atoms. The **covalent bonding** is three-dimensional since the bonds are arranged tetrahedrally around each atom. To scratch a diamond, the carbon atoms need to be separated. This is very difficult to do because they are held together by four covalent bonds.

FIGURE 7.2 Diamonds are prized for their brilliance, durability and hardness. These properties are due to their rigid internal structure.

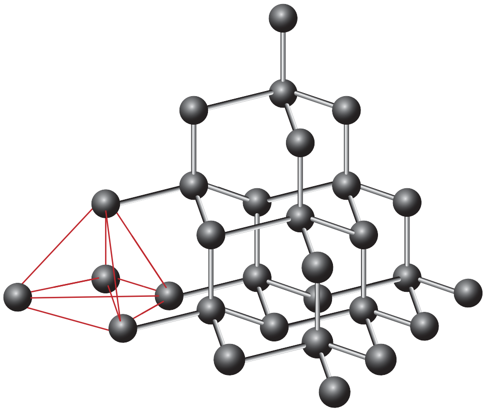


TABLE 7.1 Mohs hardness scale

Hardness	Mineral	Scratch test	Other materials
1	Talc	Very easily scratched by fingernail	Sulfur, sodium, chalk
2	Gypsum	Scratched by fingernail	Lead, **graphite**
3	Calcite	Very easily scratched with knife	Gold, silver, copper
4	Fluorite	Easily scratched with knife	Platinum, nickel
5	Apatite	Scratched with knife	Tooth enamel, glass (5.5)
6	Orthoclase	Scratches glass	Manganese, steel
7	Quartz	Scratches glass easily	Vanadium
8	Topaz	Scratches glass very easily	Cubic zirconia
9	Corundum	Cuts glass	Tungsten carbide, ruby, sapphire
10	**Diamond**	Scratches all other materials	

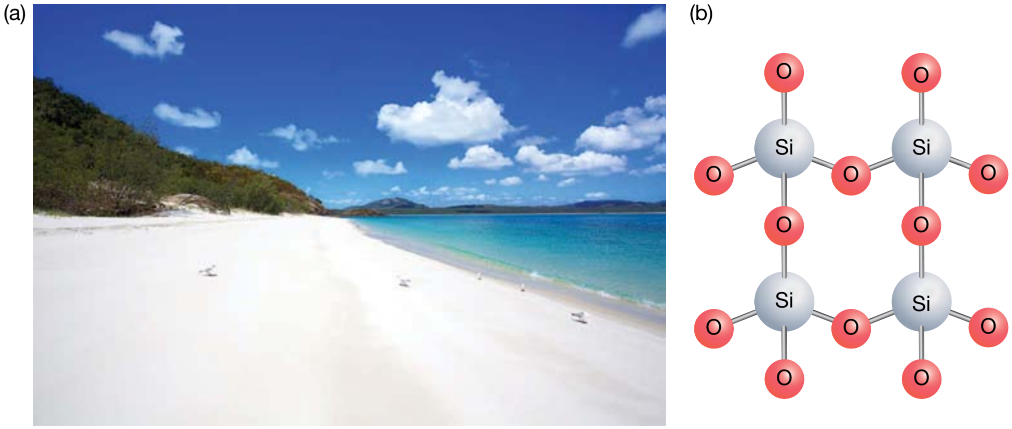
Diamonds are used in industrial situations where a very hard substance is required for a particular purpose. They are typically used as an abrasive for sawing, cutting and grinding hard substances such as glass, stone and porcelain. The tips of drilling bits used on oil rigs and other similar rigs that need to bore through hard rock are coated with diamonds. All of these uses are explained by the strong covalent bonding that exists throughout the crystals and which makes diamond exceedingly hard and suitable for these purposes.

In jewellery, cut diamonds are prized for their beauty. Diamonds have a very high **refractive index** and are cut at angles and polished so that the light entering them is refracted to make them sparkle. As can be imagined, due to diamond's extreme hardness this is no easy task and specialised techniques and equipment are required.

Silicon dioxide

Another covalent network lattice is silicon dioxide, SiO_2. Silicon dioxide is found as quartz in rocks such as granites and sandstones, and very pure varieties weather to white sand beaches, as shown in figure 7.3(a). Quartz, SiO_2, is a three-dimensional covalent network lattice shown in figure 7.3(b). Each silicon atom is bonded to four oxygen atoms, and each oxygen atom is bonded to two silicon atoms. The resulting substance is a hard, crystalline rock.

FIGURE 7.3 (a) Whitehaven beach in Queensland is renowned for its ultra-white sand, which consists of 98% pure silicon dioxide (quartz). (b) The structure of silicon dioxide.



Properties of covalent network lattices

Covalent network lattices have the following general properties:

- Since strong covalent bonding extends through the crystal structures of covalent network lattices, the particles are held rigidly. These substances are, therefore, very hard, difficult to scratch and have high melting points and boiling points.
- Since no free ions or electrons are in the structure, covalent network lattices are usually non-conductors of electricity in the solid and liquid states.
- Covalent network lattices are brittle and must be cut in a specific way or they can shatter. When the covalent bonds break, the lattice is distorted.
- Covalent network lattices are chemically inert and are insoluble in water and most other solvents.
- Covalent network lattices such as diamond are good conductors of heat. The conduction of heat is a mechanical process due to the collision of particles. A collision at one end of a diamond will be effectively transmitted through the crystal structure because the atoms are so rigidly bonded together — there is little 'give' in the bonds to soak up the energy of the collision.

SAMPLE PROBLEM 1

Pure germanium (Ge) forms crystals in the same way as diamond.

a. Describe the bonding that occurs in a crystal of germanium with respect to type and orientation.

b. Describe one difference between the bonding in germanium compared to diamond.

Teacher-led video: SP1 (tlvd-0535)

THINK	WRITE
a. Recall that in diamond each carbon has four covalent bonds that are bonded to four other C atoms, and so on. These bonds are arranged tetrahedrally. No discrete molecules are present.	Germanium atoms will be covalently bonded to four other germanium atoms, which in turn will be bonded to four more each, and so on. The arrangement of bonds around each germanium atom will be tetrahedral.

b. Germanium is in period 4 compared to carbon in period 2, so its bonding electrons are further from the nucleus than carbon, which makes the bond length longer.

The bonds are longer in germanium than in diamond.

PRACTICE PROBLEM 1

Pure silicon (Si) forms crystals in the same way as diamond.

a. Describe the bonding that occurs in a crystal of silicon with respect to type and orientation.

b. Describe one difference between the bonding in silicon compared to diamond.

7.2.3 Covalent layer lattices

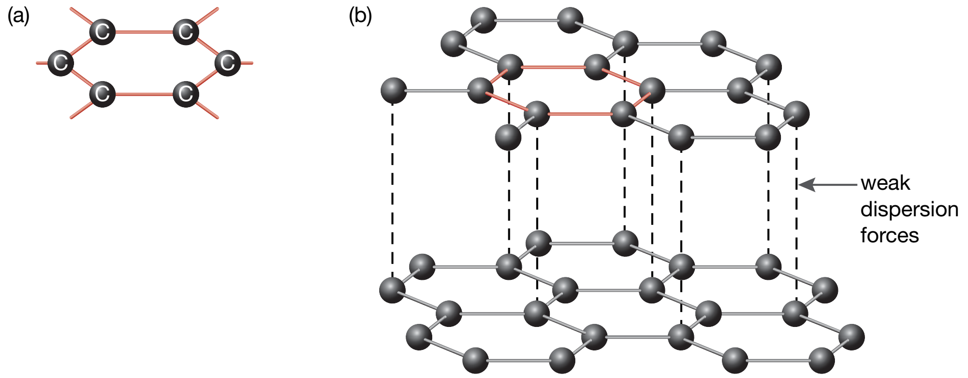
Covalent layer lattices consist of countless atoms held strongly together in planes by covalent bonds. Covalent bonding is, therefore, present in *two dimensions only*. These layers are held together by weaker dispersion forces. Graphite is a common example of a covalent layer lattice.

Graphite

Graphite is an oily, black, opaque solid with a metallic sheen. Like diamond, graphite is made solely of carbon atoms.

Graphite, however, looks and behaves quite differently from diamond, and this can be explained by its structure. Each carbon atom in graphite forms three covalent bonds with three other carbon atoms within the same plane. This forms layers of hexagonal rings (six atoms) with strong covalent bonds forming flat sheets in two dimensions. Since only three of the four outershell electrons are used in the covalent bonds, one outershell electron remains. This electron from each atom becomes **delocalised** and can move *across* the graphite layers. Layers are stacked on top of each other in the crystal lattice and are held together by weak dispersion forces.

FIGURE 7.4 Graphite is made of flat sheets of carbon atoms. Each carbon atom covalently bonds to three other carbon atoms, producing rings of six atoms that join to form flat sheets, as shown in (a). These sheets of carbon atoms lie on top of each other, held together by dispersion forces, as shown in (b).



Properties and uses of graphite

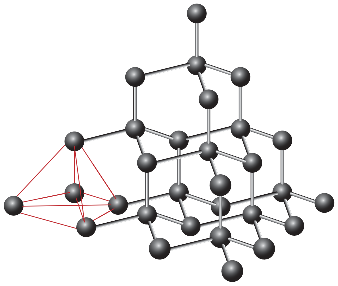
- Graphite is a good conductor of electricity because it contains charged particles that are free to move. These are the delocalised electrons left over from the bonding into the hexagonal rings that form its sheets.
- Graphite is a good conductor of heat also due to its delocalised electrons. The energy of collisions in one location can easily be transmitted through further collisions between these electrons and others to a different location.

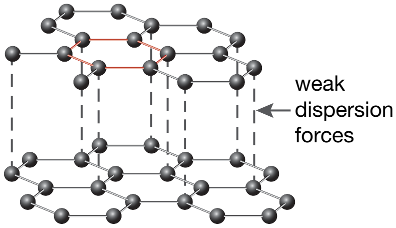
- Graphite is soft and feels slippery. Although the layers in graphite are held together by strong covalent bonds, only weak dispersion forces exist between them. The layers, therefore, slide over each other very easily. This makes it feel slippery and, in powdered form, this feeling is further enhanced. The most familiar use of graphite is as 'lead' in pencils. Here it is mixed with varying proportions of clay to make the different grades of pencil. The higher the proportion of graphite, the softer and darker the pencil will be. The weakly bonded layers slide easily over each other and are left behind as marks on the paper. This property also explains its use as a dry lubricant for machine parts and in locks.
- Graphite is inert (it can withstand chemical change) due to its stable hexagonal rings and delocalised electrons. It is used extensively as an inert electrode material in batteries and electrolysis applications due to its electrical conductivity and chemical inertness. It is also used as a moderator in nuclear reactors, where it absorbs fast-moving neutrons. It is also used in the making of graphene and synthetic diamonds.

7.2.4 Comparison of allotropes of carbon

Table 7.2 compares different allotropes of carbon.

TABLE 7.2 Various allotropes of carbon

Name	Structure	Property	Use
Diamond 	• Covalent network lattice: each carbon atom bonds with four single covalent bonds to four other carbon atoms, creating tetrahedral structure. 	• Hardest known substance due to rigid crystal structure • High refractive index • High melting and boiling points • Non-conductors of electricity due to lack of free ions or electrons in the lattice • Brittle • Chemically inert • Insoluble in water • Good conductor of heat due to collision of particles through crystal structure • Formed at very high pressures	• Sawing, cutting, drilling • Jewellery
Graphite 	• Covalent layer lattice: layers of hexagonal carbon rings with strong covalent bonds forming two-dimensional flat sheets, with sheets held together by weak dispersion forces.	• Soft • Low refractive index (opaque) • High melting and boiling points • Good conductor of electricity due to free moving delocalised electrons	• Dry lubricant • Pencils • Inert electrodes • Moderators in nuclear reactors

Name	Structure	Property	Use
	• Remaining outershell electron is delocalised and can move across graphite layers.  	• Slippery feel because layers can slide over each other easily • Chemically inert • Good conductor of heat due to free moving delocalised electrons	
Amorphous carbon 	• No consistent structure: impure forms of carbon including coal and soot are produced from incomplete combustion of organic material.	• Reactive • Conductive • No permanent crystal structure	• Charcoal fuel • Printer (toner) ink • Activated charcoal for water filtration, and treatment of poisoning and drug overdoses

on Resources

Weblinks Diamond and silicon dioxide

Graphite

7.2 EXERCISE

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1. Gemstones that are similar in appearance and structure to diamonds can be produced from a form of silicon carbide (SiC). These are called *synthetic moissanites*.
   (a) Describe the bonding that occurs in a crystal of synthetic moissanite with respect to type and orientation.
   (b) Describe one difference between the bonding in synthetic moissanite compared to diamond.
2. Why is carbon able to form such a variety of compounds?
3. The density of diamond is 3.5 g cm^{-3} and that of graphite is 1.9–2.3 g cm^{-3}. Explain why they are different.
4. Explain why the following substances are used in the way described.
   (a) Graphite in pencils
   (b) Diamond in oil drilling
5. Graphite is used to make electrodes in batteries. Explain why graphite is a good conductor of electricity and diamond is not.

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7.3 Carbon nanomaterials

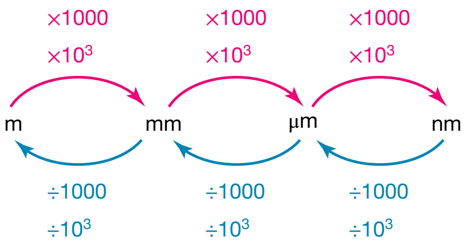
KEY CONCEPT

- The structures, properties and applications of carbon nanomaterials, including graphene and fullerenes

7.3.1 What are nanomaterials?

Nanomaterials were introduced in topic 1 and topic 3. They are generally considered to be materials that contain at least one dimension that is between 1 and 100 nanometres. A nanometre (nm) is a billionth of a metre. Nano is a metric prefix just like micro and milli. It represents a factor of 10^{-9}. A review of unit conversion is presented in figure 7.5 and in sample problem 2.

FIGURE 7.5 Converting between metres and nanometres



SAMPLE PROBLEM 2

A particular nanostructure has a thickness of 10 nanometres. Convert this to millimetres.

Teacher-led video: SP2 (tlvd-0536)

THINK	WRITE
Nano represents 10^{-9}. Milli represents 10^{-3}. A factor of 10^6 is, therefore, between them. 1 nm = 10^{-6} mm	$10\text{ nm} = 10 \times 10^{-6}\text{ mm}$ $= 1.0 \times 10^{-5}\text{ mm}$

PRACTICE PROBLEM 2

A particular nanostructure has a thickness of 1.0×10^{-4} mm. Convert this to nanometres.

The small size of nanomaterials means that the surface properties are much more important in determining properties of a nanomaterial, and the properties that result are often quite different to the more familiar larger scale properties of the same substance.

7.3.2 Fullerenes

Fullerenes are nanoscale allotropes of carbon. They are characterised by the fact that they have some type of hollow shape. Commonly observed shapes include spheres, ellipses and tubes.

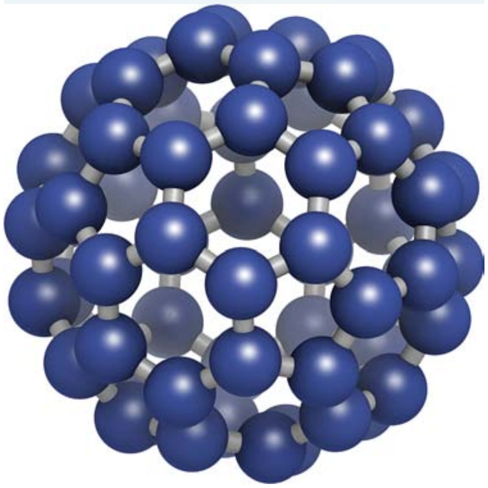
Buckyballs

Buckminsterfullerenes (or ‘buckyballs’ for short) are cage-like molecules of carbon. Small quantities of these unusual molecules exist in nature; however, in 1985 they were synthetically produced by vaporising graphite with a laser. They are formed in rings of carbon (in arrangements other than six carbons), and this means the arrangement is not flat (as is the case for graphite). The curvature induced by this arrangement leads to the formation of the large cage-like structures that are displayed by buckyballs. Despite this cage-like molecular structure, they are not considered giant lattice structures. The most common is C_{60}, which consists of a curved surface in the shape of a soccer ball made up of alternating hexagonal and pentagonal rings of covalently bonded carbon atoms. Other forms include C_{70}, C_{76} and C_{84} molecules.

Buckyballs are stable solids at room temperature, soft and resistant to collision, and have low melting points. They are insoluble in water but soluble in methyl benzene. In different forms, C_{60} can act as an insulator, a conductor, a semiconductor or, when doped with another element at low temperatures, a superconductor.

To date not many uses have been developed for buckyballs. However, research continues with possible future applications centring on the utilisation of the internal space for applications such as transporting medication and hydrogen storage for fuel cell cars.

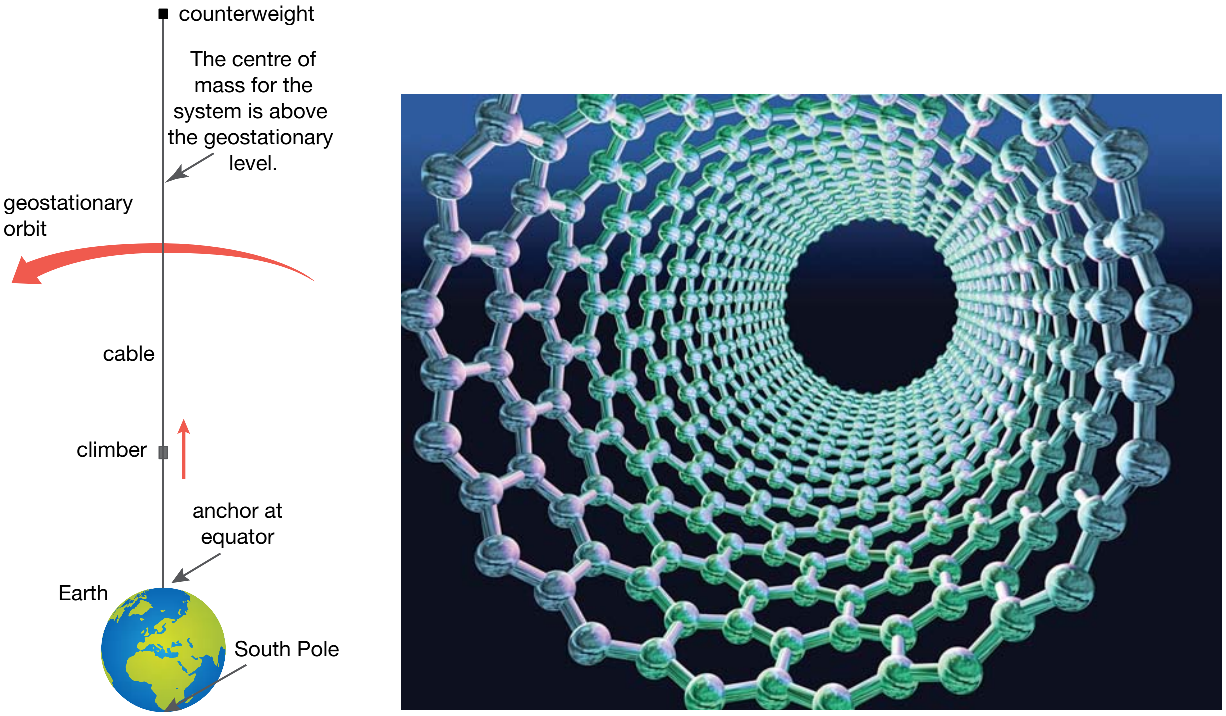
FIGURE 7.6 Buckyballs are cage-like molecules of carbon.



Nanotubes

An elevator into space is not as unrealistic as it might appear if suitable materials are available. **Carbon nanotubes** would be ideal for this purpose. They are continuous tubes made up of flat sheets of hexagonal rings of carbon atoms. These sheets are similar to those in graphite but rolled into a cylinder; they may form into a single layer or multiple layers. A human hair is about 10 000 times thicker than one of these very strong fibres. Nanotubes are about 100 times stronger than steel with one-fifth the density. Once established, the elevator could transport materials into space at a fraction of the cost of the massive amount of fuel used in rockets.

FIGURE 7.7 A space elevator made of carbon nanotubes anchored to an offshore sea platform would stretch to a small counterweight approximately 100 000 km away in space.



Nanotubes have extensive applications in many fields, and these are summarised in table 7.3.

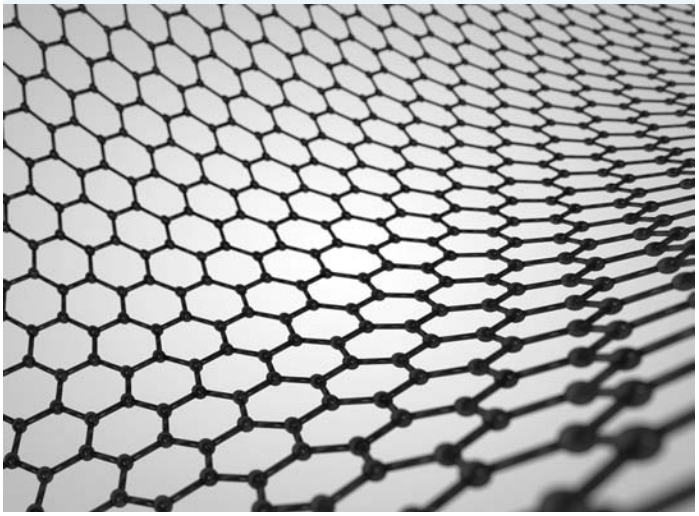
FIGURE 7.8 The Boeing 787 Dreamliner makes extensive use of carbon fibre, manufactured from carbon nanotubes, in its construction.

7.3.3 Graphene

Graphene is another nanoscale allotrope of carbon.

Imagine a mobile phone that is flexible enough to be worn around your wrist. This could indeed be possible with a material that was discovered in 2004. Graphene is similar to nanotubes but is a flat layer of carbon that is only one atom thick. This wonder material has some remarkable properties that make it superior to nanotubes for use in composite materials. Graphene's large surface area allows it to have closer contact with other materials, and its rippled surface enables interlocking with surrounding polymer atoms to readily form very useful composites.

FIGURE 7.9 Graphene is a flat layer of carbon that is only one atom thick.



Graphene is obtained from readily available graphite; it is much stronger than steel but light, transparent and flexible. It conducts electricity and heat more effectively than any metal. Potential applications based on these properties include lower cost mobile phone screens. Phones could be charged in minutes because of its ability to store and release energy quickly. Although impermeable, tiny pores could be introduced in graphene sheets and water could be economically desalinated or purified. Composite materials can use the lightness and strength of graphene for protective clothing or to make transport lighter and more fuel efficient.

7.3.4 A summary of properties and uses of carbon nanomaterials

Table 7.3 list some of the properties and uses of carbon-based nanomaterials.

TABLE 7.3 Properties and examples of uses for carbon nanomaterials

Property	Use
Buckminsterfullerene (buckyballs)	
• Internal cage-like structure • Easy to hydrogenate and dehydrogenate • High affinity for electrons	• Drug delivery • Hydrogen storage • Solar cells
Carbon nanotubes	
• High tensile strength • Electrically conductive • Very large surface area	• Tennis racquets and carbon-fibre bicycles • Composite materials reinforced with carbon fibres used for airplane fuselage and wings • Bulletproof vests and composite storage material • Huge surface area and low electrical resistance increases efficiency of electrodes. • Flat screen displays • Possible conduction of nerve impulses in health applications, artificial muscle tissue, sensors and targeted drug delivery. • Extend battery energy storage capacity • Can be used to attach catalysts for more efficient performance • Small size and even walls enable easier water purification by filtration
Graphene	
• Very strong • Biologically inert • Very large surface area • Electrically conductive • Adaptive properties	• Repair of muscle tendons • Killing of bacteria through graphene oxide (graphene with molecular oxygen) surrounding bacteria and puncturing its membrane • Attach catalysts for more efficient performance • Optical lenses • Solar cells • Bullet proof vests (can self-repair holes in sheets) • Water purification (can have pores introduced into sheets) • Lubricants

on Resources

Video eLesson	Allotropes of carbon (eles-2476)
Weblinks	Discovery of fullerenes
	Interactive molecules: Fullerene
	Uses for graphene

7.3 EXERCISE

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1. (a) How many micrometres are in a metre?
   (b) How many nanometres are in a metre?
   (c) How many nanometres are in a micrometre?
2. Human hair can vary in its thickness for a number of reasons. However, a typical value is about 0.025 mm. Express this in nanometres.
3. Both graphite and carbon nanotubes are made up of layers containing hexagonal rings of carbon atoms. Explain why some of their properties are so different.
4. Explain the difference in structure between buckyballs, nanotubes and graphene.

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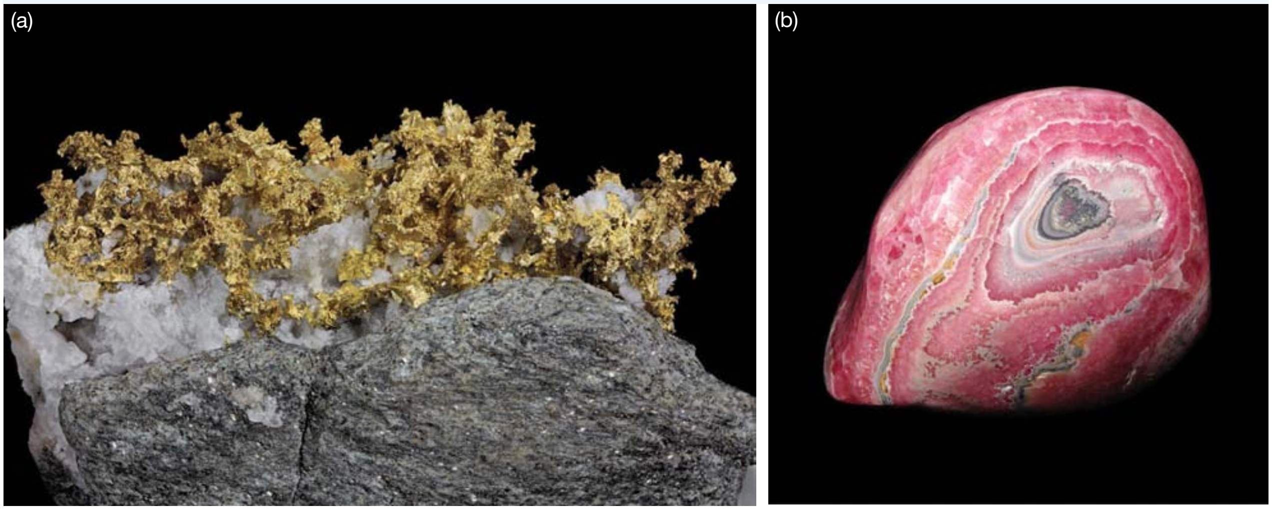
7.4 A review of bonding

BACKGROUND KNOWLEDGE

- A review of the different types of bonding between atoms, the different models for the structure of materials, and methods of classification using properties and composition

7.4.1 The basis of bonding

FIGURE 7.10 (a) Rare gold crystals and (b) red crystals of rhodochrosite (manganese carbonate)



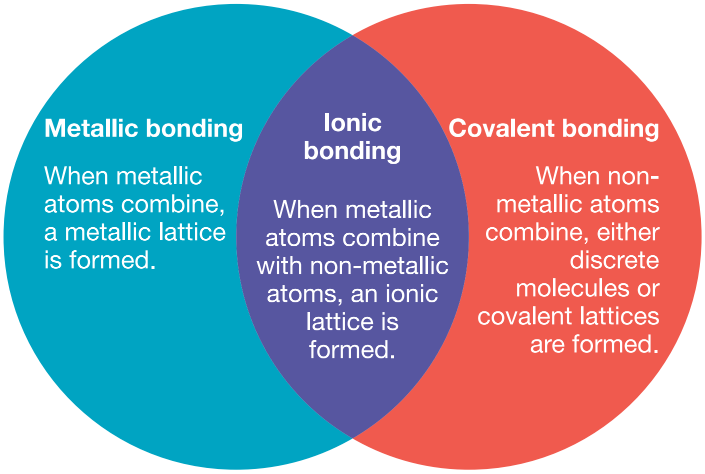
We have seen from the previous topics that only the noble gases exist as separate atoms due to their stable outer shell containing eight electrons. The atoms of all other elements form chemical bonds with each other in order to attain a more stable outershell configuration.

Atoms can become stable in one of three ways:

1. by giving electrons to another atom
2. by taking electrons from another atom
3. by sharing electrons with another atom.

When atoms combine to achieve more stable structures, three types of bonding are possible. This is illustrated in figure 7.11.

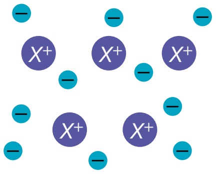
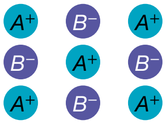
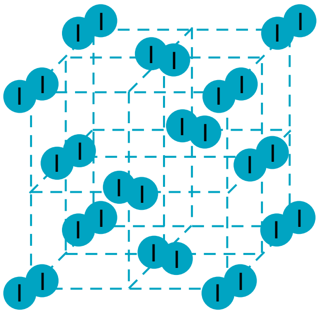
FIGURE 7.11 The three main types of bonding



7.4.2 Lattice structures

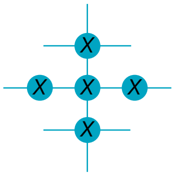
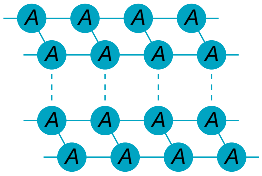
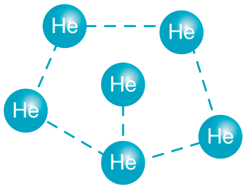
Atoms may bond together to form crystalline solids. The bonding between atoms gives rise to one of six lattice types as seen in table 7.4.

TABLE 7.4 The structures of different crystalline lattices

Type of lattice structure	Type of substance	Particles in the structure	Pictorial representation of the structure	Bonding types present	Examples
Metallic	Element or alloy that contains metals only	Cations and electrons		Metallic	Fe, Ca
Ionic	Compound that contains both a non-metal and a metal	Cations and anions		Ionic	NaCl, KF
Covalent molecular	Elements or compounds that contain non-metals only	Molecules		Covalent, dispersion forces; may also be dipole–dipole and hydrogen bonding	I_2

(Continued)

TABLE 7.4 The structures of different crystalline lattices *(Continued)*

Type of lattice structure	Type of substance	Particles in the structure	Pictorial representation of the structure	Bonding types present	Examples
Covalent network	Elements or compounds that contain non-metals only	Atoms		Covalent	SiO_2, C (diamond)
Covalent layer	Elements or compounds that contain non-metals only	Atoms		Covalent, dispersion forces	C (graphite)
Atomic	Noble gases	Atoms		Dispersion forces	Ne, Ar, Kr

7.4.3 Identifying the bonding in different substances

A material's properties depend on its structure, and its structure is determined by the nature of the constituent particles and the bonds holding them together. We can, therefore, identify the type of bonding in different substances using:

1. the properties of a substance
2. the composition of a substance.

Properties

The properties of different crystalline solids are a consequence of the type of bonding present in the lattice (see table 7.5).

TABLE 7.5 The properties of different crystalline lattices

			Electrical conductivity		Solubility in	
Type of lattice structure	**Melting point**	**Hardness**	**Solid**	**Molten or aqueous**	**Water (polar)**	**Petrol (non-polar)**
Metallic	High	Varies	Good	Good	Insoluble	Insoluble
Ionic	High	Hard	Poor	Good	Most are soluble	Most are insoluble
Covalent molecular	Low	Soft	Poor	Poor	Soluble if polar	Soluble if non-polar
Covalent network	Very high	Hard	Poor	N/A	Insoluble	Insoluble
Covalent layer	Very high	Soft	Good	N/A	Insoluble	Insoluble

SAMPLE PROBLEM 3

Identify the bonding in substance *A* and substance *B*, given the following information.

a. **Substance *A* is hard, melts at 890 °C and conducts electricity in the molten form but not in the solid form.**

b. **Substance *B* is soft, melts at −183 °C, and is a poor conductor of electricity in solid or liquid form.**

Teacher-led video: SP3 (tlvd-0537)

THINK	WRITE
a. Electrical conduction of *A* suggests it is ionic, and this is consistent with hardness and melting point.	Substance *A* has an ionic lattice structure in its solid state.
b. The low melting point of *B* suggests it is a molecular lattice, and this is consistent with the softness and electrical conductivity.	Substance *B* has a covalent molecular lattice structure in its solid state.

PRACTICE PROBLEM 3

Identify the bonding in substance *C* and substance *D*, given the following information.

a. **Substance *C* is hard, melts at 650 °C and conducts electricity in both the solid and liquid states.**

b. **Substance *D* is soft and is a good conductor of electricity in the solid state. It melts at 1700 °C.**

SAMPLE PROBLEM 4

Using the flow chart in figure 7.12, identify the bonding in:

a. **magnesium (a solid metal)**

b. **magnesium chloride (a solid salt)**

c. **chlorine gas.**

Teacher-led video: SP4 (tlvd-0538)

THINK	WRITE
a. Using the classification of substances flow chart, magnesium is a pure metal, so the bonding is metallic.	Magnesium displays metallic bonding.
b. Magnesium chloride is made up of both metal and non-metal elements; according to the flow chart, the bonding is ionic.	Magnesium chloride displays ionic bonding.
c. Chlorine gas is composed of diatomic molecules of non-metallic atoms. According to the flow chart, the bonding is covalent within the chlorine molecule. Because it is non-polar, dispersion forces also exist between the chlorine molecules.	Chlorine displays covalent bonding within its molecules and dispersion forces between them.

PRACTICE PROBLEM 4

Using the flow chart in figure 7.12, identify the bonding in:

a. solid copper

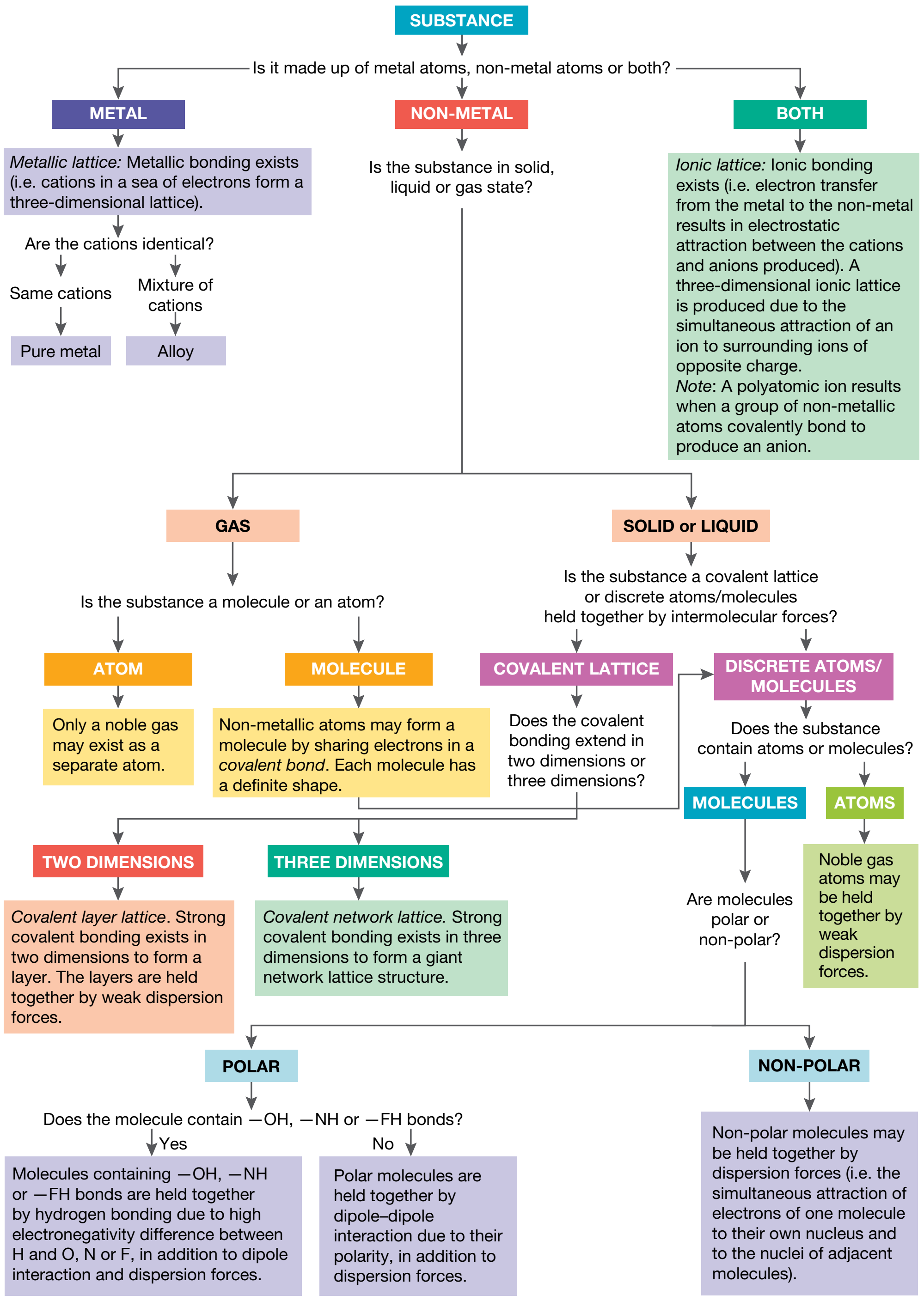
b. copper(II) oxide

c. oxygen gas.

Composition

Figure 7.12 shows how substances can be identified according to their composition.

FIGURE 7.12 Classification of substances according to composition



7.4 EXERCISE

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1. Identify the bonding in substance *E* and substance *F*, given the following information: substance E is soluble in water and has a melting point of –25 °C. Substance *F* is hard and is a poor conductor of electricity. It melts at 2500 °C.
2. Classify each of the following as covalent molecular, ionic or metallic. Justify your responses.
   (a) A yellow solid that melts at 105 °C to form a clear yellow liquid. Both solid and liquid are poor conductors of electricity.
   (b) A solid that melts at 99 °C to form a silvery liquid. Both solid and liquid are good conductors of electricity.
   (c) A dark shiny solid that sublimes to form a vapour. It is a poor conductor of electricity and heat.
   (d) A white solid that melts at 872 °C to form a colourless liquid. The solid does not conduct electricity, whereas the liquid does.
3. Consider the substances in the following table.

		Electrical conductivity		
Substance	**Melting point (°C)**	**Solid**	**Molten**	**Dissolved in water**
L	−115	—	—	—
M	1260	—	Conducts	Conducts
N	3600	—	—	Insoluble
O	2468	Conducts	Conducts	Insoluble

   (a) Which substance is metallic?
   (b) Which substance is ionic?
   (c) Which substance has a covalent molecular structure?
   (d) Which substance has a covalent network structure?
4. Use figure 7.12 to identify the bonding in the following substances.
   (a) Copper
   (b) Copper(II) chloride
   (c) Argon
   (d) Carbon tetrachloride liquid
   (e) Sulfur dioxide
5. Use figure 7.12 to identify the bonding in the following substances.
   (a) Steel
   (b) Diamond
   (c) Graphite
   (d) Ammonia gas
   (e) Iron(II) sulfate
6. Classify the type of lattice structure found in the following solids.
   (a) Sulfur
   (b) Silicon dioxide
   (c) Aluminium fluoride
   (d) Bronze
   (e) Tungsten
   (f) Dry ice
7. (a) Use the flow chart shown in figure 7.12 to classify the bonding within and between the particles that make up the following compounds.
   (b) Draw the shape diagram for each.
      i. CO_2
      ii. NH_3
      iii. CH_4
      iv. CH_3OH

▶

8. Which one of the following bonds has the strongest dipole? Justify your response.
   - **A.** C—H
   - **B.** O—O
   - **C.** N—Cl
   - **D.** H—F.
9. (a) Discuss the bonding in sodium chloride (table salt).
   (b) Why are salt crystals brittle?
   (c) What property of sodium chloride makes it useful as a flavouring? Relate this property to its structure and bonding.
10. Explain why, in terms of structure and bonding and their relationship to properties, the following materials have the uses given.
   (a) Gold is used to make jewellery.
   (b) Graphite is used as a lubricant in locks.

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7.5 Review

7.5.1 Summary

Covalent lattice bonding

- Carbon is a unique and ubiquitous element. It is essential to life and found in rocks, organisms, the atmosphere and fuels. It forms millions of compounds because it forms 4 bonds readily with itself and other elements. It can be found in small and large molecules as well as lattices.
- A covalent network lattice is a three-dimensional network lattice where many atoms are covalently bonded to each other. Examples are diamond and silicon dioxide. Such substances have similar properties.
- Covalent network lattices:
  - exist as solids and have very high melting points due to the strong covalent bonds between atoms
  - do not conduct electricity because they have no free electrons
  - are hard and brittle
  - are chemically inert
  - are insoluble in water.
- Covalent layer lattices, such as graphite, are made up of many atoms held strongly in two-dimensional layers by covalent bonds. These layers are held together by weak dispersion forces.
- Graphite:
  - is solid with a high melting point due to strong covalent bonds
  - conducts electricity due to the presence of delocalised electrons
  - is slippery to the touch because weak dispersion forces between layers allow the layers to slide over each other
  - has a metallic sheen due to the interaction between light and the delocalised electrons.
- Diamonds are used for cutting, sawing and drilling through hard materials where their hardness is a required property. They are also used in jewellery for their ability to refract light.
- Graphite has many uses. Most of these depend on its electrical conductivity and the ability of its layers to slide easily over one another. Typical of these uses are as the ‘lead’ in pencils, as inert electrodes and as a lubricant.

Carbon nanomaterials

- Diamond, graphite, graphene and fullerenes are allotropes of carbon and have very different properties and uses.

- Fullerenes, such as buckyballs, are spherical molecules of carbon that are shaped like a soccer ball. Each carbon is covalently bonded to three other carbon atoms. The most common type of buckyball is C_{60}.
- Carbon nanotubes consist of one or more layers of carbon atom rings rolled into a sheet. They are good conductors and strong.
- Graphene consists of a single layer of carbon atoms arranged similarly to graphite but only one atom thick. It is two dimensional and almost transparent and is also an excellent conductor and stronger than steel.
- Uses for carbon nanomaterials include as carbon fibres in lightweight but strong composite materials; for example, in high-quality sports equipment and in the transport industry. They are also used in medicine for the repair of joints and are being investigated for the targeted delivery of drugs and medicine. New uses for these materials are constantly being developed.

A review of bonding

- Comparison of networks:
  - metallic: cations and electrons, high melting point, hard, malleable, ductile, good conductor in solid or liquid state
  - ionic: cations and anions, high melting point, brittle, non-conductor as solid, conductor in liquid state
  - covalent molecular: molecules, low melting point, soft, non-conductor in any state
  - covalent network: atoms, very high melting point, hard, non-conductor
  - covalent layer: atoms, very high melting point, soft, conductor.

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7.5.2 Key terms

allotrope different forms of an element
carbon nanotube structure in the shape of a tube composed of carbon atoms
covalent bonding bonding between non-metal atoms that involves electron sharing
covalent layer lattice a substance that displays covalent bonding in three dimensions without the formation of discrete molecules
delocalised describes electrons that are not bound to any one atom but are free to move throughout a lattice
fullerenes nanoscale allotropes of carbon, characterised by having some type of hollow shape
Mohs scale a scale used to measure the hardness of substances
nanomaterial a material that has at least one dimension from 1 to 100 nanometres
network lattice a substance that displays covalent bonding in three dimensions without the formation of discrete molecules
organic chemistry the study of carbon-containing compounds and their properties
refractive index a measure of how much a substance refracts (bends) light when it enters it

on Resources

Digital document Key terms glossary — Topic 7 (doc-30950)

7.5 Exercises

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7.5 Exercise 1: Multiple choice questions

1. Why is diamond an electrical insulator?
   - **A.** It is a crystalline substance.
   - **B.** All the outershell electrons are delocalised.
   - **C.** It consists of non-conducting carbon atoms only.
   - **D.** All the outershell electrons are involved in single covalent bonds.
2. Crystalline diamond, C, and crystalline quartz, SiO_2, are alike in some respects. Which one of the following statements is *incorrect*?
   - **A.** Both substances are above average in hardness.
   - **B.** At room temperature, neither substance dissolves in water.
   - **C.** In pure form, both substances may be colourless and transparent.
   - **D.** Diamond has covalent bonds between carbon atoms, and quartz has covalent bonds between silicon atoms only.
3. The addition of powdered graphite to a lock that has been sticking often allows the key to turn more easily. Why does graphite have this effect?
   - **A.** Each carbon atom has already formed four bonds, and hence can form an unreactive film over which metal can slide.
   - **B.** Its infinite lattice structure is very hard and strong, enabling moving parts to slide over a thin layer of graphite.
   - **C.** The tetrahedral arrangement of bonds around each carbon atom causes neighbouring carbon atoms to slide past one another.
   - **D.** The carbon atoms are bonded into two-dimensional sheets that can slide freely over one another.
4. What makes carbon unique?
   - **A.** Carbon has two allotropes.
   - **B.** Carbon forms four bonds.
   - **C.** Carbon forms covalent compounds.
   - **D.** Carbon can bond to itself to form straight chains, branched chains and rings more than any other element.
5. Diamond, graphite and graphene are all
   - **A.** isotopes of carbon.
   - **B.** isomers of carbon.
   - **C.** allotropes of carbon.
   - **D.** compounds of carbon.
6. How does the structure of nanotubes make them suitable as catalysts?
   - **A.** They have a large surface area.
   - **B.** They are made from reactive carbon atoms.
   - **C.** They have strong covalent bonds.
   - **D.** They are tubular in shape.
7. How can graphite be described?
   - **A.** It is made of carbon and hydrogen atoms only.
   - **B.** It is a metal.
   - **C.** It is hard and crystalline.
   - **D.** It is soft and slippery.

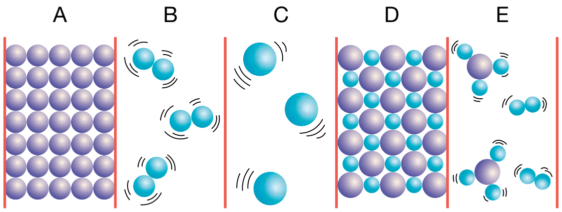
8. Why might fullerenes be used in new drug delivery systems?
   - **A.** They are made from carbon atoms.
   - **B.** They are hollow.
   - **C.** They are very strong.
   - **D.** They are allotropes of carbon.
9. Which of the following sets of substances have only covalent bonds?
   - **A.** Copper(II) oxide, ammonia, methane and silicon dioxide
   - **B.** Copper(II) oxide, sodium chloride, calcium fluoride and hydrogen chloride
   - **C.** Ammonia, methane, silicon dioxide and hydrogen chloride
   - **D.** Ammonia, calcium fluoride, methane and silicon dioxide
10. Why do metals conduct electricity in the solid state?
   - **A.** They contain many mobile electrons.
   - **B.** They contain equal numbers of cations and electrons.
   - **C.** They contain many mobile anions.
   - **D.** They contain equal numbers of cations and anions.
11. Which of the following substances is the best electrical conductor when molten?
   - **A.** Diamond
   - **B.** Carbon tetrachloride
   - **C.** Potassium chloride
   - **D.** Silicon dioxide
12. The electron configurations of elements *P, Q, R* and *S* are given below.

   P: 2, 8, 2
   Q: 2, 8, 6
   R: 2, 7
   S: 2, 8, 8

   Which one of the following pairs of elements is most likely to react to form
   - **i.** an ionic compound?
   - **ii.** a covalent compound?
   - **A.** *P* and *S*
   - **B.** *P* and *Q*
   - **C.** *Q* and *R*
   - **D.** *Q* and *S*
13. Chemical bonds form for all of the following reasons *except*
   - **A.** a tendency to achieve a more stable electron configuration.
   - **B.** attractions between nuclei and electrons.
   - **C.** a tendency for atoms to fill their valence shell with electrons.
   - **D.** a tendency of bonded atoms to react more readily than unbonded atoms.
14. A difference between ionic and molecular compounds is that ionic compounds
   - **A.** dissolve in water, but molecular compounds are insoluble.
   - **B.** usually form crystals, whereas molecular compounds usually do not.
   - **C.** conduct electricity in the solid state, whereas molecular compounds usually do not.
   - **D.** usually melt at higher temperatures than molecular compounds.
15. In which of the following substances would you expect the bonding to be most ionic?
   - **A.** Ice
   - **B.** Solid ammonia
   - **C.** Solid lead bromide
   - **D.** Solid silicon dioxide

7.5 Exercise 2: Short answer questions

1. **a.** How many millimetres are in a metre?
   b. How many micrometres are in a metre?
   c. How many nanometres are in a metre?
2. **a.** Discuss the structure of diamond and graphite.
   b. What are the similarities and differences in these two structures?
   c. Why are they called ‘giant’ molecules?
   d. Provide three uses of graphite and three uses of diamond.
3. A zipper is rubbed with a pencil to make it move freely.
   a. What property of graphite is this illustrating?
   b. Why does graphite have this property?
4. What property of diamond enables it to be used to cut glass?
5. Diamonds are brittle. Discuss this.
6. When wax from a candle vaporises, some molecules burn in the tip of the flame, producing soot that contains buckyballs. Describe the structure of the C_{60} buckyball.
7. An element *X*, the atoms of which contain 12 protons, is chemically bonded to another element, *Y*, the atoms of which contain 9 protons.
   a. What type of bonding would be present between *X* and *Y*?
   b. What is the formula of the compound formed from *X* and *Y*?
8. The following figure shows various arrangements of atoms, molecules or ions. Which of these diagrams represents (a) a noble gas, (b) a mixture of gases, (c) a metallic solid, (d) an ionic solid and (e) hydrogen gas?



9. Briefly discuss the bonding in the following substances.
   a. $CaCl_2$
   b. CCl_4
   c. MgO
   d. H_2O
   e. CO_2
   f. He
   g. C_2H_2
   h. HNO_3
   i. H_2S
   j. $CaCO_3$
   k. KOH
   l. 18 carat gold
10. Discrete molecules (simple molecular substances) often have a smell. Metallic, ionic and covalent network molecules do not have a smell. Discuss.
11. An experiment is conducted in order to classify some compounds. Each compound is melted and then its electrical conductivity is tested.

Compound	Electrical conductivity
Benzoic acid	Poor
Magnesium sulfate	Good
Sodium fluoride	Good
Paraffin wax	Poor
Aluminium	Good
Sugar	Poor

- **a.** **i.** Use the electrical conductivity data to classify the bonding for each substance as ionic, metallic or covalent molecular.
  - **ii.** Would any of the substances conduct electricity in the solid state? Explain.
  - **iii.** What types of bond or force are breaking when these substances melt?
  - **iv.** Suggest whether you expect the melting point of each substance to be high or low. Give reasons for your answers.
- **b.** Explain why some substances tested did not conduct electricity in the molten state.
- **c.** Testing the electrical conductivity of a substance when it is added to water is *not* a reliable method of classification. Give reasons to explain why.

12. Diamond and tungsten both have extremely high melting points. Contrast the bonding and structure of each and explain why they both have a high melting point.

13. Some of the properties of the pure substances *E, F, G* and H are given in the following table.

Substance	Melting point (°C) of solid	Electrical conductivity		
		of solid	of liquid	of solution
E	−110	Nil	Nil	Nil
F	21	Nil	Nil	(Insoluble)
G	810	Nil	High	High
H	1640	High	High	(Insoluble)

- **a.** Which substance(s) could have ionic bonding?
- **b.** Which substance(s) could have metallic bonding?
- **c.** Which substance(s) could have covalent bonding?

14. Classify the following substances as metals, ionic substances, covalent network substances or covalent molecular substances.

- Substance *A* has a low melting point and a negligible electrical conductivity in both the solid and molten states.
- Substance *B* has a high melting point and a high electrical conductivity in both the solid and molten states.
- Substance *C* has a high melting point. It does not conduct electricity in the solid state, but the molten state has high electrical conductivity.
- Substance *D* has a very high melting point and a negligible electrical conductivity in both the solid and molten states.

15. With reference to structure and properties, discuss why each of the following materials is used in the application given.

- **a.** Carbon fibre in aircraft construction
- **b.** Aluminium in drink cans
- **c.** Gold in electrical contacts for computer parts
- **d.** Graphite in batteries

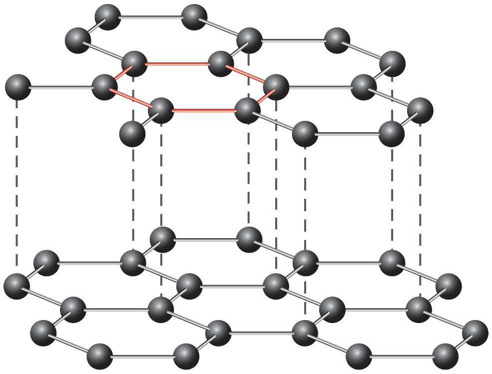
7.5 Exercise 3: Exam practice questions

Question 1 (3 marks)

Silica, SiO_2, is a solid with a high melting point, while CO_2 at room temperature is a gas. Both silicon and carbon are found in group 14 of the periodic table, and hence you would expect similar properties in their compounds. Account for the differences in physical states of SiO_2 and CO_2 in terms of the structures of the two substances.

Question 2 (4 marks)

Graphite (shown) and graphene consist of continuous rings of carbon atoms. Why are their properties so different?



Question 3 (7 marks)

When diamond is heated to high temperature, it burns to form carbon dioxide. If oxygen is excluded, however, it changes to graphite, which then sublimes at about 3500 °C. Three other substances, silicon, a form of silicon carbide, SiC, called moissanite and germanium are also able to resist very high temperatures.

a. Describe the type of bonding that you would expect in each of the substances mentioned. **1 mark**

b. Would you expect germanium to be hard or soft? Explain. **2 marks**

c. Samples of these four substances are labelled A, B, C and D. In an experiment, they are subjected to extreme heat in the absence of oxygen. It is observed that D melts before B. At the conclusion of the experiment it is noted that A has not melted and C has turned a dark grey colour. Identify which substance is which and explain your reasoning. **4 marks**

Question 4 (6 marks)

Consider each of the following solids: Na, Si, H_2S, He, HF, KF. Which would be an example of the following?

a. A solid in which hydrogen bonding exists between molecules **1 marks**

b. A solid that is a good electrical conductor **1 marks**

c. A solid that is a poor electrical conductor, but conducts on melting **1 marks**

d. A solid consisting of atoms held together by weak dispersion forces **1 marks**

e. A solid in which the atoms are covalently bonded together into a network lattice **1 marks**

f. A solid containing a V-shaped, polar molecule **1 marks**

7.5 Exercise 3: studyON Topic Test onlineonly

Fully worked solutions and sample responses are available in your digital formats.

teachon

Test maker

Create unique tests and exams from our extensive range of questions, including practice exam questions.
Access the Assignments section in learnON to begin creating and assigning assessments to students.

AREA OF STUDY 2
HOW CAN THE VERSATILITY OF NON-METALS BE EXPLAINED?

8 Organic compounds

8.1 Overview

Numerous **videos** and **interactivities** are available just where you need them, at the point of learning, in your digital formats, learnON and eBookPLUS at www.jacplus.com.au.

8.1.1 Introduction

FIGURE 8.1 Barbeques are fuelled by simple hydrocarbons.

More compounds of carbon exist than compounds of all the other elements combined. Organic chemistry is the study of these compounds. Historically, scientists thought that these carbon compounds could only be made by living organisms — hence the name *organic*. Today, however, we know that this is not the case and we use myriad organic compounds in everyday life. Some of these are natural, some are processed from natural materials and some are totally synthetic.

Some of the most familiar and widely used organic compounds are also the simplest. Transport relies heavily on hydrocarbons such as petrol and diesel, derived from crude oil and among the simplest of all organic compounds. At home, space heaters and gas barbeques run on either natural gas or bottled gas, both of which are even simpler hydrocarbons. Solid fuel barbeques often use products to light them that contain kerosene and paraffin. These are also organic products derived from crude oil.

This topic introduces how chemists sort this vast catalogue of compounds into meaningful groups or families. You will learn about the different types of hydrocarbons and use the concept of functional groups to extend this to alcohols, carboxylic acids and esters. A systematic way of naming compounds based on their structure will also be introduced. And you will learn about the refining of crude oil and how molecular formulas may be determined from experimental evidence.

This topic relies of a number of previously used concepts and skills. Among these are covalent molecular bonding, intermolecular forces and methods for representing molecules introduced in topic 6. The derivation of empirical formulas from topic 5 will be extended to the calculation of molecular formulas.

8.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- the origin of crude oil and its use as a source of hydrocarbon raw materials
- the grouping of hydrocarbon compounds into families (alkanes, alkenes, alkynes, alcohols, carboxylic acids and non-branched esters) based upon similarities in their physical and chemical properties, including

general formulas, their representations (structural formulas, condensed formulas, Lewis structures), naming according to IUPAC systematic nomenclature (limited to non-cyclic compounds up to C10, and structural isomers up to C7) and uses based upon properties
- determination of empirical and molecular formulas of organic compounds from percentage composition by mass and molar mass.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of learning and assessment. Experiments and investigations, supported by a **Practical investigation logbook** and **Teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

on Resources

Digital documents Key science skills (doc-30903)
Key terms glossary — Topic 8 (doc-30952)
Practical investigation logbook (doc-30953)

studyon

To access key concept summaries and practice exam questions download and print the **studyON: Revision and practice exam question booklet** (doc-30954).

8.2 Grouping hydrocarbons into families

KEY CONCEPT

- The grouping of hydrocarbon compounds into families (alkanes, alkenes, alkynes) based upon similarities in their physical and chemical properties, including general formulas, their representations (structural formulas, condensed formulas, Lewis structures), naming according to IUPAC systematic nomenclature (limited to non-cyclic compounds up to C10, and structural isomers up to C7) and uses based upon properties

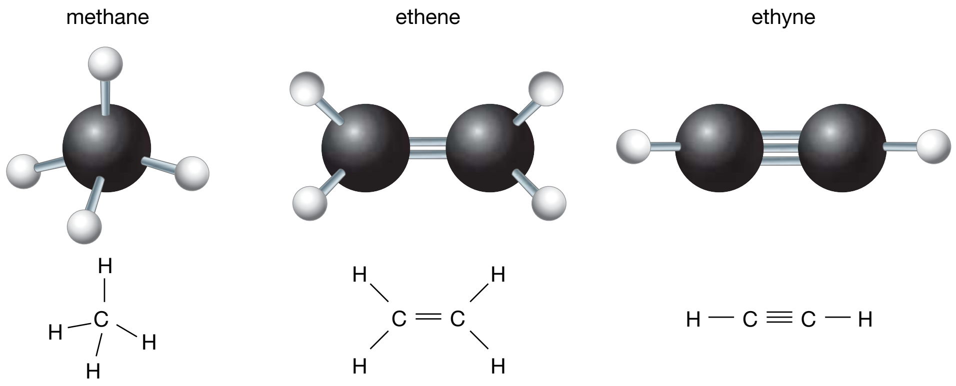
8.2.1 What is organic chemistry?

Organic chemistry is the systematic study of the compounds of carbon, excluding oxides of carbon, carbonates, carbides and cyanides. These exceptions plus the compounds of the remaining elements are considered to be inorganic. Carbon compounds comprise over 90% of all known chemicals and are the only elements with properties that make the development of all living systems possible.

The main reason for carbon's unique ability to form a wide range of chemicals is that carbon–carbon bonds are strong. Carbon normally forms four covalent bonds. These strong bonds form bonds with other non-metals. Carbon can also chemically bond with itself using single, double or triple bonds to form long chain-like structures and even rings.

FIGURE 8.2 Buckminsterfullerene, a soccerball-shaped organic molecule used for research in the pharmaceutical industry; 'Buckyballs' such as these are produced at the tip of a candle flame, where C—C bonds form in ring structures.

FIGURE 8.3 Carbon can form single, double and triple bonds with itself. The simplest compounds of hydrogen and carbon with carbon–carbon double and triple bonds are ethene and ethyne respectively.

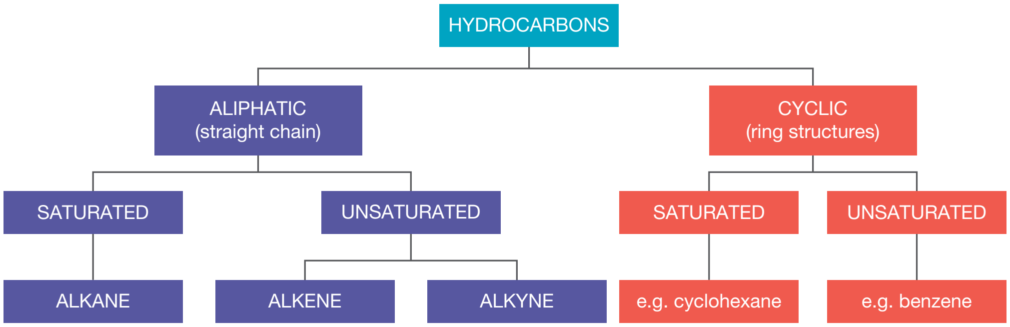


8.2.2 Introduction to hydrocarbons

A good place to begin a study of organic chemistry is with **hydrocarbons**. Hydrocarbons are molecules made up of only hydrogen and carbon. These are the simplest organic compounds and yet are among the most useful. Hydrocarbons are found in crude oil, and processing the oil yields many useful products. This process and the products that are formed are covered in more detail later in this topic.

There are different classes of hydrocarbons, based on the carbon bonding involved. These are indicated in figure 8.4.

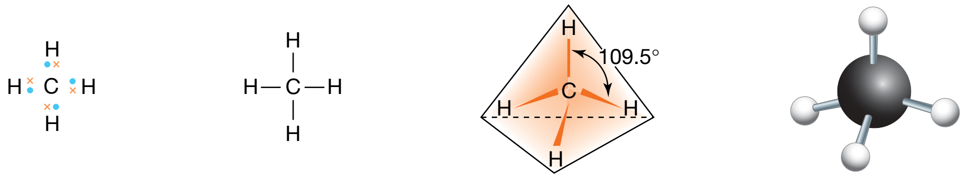
FIGURE 8.4 The classes of hydrocarbons



Saturated hydrocarbons contain only single carbon–carbon bonds. **Unsaturated hydrocarbons** contain double or triple carbon–carbon bonds. In **cyclic compounds**, the carbon atoms are arranged in a ring.

The simplest hydrocarbon is a carbon atom bonded to four hydrogen atoms. Carbon can form four equivalent bonds that are 109.5° apart, forming a tetrahedral shape. This molecule is called methane. The structure of a methane molecule was described in topic 6. Figure 8.5 demonstrates different representations of methane. Of these, structural formulas (and condensed structural formulas where appropriate) are the most commonly used representations in organic chemistry.

FIGURE 8.5 Lewis (electron dot) structure, structural formula, shape, and ball-and-stick model of methane



General properties and uses of hydrocarbons

Compounds containing C—H bonds are generally insoluble in water and do not react with it. They are more likely to be soluble in non-polar solvents. This is because the electronegativities of carbon and hydrogen atoms are similar. This means that the C—H bond is nearly non-polar and, because of the symmetry of the molecule, hydrocarbon molecules are non-polar.

This non-polar nature of hydrocarbons also explains why the boiling point of hydrocarbons increases with the length of the chain of carbon atoms. Only dispersion forces hold the molecules together, and these forces increase as the size of the molecules increases. The increase in boiling point with size reflects the increased dispersion forces between the molecules. Table 8.1 shows this trend in the **alkanes**. If a molecule is branched instead of a straight chain, this lowers the boiling point. Branches prevent the molecules from coming closer together and so the dispersion forces have less effect.

Hydrocarbons are among the most useful substances in today's society. Alkanes (see section 8.2.3) are widely used as transport fuels. They burn readily and release a large amount of energy in doing so. However, as we will see later in this topic, the crude oil from which they are derived is a finite source. Their use as a fuel also produces carbon dioxide, the levels of which are now building up in the atmosphere. **Alkenes** (section 8.2.4) are also important. The smaller alkenes especially, are widely used in the petrochemical industry as pre-cursors from which many other useful substances are made. They are also the raw materials for the manufacture of a wide range of plastics (see topic 9). The most useful **alkyne** is ethyne (known in everyday use as acetylene), a gas used with oxygen for high temperature cutting and welding.

8.2.3 Alkanes

The alkanes are a family of hydrocarbons containing only **single bonds** between the carbon atoms. Notice how the name of each alkane has the same ending: *-ane*. This is how alkanes are identified. The different prefixes, *meth-*, *eth-*, *prop-*, etc., show how many carbon atoms are present in each carbon chain (see table 8.2). Table 8.1 summarises some common straight chain alkanes and their uses.

TABLE 8.1 Common straight chain alkanes

Formula	Name	Phase	Boiling point	Typical use
CH_4	Methane	Gas	−161.5°C	Natural gas
C_2H_6	Ethane	Gas	−89°C	Refrigerant
C_3H_8	Propane	Gas	−42°C	Liquid petroleum gas
C_4H_{10}	Butane	Gas	−1°C	Manufacture of synthetic rubber
C_5H_{12}	Pentane	Liquid	36.1°C	Solvent
C_6H_{14}	Hexane	Liquid	68°C	Filling for thermometers
C_7H_{16}	Heptane	Liquid	98°C	Testing engine knocking
C_8H_{18}	Octane	Liquid	125.6°C	Automobile petroleum
C_9H_{20}	Nonane	Liquid	151°C	Gasoline
$C_{10}H_{22}$	Decane	Liquid	174.1°C	Gasoline
$C_{16}H_{34}$	Hexadecane	Liquid	286.8°C	Lubricating oil
$C_{20}H_{42}$	Icosane	Solid	343.1°C	Wax candles
$C_{28}H_{58}$	Octacosane	Solid	440°C	Tar

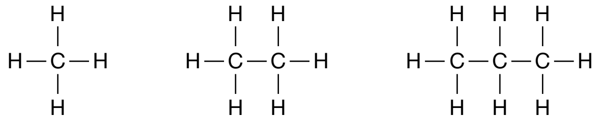
TABLE 8.2 Standard prefixes for indicating numbers of carbon atoms

Number of carbon atoms	Prefix
1	*meth-*
2	*eth-*
3	*prop-*
4	*but-*
5	*pent-*
6	*hex-*
7	*hept-*
8	*oct-*
9	*non-*
10	*dec-*

Any series of organic compounds in which each successive member differs by CH_2 from the previous one is called a **homologous series**. Alkanes can be represented by the general formula C_nH_{2n+2}, where n is the number of carbon atoms in the molecule.

Alkanes are also called saturated hydrocarbons. This refers to the fact that all the carbon to carbon bonds are single bonds. This means that, for a given number of carbons, the maximum number of hydrogens possible is in each molecule.

FIGURE 8.6 Structural and molecular formulas of the first three alkanes



methane, CH_4 ethane, C_2H_6 propane, C_3H_8

A homologous series of organic compounds have the same structure but differ between members by CH_2.

Alkanes are saturated hydrocarbons with single bonds between the carbon atoms, they have the general formula C_nH_{2n+2}.

SAMPLE PROBLEM 1

Write the formula for the alkane that has 12 carbon atoms per molecule.

Teacher-led video: SP1 (tlvd-0539)

THINK	WRITE
Alkanes have the general formula C_nH_{2n+2}. If 12 carbons are present, $n = 12$. The number of hydrogen atoms will be $2n + 2 = 2 \times 12 + 2 = 26$. This molecule is called dodecane.	$C_{12}H_{26}$

PRACTICE PROBLEM 1

Write the formula for the alkane that has 16 carbon atoms per molecule.

Reactions of alkanes

Alkanes burn (combust) in a plentiful supply of oxygen to produce carbon dioxide and water. However, if the supply of oxygen is limited, carbon monoxide or even just carbon may be produced in place of carbon dioxide. Chemists use chemical equations to summarise chemical reactions. Using methane as an example, the reactions mentioned may be summarised as follows.

In a plentiful supply of oxygen

$$\begin{array}{ccccccc} CH_4(g) & + & 2O_2(g) & \rightarrow & CO_2(g) & + & 2H_2O(g) \\ \text{methane} & + & \text{oxygen} & \rightarrow & \text{carbon dioxide} & + & \text{water} \end{array}$$

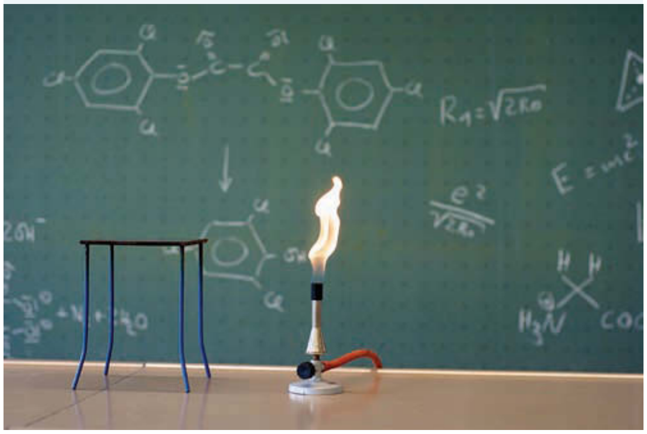
and in a limited supply of oxygen

$$\begin{array}{ccccccc} CH_4(g) & + & 1\tfrac{1}{2}O_2(g) & \rightarrow & CO(g) & + & 2H_2O(g) \\ \text{methane} & + & \text{oxygen} & \rightarrow & \text{carbon monoxide} & + & \text{water} \end{array}$$

or

$$\begin{array}{ccccccc} CH_4(g) & + & O_2(g) & \rightarrow & C(s) & + & 2H_2O(g) \\ \text{methane} & + & \text{oxygen} & \rightarrow & \text{carbon} & + & \text{water} \end{array}$$

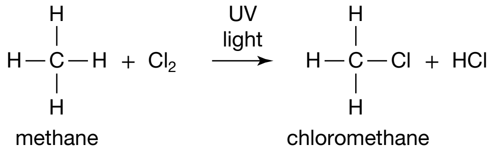
FIGURE 8.7 Methane burning with a limited supply of oxygen produces a yellow flame.



The effect of these reactions can be seen when using a Bunsen burner. When the air hole of a Bunsen burner is open and oxygen is plentiful, methane burns with a blue flame. When the air hole is closed, limiting the supply of oxygen, a yellow flame is the result of the particles of carbon produced (figure 8.7).

Alkanes react with chlorine or fluorine in a reaction called a **substitution reaction**. In this type of reaction, one or more of the chlorine or fluorine atoms take the place of one or more of the hydrogen atoms. The products of these reactions are known as **haloalkanes**. For example, if chlorine is used and enough is present, the hydrogen atoms may be substituted in sequence by chlorine to produce a mixture of CH_3Cl, CH_2Cl_2, $CHCl_3$ and CCl_4. The first step in this process is shown in figure 8.8.

FIGURE 8.8 Methane reacts with chlorine gas under UV light to form chloromethane.



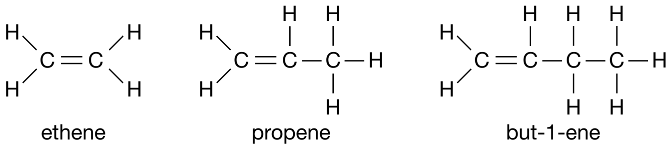
8.2.4 Alkenes

The **alkenes** make up a family of hydrocarbons that each contains a **double bond** *between two carbon atoms*.

The structural formulas of the first three alkenes are shown in figure 8.9. Notice that each of the names ends with *-ene*. This is the characteristic ending used for the alkenes.

The alkenes have a general formula C_nH_{2n} and are another example of a homologous series. The double bond between the carbon atoms reduces the number of hydrogen atoms in an alkene compared to their corresponding alkane.

FIGURE 8.9 Structural formulas of the first three alkenes



Alkenes are hydrocarbons with a double bond between two carbon atoms, they have the general formula C_nH_{2n}.

on Resources

Video eLesson Homologous series of alkenes (eles-2477)

SAMPLE PROBLEM 2

Write the formula for the alkene that has eight carbon atoms per molecule.

Teacher-led video: SP2 (tlvd-0540)

THINK	WRITE
Alkenes have the general formula C_nH_{2n}. If eight carbons are present, $n = 8$. The number of hydrogen atoms will be $2n = 2 \times 8 = 16$. This molecule is called octene.	C_8H_{16}

PRACTICE PROBLEM 2

Write the formula for the alkene that has six carbon atoms per molecule.

Reactions of alkenes

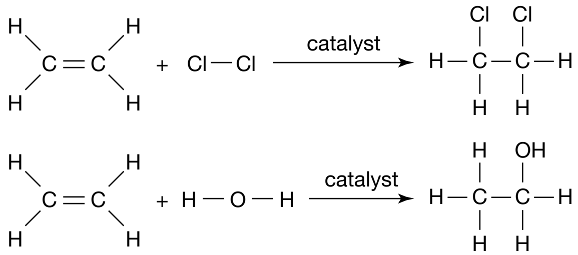
Alkenes undergo combustion reactions with oxygen and, like alkanes, form carbon dioxide and water.

$$\begin{array}{llll} C_2H_4(g) + 3O_2(g) \rightarrow & 2CO_2(g) & + 2H_2O(g) \\ \text{ethene} + \text{oxygen} \rightarrow & \text{carbon dioxide} + & \text{water} \end{array}$$

However, this reaction has little commercial value. Small alkenes such as ethene are far more valuable to the chemical industry as raw materials for the manufacture of other chemicals.

Alkenes can undergo **addition reactions**. In these reactions alkenes react with small molecules such as hydrogen or a halogen. They can also react with water and hydrogen halides. The small molecule attacks and breaks one of the double bonds. In the process, the small molecule is split and is added across the site of the former double bond. With chlorine and water (along with suitable catalysts) the reactions shown in figure 8.10 take place.

FIGURE 8.10 Alkenes break their double bond when they undergo addition reactions.



Alkenes can undergo self-addition in which the alkene molecules join together to make long chains called **polymers**. This will be discussed in topic 9.

Alkenes (and also alkynes) can be referred to as unsaturated hydrocarbons. This is because they contain at least one carbon-to-carbon multiple bond. Consequently, for a given number of carbons, fewer than the maximum number of hydrogen atoms are present per molecule.

8.2.5 Alkynes

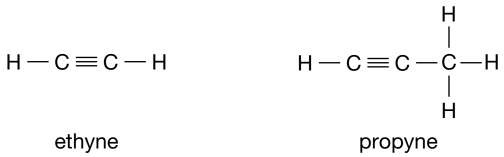
Alkynes are hydrocarbons that contain a **triple bond** *between two carbon atoms*.

Note that the ending for each of the two alkynes shown in figure 8.11 is *-yne*. This ending shows that a triple bond is present in the molecule.

The general formula for the alkynes is C_nH_{2n-2}. The alkynes are another homologous series.

Alkynes are hydrocarbons with a triple bond between two carbon atoms, they have the general formula C_nH_{2n-2}.

FIGURE 8.11 Structural formulas of two alkynes



ethyne

propyne

Reactions of alkynes

Alkynes can undergo combustion (oxidation) and addition reactions. Like all hydrocarbons, the products of the combustion reaction are carbon dioxide and water. The equation for the combustion of ethyne in pure oxygen is:

$$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(g).$$

FIGURE 8.12 The combustion of ethyne (acetylene) is used in oxyacetylene welding. The extremely high temperature generated when ethyne is burned with pure oxygen can melt metal and allow surfaces to be welded together.

CYCLIC HYDROCARBONS

The carbon backbone of some hydrocarbons can form a ring. Such compounds are known as cyclic hydrocarbons and have the prefix *cyclo*- before their name.

An important category of the hydrocarbons is the unsaturated cyclic compounds making up the aromatic series. The simplest member of this series is benzene, C_6H_6. This molecule has been shown to have two equivalent forms, as shown in figure 8.14 (a) and (b).

Benzene is found naturally in crude oil but is usually synthesised from other hydrocarbons when used in making pharmaceuticals.

FIGURE 8.13 Structural formulas for (a) straight chain hexane and (b) cyclic hexane

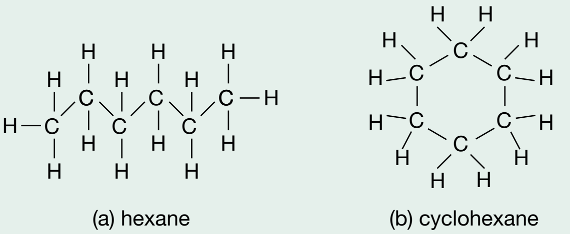
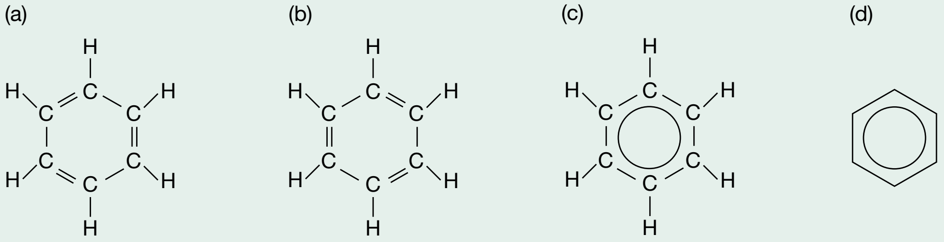


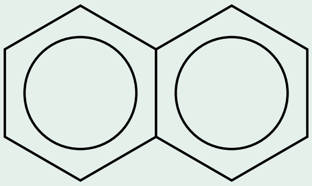
FIGURE 8.14 Four structural representations of benzene



Scientists have shown experimentally that all carbon–carbon double bonds in benzene are identical and somewhere between single and double bonds in length and bond strength. The equivalence of the bonds is shown by the alternative structure shown in figure 8.14 (c). The circle represents the even distribution of the six electrons around the benzene ring. The six electrons are said to be delocalised. Figure 8.14 (d) shows a shorthand representation of benzene.

Given the information provided for benzene, try to work out the molecular formula for napthalene, represented by the structure in figure 8.15.

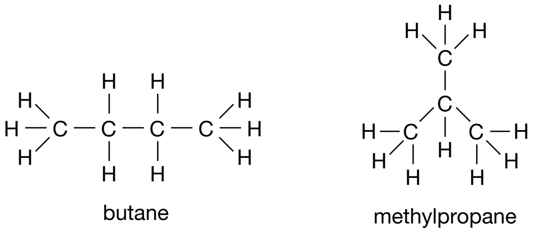
FIGURE 8.15 Structural formula of napthalene



8.2.6 Isomers

Isomers are molecules that have the same molecular formula but have differences in the way the atoms are joined together or arranged. Isomerism is very common in organic chemistry and is one of the main reasons that carbon can form such a huge range of compounds. As the number of carbons in a molecule increases, so too does the number of ways that all the atoms can be joined together. In other words, the number of isomers increases. One particular type of isomerism is structural isomerism. **Structural isomers** have the same molecular formula but different structural formulas. For example, C_4H_{10} can be drawn in two ways, as shown in figure 8.16. When describing structures such as these, the terms ‘straight chain’ and ‘branched chain’ are often used. Straight chain refers to when the carbons are joined in such a way that an unambiguous path exists from the first carbon atom in the structure to the last. Branched chain, on the other hand, refers to where a branch occurs at some stage as you progress along the carbon backbone. In figure 8.16, butane would be described as a straight chain hydrocarbon, whereas methylpropane would be described as a branched chain hydrocarbon.

FIGURE 8.16 The two isomers of C_4H_{10}



on Resources

Video eLesson Isomers (eles-2478)

When considering isomers of hydrocarbons, it is useful to think of carbon atoms attached to the hydrogen atoms as groups. A methyl group is one carbon attached to three hydrogens, —CH_3, and they occur at each end of a straight chain hydrocarbon. As demonstrated in sample problem 3, changing the position of a methyl group produces different isomers of the same molecule. Similarly, a group with two carbon atoms, one attached to two hydrogens and another attached to three hydrogens, —CH_2 —CH_3, is an ethyl group. Methyl and ethyl groups are alkyl groups, and they have one less hydrogen than the alkane of the same name.

Structural isomers have difference names and chemical properties because the molecules have different shapes. In figure 8.16, for example, the only forces between the molecules of each isomer are dispersion forces. These would be about the same strength because the isomers are of similar size. However, molecules of butane can get closer together than can molecules of methyl propane (because butane is a straight chain hydrocarbon), meaning that its dispersion forces are slightly more effective. This is reflected in the slightly higher boiling point for butane (−0.5 °C) compared to methyl propane (−11.7 °C).

SAMPLE PROBLEM 3

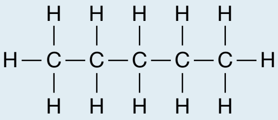
Draw three possible isomers for pentane, C_5H_{12}.

Teacher-led video: SP3 (tlvd-0541)

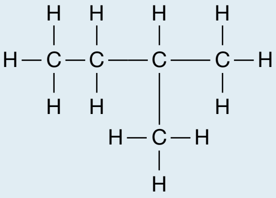
THINK

WRITE

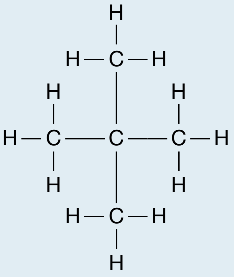
1. All structures must link five carbons to 12 hydrogens. Each hydrogen must only have one bond. Each carbon must have four bonds.
   Start with a straight chain of five C atoms, linked to the 12 H atoms.



2. Move one methyl group (one C and three H) to the third C in the chain. The numbering of the C atoms in the chain does not matter. The —CH_3 group could be moved to the second C in this case, and the molecule would be the same.



3. Move another —CH_3 group to the third C atom in the chain to create the final isomer of pentane.



PRACTICE PROBLEM 3

Draw three possible isomers for hexane, C_6H_{14}.

8.2.7 Naming hydrocarbons

Over the years, as organic chemistry grew and more and more compounds were either discovered or synthesised, naming them increasingly became a problem. To solve this problem, the International Union of Pure and Applied Chemistry (IUPAC) developed a set of rules by which any organic substance could be named depending on the structure of its molecules. This system is in widespread use today, although many common substances still retain their ‘old’ or ‘trivial’ names in everyday use. For example, ethyne, mentioned earlier in this topic, is still called acetylene in everyday use.

Rules for naming hydrocarbons:

- **Rule 1: Determine the longest chain of carbon atoms. This will form the basis of the name.**
- **Rule 2: Determine which end is nearest to a branch, a double bond or a triple bond. (A double or triple bond takes precedence over a branch if they are equidistant from either end of the chain.)**
- **Rule 3: Number the carbon atoms from the end chosen.**
- **Rule 4: Name any branches first with the ending *-yl* (for example, methane becomes *methyl* and ethane becomes *ethyl*), then the longest chain, and then any single or double bond.**
- **Rule 5: When two or more branches occur on the same carbon atom, the number of the carbon atom is indicated for each branch, with the names given in alphabetical order (ignoring the prefixes described in rule 6).**
- **Rule 6: When two or more identical branches occur on different carbon atoms, the prefixes *di-*, *tri-* and *tetra-* are used.**

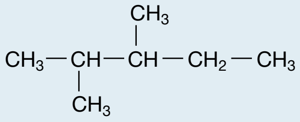
TIP: **Avoid the following common errors when naming hydrocarbons:**

- **not identifying the longest chain possible**
- **not listing the side branches in alphabetical order**
- **omitting the prefixes *di-*, *tri-* and *tetra-* when they are required.**

This system uses a set of standard prefixes by which the number of carbons is indicated. These are referred to as alkyl prefixes; for example, methyl, ethyl and propyl. Table 8.2 (section 8.2.3) shows these prefixes for up to ten carbons.

SAMPLE PROBLEM 4

Use the rules for naming organic compounds to systematically name the following compound.

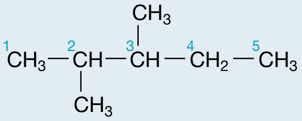


$$CH_3-\underset{\underset{\displaystyle CH_3}{|}}{CH}-\overset{\overset{\displaystyle CH_3}{|}}{CH}-CH_2-CH_3$$

Teacher-led video: SP4 (tlvd-0542)

THINK	WRITE
1. The longest unbranched carbon chain contains five carbon atoms separated by single bonds.	Pentane

2. The carbon atoms of the longest unbranched chain are numbered from left to right, because the branches are closest to the left.



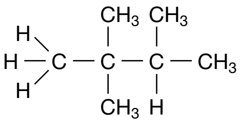
$^{1}CH_3-^{2}CH(CH_3)-^{3}CH(CH_3)-^{4}CH_2-^{5}CH_3$

3. The branching chains are —CH_3, so they are methyls and attached to carbon atoms 2 and 3. The two branches are both —CH_3, so they are named dimethyl.

The name of the molecule is 2,3-dimethylpentane.

PRACTICE PROBLEM 4

Use the rules for naming organic compounds to systematically name the following compound:



SAMPLE PROBLEM 5

Draw the structure for 4-ethyl-2-methylhex-1-ene.

Teacher-led video: SP5 (tlvd-0543)

THINK

WRITE

1. The prefix 'hex' indicates the molecule will have six carbons.

$C-C-C-C-C-C$

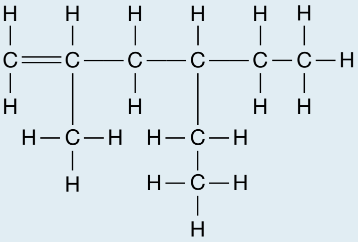
2. The 1-ene indicates the molecule is an alkene and a double bond exists between carbons 1 and 2.

$C{=}C-C-C-C-C$

3. Number the carbon chain from the double bond end.

$^{1}C{=}^{2}C-^{3}C-^{4}C-^{5}C-^{6}C$

4. 4-ethyl-2-methyl indicates an ethyl group (—CH_2—CH_3) to carbon number 4 and a methyl group (—CH_3) to carbon number 2. Ensure sufficient hydrogen atoms have been added to all C atoms.

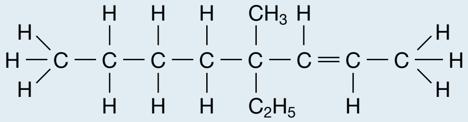


PRACTICE PROBLEM 5

Draw the structure for a molecule of 4-ethyl-2,3,3-trimethylheptene.

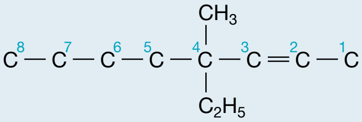
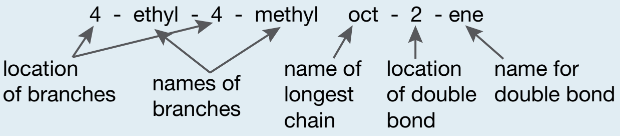
SAMPLE PROBLEM 6

Name the following compound.



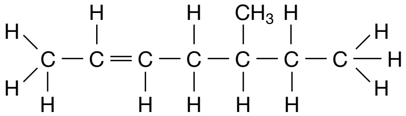
$CH_3CH_2CH_2CH_2C(CH_3)(C_2H_5)CH{=}CHCH_3$

Teacher-led video: SP6 (tlvd-0544)

THINK	WRITE
1. The longest chain has eight carbons, so the *oct* name is given. The double bond is nearest to the right end, so number accordingly (starting from the right). The double bond on the carbon atom number 2 is represented as *2-ene*.	Oct-2-ene $\overset{8}{C}-\overset{7}{C}-\overset{6}{C}-\overset{5}{C}-\overset{4}{C}-\overset{3}{C}=\overset{2}{C}-\overset{1}{C}$
2. The upper branch on carbon atom number 4 contains one carbon atom, so it is given the name *4-methyl*. The lower branch contains two carbon atoms, so it is named *4-ethyl*.	 $\overset{8}{C}-\overset{7}{C}-\overset{6}{C}-\overset{5}{C}-\overset{4}{C}(CH_3)(C_2H_5)-\overset{3}{C}=\overset{2}{C}-\overset{1}{C}$
3. The two branches on carbon atom number 4 both have the number indicated and are written alphabetically as *4-ethyl-4-methyl*.	 The name of the compound is 4-ethyl-4-methyloct-2-ene.

PRACTICE PROBLEM 6

Name the following compound.



$CH_3CH{=}CHCH_2CH(CH_3)CH_2CH_3$

on Resources

Video eLesson Naming alkanes (eles-2484)

8.2.8 Representing organic molecules

Organic molecules can be represented using Lewis structures, structural formulas and condensed formulas, among others.

Lewis structures

Lewis structures (electron dot formulas) were introduced in topic 6. Dashes (—) can be used to replace each pair of bonding electrons but lone pairs are often left as dots. There may or may not be an attempt to show the actual shape of the molecule.

The following steps outline how to draw a Lewis structure, with figures 8.17 and 8.18 showing how the Lewis structure may be drawn for water, H_2O, and methanal, H_2CO respectively.

1. Add up the total number of valence electrons in the molecule.
2. Place the central atom(s) in position. This will *usually* be the least electronegative atom, or might be the one with the fewest number of valence electrons. Note that this is only a guide and exceptions to this can occur.
3. Connect all atoms with single bonds as dashes (—) or as shared electron pairs (•×).
4. Add electron dots to the peripheral atoms as *lone pairs* to make octets. Remember that hydrogens do not have lone pairs.
5. Place remaining electrons as lone pairs around the central atom(s).
6. Check for octets around the central atom(s). If they are not present, use multiple bonds and adjust lone pairs accordingly. Continue until all octets are generated.
7. If after Step 6 electrons are still left over, add them to the central atom, even if this violates the octet rule. Remember that the octet rule is a guide only and can the broken for atoms in the third period and beyond.

FIGURE 8.17 Drawing the Lewis structure for water, H_2O

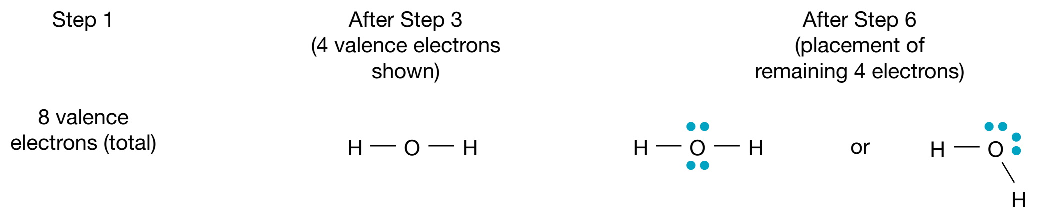
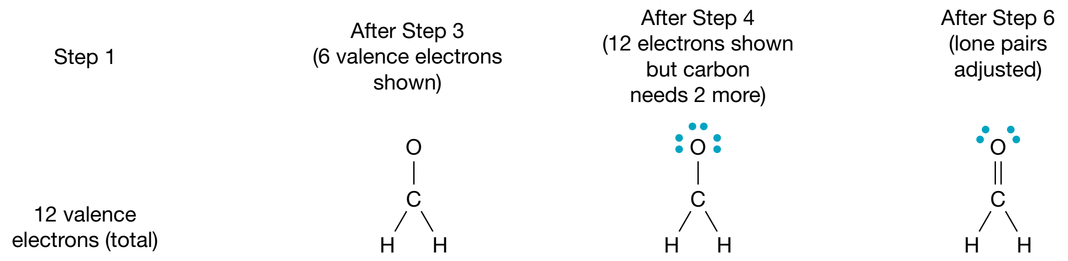


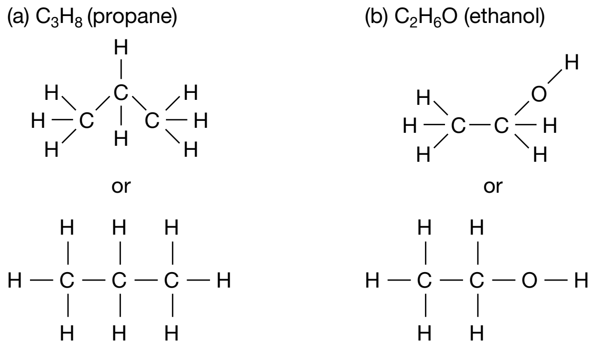
FIGURE 8.18 Drawing the Lewis structure for methanal, H_2CO



Structural formulas

As you saw in topic 6, the **structural formula** of a substance shows the way the atoms in a molecule are connected; that is, it shows the bonds. For small molecules, there may be an attempt to show shape although this is not always necessary. Figure 8.19 shows the structural formulas for propane and ethanol.

FIGURE 8.19 Structural formulas for (a) propane and (b) ethanol



(a) C_3H_8 (propane)

(b) C_2H_6O (ethanol)

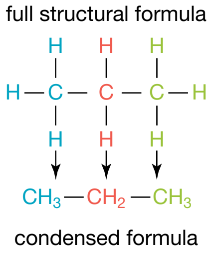
on Resources

Video eLesson Molecular representations of butan-1-ol (eles-2485)

Condensed formulas

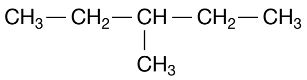
As the name suggests, a **condensed formula** (also known as a *semi-structural formula*), is an abbreviated form of the structural formula. It is frequently used in organic chemistry, especially as molecules become larger. A condensed formula avoids the tedious task of showing all bonds and all atoms by grouping sets of atoms together. It shows an abbreviation of the structural formula, *on a single line*. For example, ethane, C_2H_6, can be represented by the condensed formula CH_3CH_3. The condensed formula of propane, C_3H_8, is $CH_3CH_2CH_3$. When abbreviating a structural formula to its condensed formula equivalent, go ‘carbon by carbon’.

FIGURE 8.20 Condensing the full structural formula of propane into the condensed formula



Brackets are used:

- to indicate side chains, which are written after the carbon to which they are attached — for example:

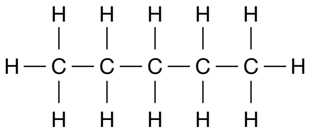


$$CH_3-CH_2-\underset{\displaystyle CH_3}{\underset{|}{CH}}-CH_2-CH_3$$

3-methylpentane

$CH_3CH_2CH(CH_3)CH_2CH_3$

- if there are repeating CH_2 groups — for example:



pentane

$CH_3CH_2CH_2CH_2CH_3$ or $CH_3(CH_2)_3CH_3$

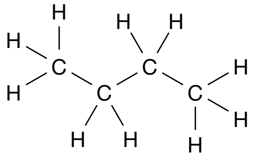
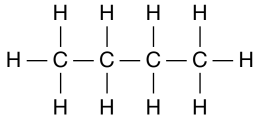
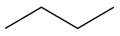
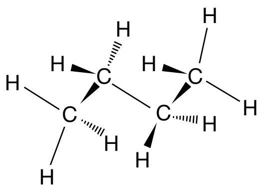
Other formula types

Three other commonly used types of formulas are **empirical formulas**, **molecular formulas**, **skeletal formulas** and three-dimensional structural formulas.

Empirical formulas were introduced in topic 5. Their extension to molecular formulas is discussed later in this topic. Skeletal formulas are particularly useful when dealing with large, complicated molecules. These use lines and vertices to simplify a structural formula. It is assumed that a carbon atom (and enough hydrogens to satisfy carbon's valency) is present at each vertex (and also at the ends). Double bonds and other different types of atoms are specifically shown. These will be encountered further in units 3 and 4. Three-dimensional structural formulas attempt to show the three dimensional shape using wedges and dotted lines, as discussed in topic 6.

Table 8.3 shows the different ways a butane molecule can be shown using these different types of formulas.

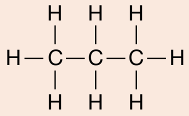
TABLE 8.3 Different ways of representing butane molecules

Formula	Example	Notes
Molecular	C_4H_{10}	This shows the number and kinds of atoms in a molecule.
Empirical	C_2H_5	This shows the simplest whole number ratio of atoms in a molecule.
Structural	 or 	This shows the actual arrangement of atoms in a molecule. As molecules become longer, the second example of a structural formula tends to be used.
Semi-structural or condensed	$CH_3CH_2CH_2CH_3$ or $CH_3(CH_2)_2CH_3$	These can be written on a single line, with each carbon atom being followed by the atoms that are joined to it. Repeated CH_2 groups can be collected together in brackets with a subscript as shown.
Skeletal		A carbon atom is present at each join and at the ends of the line. Enough hydrogen atoms are present to complete the octet of electrons around each carbon atom. This representation is usually used for more complex molecules.
3D structural or shape diagram		This shows the 3D arrangement of atoms, with the bonds represented as follows: • The continuous line is in the plane of the paper. • The dashed line extends to the back of the paper. • The solid wedge comes out of the plane of the paper.

8.2 EXERCISE

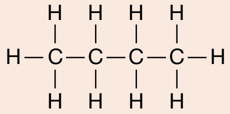
To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Write the formula for the alkane that has 10 carbon atoms per molecule.
2. Write the formula for the alkene that has 13 carbon atoms per molecule.
3. Draw Lewis (electron dot) structures for methane, ethane and ethene.
4. What are the products for the combustion of propane in the following circumstances?
   (a) A plentiful supply of air
   (b) A limited supply of air
5. State two types of reactions that are undergone by
   (a) alkanes.
   (b) alkenes.
6. As the molecular size of the alkanes increases, a change occurs from the gaseous phase to the solid phase. Explain why this happens.
7. Which of the following compounds would you expect to have the higher boiling point? Give reasons for your answer.
   (a)



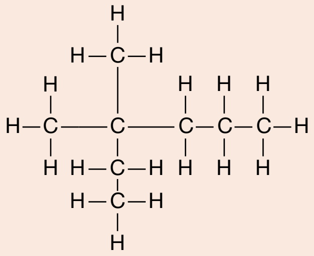
propane, C_3H_8

   (b)



butane, C_4H_{10}

8. Hexane and 2,3-dimethylbutane both have the formula C_6H_{14}. Their boiling points are –95.3 °C and –128.6 °C. Explain with reference to their structure and bonding why their boiling points are so different.
9. Draw two possible isomers for butane, C_4H_{10}.
10. For each of the following compounds labelled (i), (ii) and (iii), complete the following.
    (a) How many isomers are there?
    (b) Draw the isomers from part (a).
        i. Pentane, C_5H_{12}
        ii. Butene, C_4H_8
        iii. Propyne, C_3H_4
11. Explain the following.
    (a) Why are numbers not required to correctly name methylbutane?
    (b) Why is the name propan-3-ol incorrect?
12. What is the systematic name for the following compound?



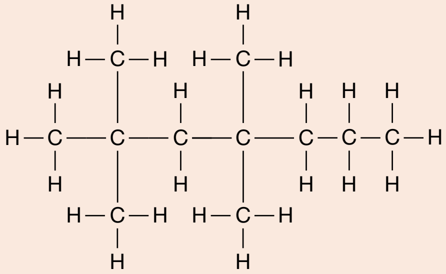
13. Draw the structure for a molecule of 4-ethyl-2,3-dimethylheptane.
14. Bromine can react with both alkanes and alkenes (although the conditions necessary for reaction are different).
    (a) Write the structural formula of a product that is formed when bromine, Br_2, reacts with ethane.
    (b) Write the structural formula of the product that forms when bromine, Br_2, reacts with ethene.
    (c) What types of reaction is each of these?

15. (a) Write structural formulas for the following.

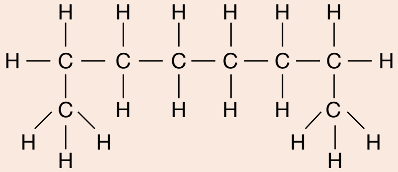
i. Hex-2-ene
ii. Methlpropene
iii. Pent-2-yne

(b) Name each of the following.

i.



ii.



(c) Write condensed formulas for each substance from parts a. and b.

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8.3 Functional groups

KEY CONCEPT

- The grouping of hydrocarbon compounds into families (alcohols, carboxylic acids and non-branched esters) based upon similarities in their physical and chemical properties, including general formulas, their representations (structural formulas, condensed formulas, Lewis structures), naming according to IUPAC systematic nomenclature (limited to non-cyclic compounds up to C10, and structural isomers up to C7) and uses based upon properties

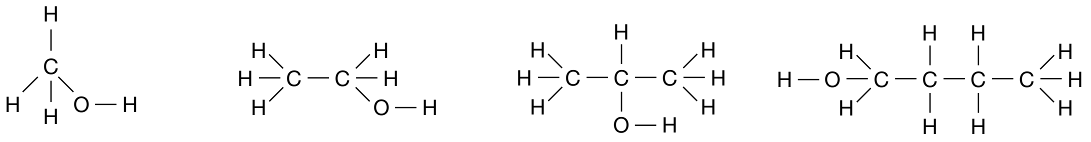
8.3.1 Functional groups

Chemists use the concept of **functional groups** to simplify and assist with the classification of organic compounds. This allows them to reduce the large number of different compounds into more manageable groups that share common properties. A functional group is the bond, atom or group of atoms that give a molecule its specific properties, and its name. For example, the functional group of the alkenes is C═C; the functional group of the alcohols is —OH. The oxygen–hydrogen bond in alcohols and carboxylic acids gives them significantly different properties from those of hydrocarbons.

8.3.2 Alcohols

Alcohols are carbon chains containing one or more —OH groups. The —OH group is called the 'hydroxy' group. Names of alcohols have the ending *-ol*.

FIGURE 8.21 The first four alcohols shown as their structural formula and their semi-structural formula. Common uses are also listed.



CH_3OH

methanol — solvent for varnishes, windscreen washer fluid

CH_3CH_2OH

ethanol — component of alcoholic beverages

$CH_3CH(OH)CH_3$

propan-2-ol — rubbing alcohol, solvent

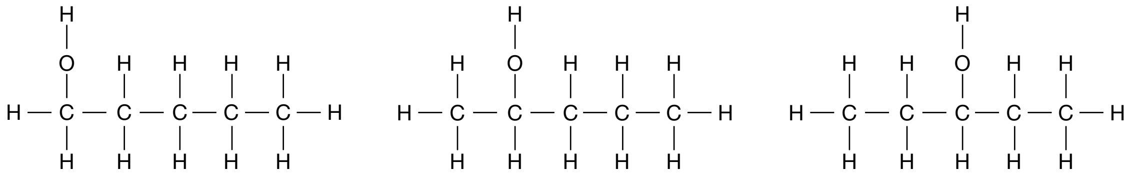
$CH_3CH_2CH_2CH_2OH$

butan-1-ol — solvent, hydraulic fluid

Alcohols are named on the basis of the alkane that they resemble. The *-e* on the end is dropped and replaced by *–ol*; for example, ethanol is the alcohol that resembles ethane. For instances where there is more than one location to attach the —OH group, the number of the carbon to which it is attached must be specified. The general formula for an alcohol is $C_nH_{2n+1}OH$.

Figures 8.21 and 8.22 show further examples of how alcohols are named.

FIGURE 8.22 The structural and semi-structural formulas of the isomers of pentanol, $C_5H_{12}O$



$CH_3CH_2CH_2CH_2CH_2OH$

pentan-1-ol

$CH_3CH(OH)CH_2CH_2CH_3$

pentan-2-ol

$CH_3CH_2CH(OH)CH_2CH_3$

pentan-3-ol

Alcohols contain the hydroxy functional group, —OH. They have the general formula $C_nH_{2n+1}OH$.

Properties, reactions and uses of alcohols

In the smaller alcohols, the hydrogen bonding between the alcohol molecules is stronger than the dispersion forces between hydrocarbon molecules in alkanes, alkenes and alkynes. This results in higher boiling points than the corresponding alkane or alkene. More energy is need to separate methanol molecules than methane molecules.

In addition, the boiling points of alcohols increase with molecular size as the non-polar portion of the molecule increases; this diminishes the effect of hydrogen bonding but increases the influence of the dispersion forces.

Smaller alcohols are soluble in water; hydrogen bonds form between the water molecules and the alcohol molecules (figure 8.23). Solubility decreases with increasing size due to the increasing non-polar section of the alcohol molecules (figure 8.24).

FIGURE 8.23 Smaller alcohols are soluble in water due to hydrogen bonds that form between the alcohol and water molecules.

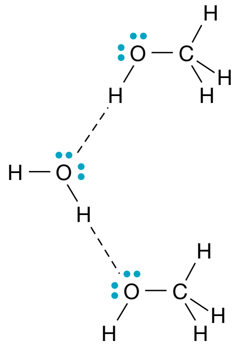
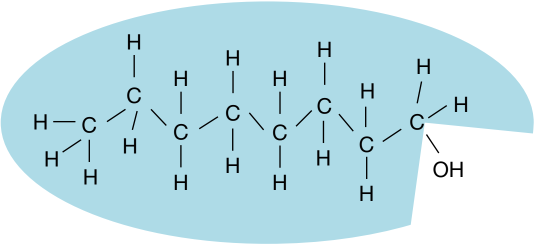


FIGURE 8.24 Larger alcohols become less and less soluble in water due to the increasing influence of the non-polar hydrocarbon portion of the molecule (shaded). Octan-1-ol is not soluble in water.



Alcohols undergo combustion reactions in air to produce carbon dioxide and water. Smaller alcohols such as methanol and ethanol are increasingly being used as fuels. Much research is currently occurring on developing efficient ways to produce ethanol from plant materials in a renewable manner (making so called bioethanol), in order to decrease our dependence on fossil fuels. Many service stations now sell 'E10' — a blend of petrol and up to 10% ethanol.

Using ethanol as an example, the combustion of alcohols may be summarised as follows.

$$\begin{array}{ccccccc} CH_3CH_2OH(l) & + & 3O_2(g) & \rightarrow & 2CO_2(g) & + & 3H_2O(g) \\ \text{ethanol} & + & \text{oxygen} & \rightarrow & \text{carbon dioxide} & + & \text{water} \end{array}$$

Note that ethanol is shown using its semi-structural formula rather than its molecular formula. This is a common practice in organic chemistry.

Ethanol is also used as a solvent (it is widely used and sold as methylated spirits, which is ethanol with a small amount of methanol added) and is present in many products such as perfumes and cosmetics, where a fast rate of evaporation is desirable. It is the type of alcohol in alcoholic drinks. Other commonly used alcohols are methanol which, besides its use as a fuel, is used to make a number of other important chemicals. Also widely used is propan-2-ol (often referred to by its alternative name, isopropanol). This is used as a skin disinfectant and to dissolve oil. It is frequently used to sterilise the skin prior to injections and surgical procedures, and is also used as 'rubbing alcohol'. Its rapid evaporation rate makes it an ideal choice in these situations.

FIGURE 8.25 Many service stations now sell a blend of petrol and ethanol.

Alcohols also react with carboxylic acids to produce esters. These are discussed below.

SAMPLE PROBLEM 7

Consider the following alcohols: ethanol, butan-1-ol and decan-1-ol.

a. **Which would have the highest boiling point?**

b. **Which would be the most soluble in water?**

Teacher-led video: SP7 (tlvd-0545)

THINK	WRITE
a. All alcohols have hydrogen bonding and dispersion forces between their molecules. The larger alcohols will have stronger dispersion forces between their molecules and, therefore, the higher boiling points. Decan-1-ol is the largest molecule from those given (with a 10-carbon chain).	Decan-1-ol will have the highest boiling point.
b. As the hydrocarbon component of the alcohol increases, alcohols become less soluble in water due to the increasing non-polar nature that results. Hence, the molecule with the smallest hydrocarbon component will be the most soluble. This is ethanol (with a 2-carbon chain).	Ethanol is the most soluble in water.

PRACTICE PROBLEM 7

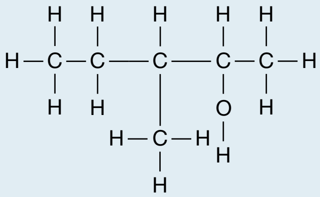
Consider the following alcohols: propan-1-ol, octan-1-ol and pentan-1-ol.

a. Which would have the highest boiling point?

b. Which would be the most soluble in water?

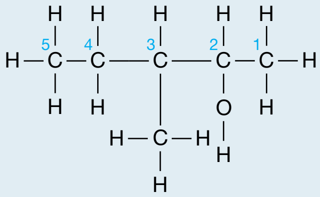
SAMPLE PROBLEM 8

a. Name the following molecule.



b. Write the condensed structural formula for this molecule.

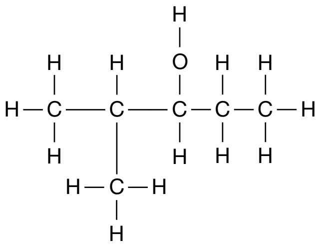
Teacher-led video: SP8 (tlvd-0546)

THINK	WRITE
a. 1. Examine the molecule to find the longest carbon chain and number the chain to give the functional group (—OH) the smallest number. The molecule has five carbons in its longest chain, with the —OH attached to carbon number 2.	 This gives pentan-2-ol.
2. Identify any other hydrocarbon groups on the chain. A methyl group is also attached to carbon number 3, this is denoted by 3-methyl.	
3. Write the name of the compound.	3-methylpentan-2-ol

b. Go carbon-by-carbon along the longest chain, abbreviating the structural formula by grouping the atoms together. Show branches in brackets after the carbon they are attached to.	$CH_3CH(OH)CH(CH_3)CH_2CH_3$

PRACTICE PROBLEM 8

a. Name the following molecule.



b. Write the condensed structural formula for this molecule.

8.3.3 Carboxylic acids

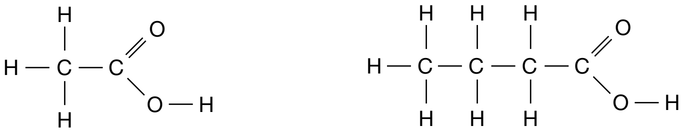
Carboxylic acids are another homologous series. These compounds contain the functional group —COOH at the end of a chain. The —COOH group is called the carboxyl group. When naming a carboxylic acid, add the ending *-oic acid* to the alkyl prefix. Remember to count the carbon atom in the —COOH group. For example, methane becomes methanoic acid. Figure 8.27 shows the structure and names of two further carboxylic acids.

Remember: when identifying and numbering the longest chain in a carboxylic acid, always *start* with the –COOH group. The C in this will *always* be number 1 in the sequence.

Carboxylic acids have the general formula $C_nH_{2n+1}COOH$.

FIGURE 8.26 Painful bee and ant stings contain methanoic acid, HCOOH.

FIGURE 8.27 The structural and semi-structural formulas for two carboxylic acids



CH_3COOH

ethanoic acid (acetic acid)

$CH_3CH_2CH_2COOH$

butanoic acid

Carboxylic acids are a homologous series that contain the carboxyl functional group, —COOH. They have the general formula $C_nH_{2n+1}COOH$.

Properties, uses and reactions of carboxylic acids

As their name suggests, carboxylic acids show typical acid properties. They are weak acids and some, such as ascorbic and citric acids, are present in foods. Familiar reactions include their reaction with bases to form a salt plus water, and their reaction with reactive metals (such as magnesium) to produce a salt plus hydrogen gas. They also react with carbonates, making a salt, carbon dioxide and water. Acid–base behaviour is examined in more detail in unit 2.

Carboxylic acids show trends similar to alcohols with respect to their boiling points and solubility. Their relatively high boiling points are due to the hydrogen bonding that is present; the hydrogen bonding forms due to the —OH component of the —COOH group. Their decreasing solubility is due to the increasing influence of the non-polar hydrocarbon component as their molecules become larger.

TABLE 8.4 Sources and uses of some common organic acids

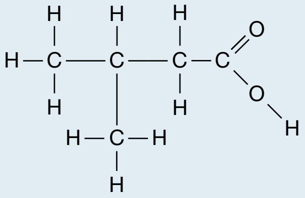
Name	Sources and uses
Methanoic acid (formic acid)	Responsible for sensation caused by nettle, bee and ant stings; used in medicine, food preservation and textile industry
Ethanoic acid (acetic acid)	Main component of vinegar; used as a solvent; salts of the acid used in insecticides and fungicides
Propanoic acid (propionic acid)	Used as an antifungal agent in the baking industry and in ointments (either in salt or acid form)
Butanoic acid (butyric acid)	Odour-causing component of rancid butter, rotten socks and body odour
2-hydroxybenzoic acid (salicylic acid)	Used as a food preservative
2-acetoxybenzoic acid (acetylsalicylic acid, aspirin)	Used for relief of fever, pain and rheumatic conditions

Note: common names are given in brackets

Carboxylic acids also react with alcohols to make the important class of compounds called esters.

SAMPLE PROBLEM 9

a. **Name the following carboxylic acid.**



b. **Write the condensed formula for this compound.**

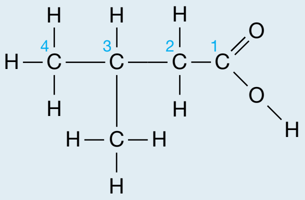
Teacher-led video: SP9 (tlvd-0547)

THINK

a. 1. Carboxylic acids are always numbered from the COOH end. Examine the molecule to find the longest carbon chain, numbering from the C in the COOH.

The longest chain is four carbons (counting the C in the COOH).

WRITE

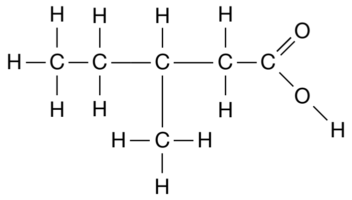


This is a butanoic acid.

2. Identify any other hydrocarbon groups on the chain. A methyl is attached to carbon number 3 (this is denoted by 3-methyl). Write the name of the compound.	3-methylbutanoic acid
b. Go carbon-by-carbon, abbreviating the structural formula by grouping the atoms together. Put side groups in brackets. The standard abbreviation for the carboxyl group is —COOH.	$CH_3CH(CH_3)CH_2COOH$

PRACTICE PROBLEM 9

a. Name the following carboxylic acid.



b. Write the condensed formula for this compound.

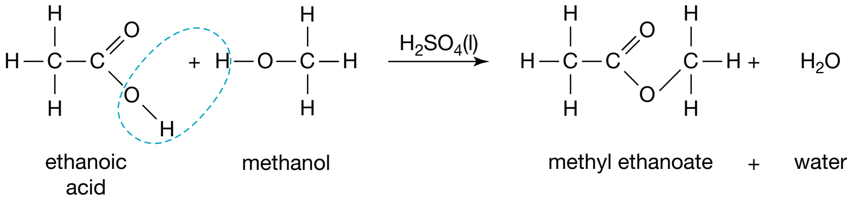
8.3.4 Esters

Esters are formed when an alcohol reacts with a carboxylic acid. Esters are a group of compounds that give the pleasant 'fruity' smell to various fruits, and synthetic esters are often used as an additive to create this pleasant smell. This is called a condensation or esterification reaction. Concentrated sulfuric acid is used to link the alcohol and carboxylic acid together by removing a water molecule.

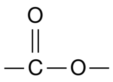
$$\text{carboxylic acid} + \text{alcohol} \xrightarrow{H_2SO_4} \text{ester} + \text{water}$$

$$CH_3COOH(l) + CH_3OH(l) \xrightarrow{H_2SO_4} CH_3\,COOCH_3(l) + H_2O(l)$$

$$\text{ethanoic acid} + \text{methanol} \xrightarrow{H_2SO_4} \text{methyl ethanoate} + \text{water}$$



As can be seen from the preceding example, the functional group for an ester can be represented as



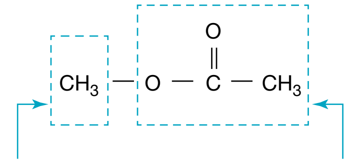
$$-\overset{\overset{\displaystyle O}{\|}}{C}-O-$$

This is abbreviated to COO when writing condensed formulas.

When naming an ester, the alcohol (alkyl part) name is given first, followed by the acid part. For example, the alcohol methanol becomes *methyl*. The -*oic* ending of the acid is replaced with -*oate*. (For example, the carboxylic acid called ethanoic acid becomes *ethanoate*.) So, methanol reacting with ethanoic acid forms the ester methyl ethanoate.

Esters are formed when an alcohol reacts with a carboxylic acid. They form an homologous series that contains the functional group, —COO.

FIGURE 8.28 Naming methyl ethanoate



The alkyl part of the alcohol named methyl

The part considered derived from the organic acid is named ethanoate

Full name of ester: methyl ethanoate

SAMPLE PROBLEM 10

Name the esters produced from the following alcohols and carboxylic acids and write their molecular formula.

a. Propanoic acid and ethanol
b. Methanoic acid and ethanol
c. Ethanoic acid and methanol
d. Which two esters from those listed are isomers?

▨ **Teacher-led video:** SP10 (tlvd-0548)

THINK	WRITE
a. 1. When naming an ester, state the alcohol part first, and then the carboxylic acid part. Change ending to –*oate*. This means ethanol becomes ethyl and propanoic acid becomes propanoate.	The ester produced is ethyl propanoate.
2. Determine the condensed formula of ethyl propanoate by writing the condensed formulas of the reaction between propanoic acid and ethanol.	$CH_3CH_2COOH + CH_3CH_2OH \rightarrow$ $CH_3CH_2COOCH_2CH_3 + H_2O$
3. Write the molecular formula of ethyl propanoate.	$C_5H_{10}O_2$
b. 1. When naming an ester, state the alcohol part first, and then the carboxylic acid part. Change ending to –*oate*. This means ethanol becomes ethyl and methanoic acid becomes methanoate.	The ester produced is ethyl methanoate.
2. Determine the condensed formula of ethyl methanoate by writing the condensed formulas of the reaction between methanoic acid and ethanol.	$HCOOH + CH_3CH_2OH \rightarrow$ $HCOOCH_2CH_3 + H_2O$
3. Write the molecular formula of ethyl methanoate.	$C_3H_6O_2$
c. 1. When naming an ester, state the alcohol part first, then the carboxylic acid part. Change ending to –*oate*. This means methanol becomes methyl and ethanoic acid becomes ethanoate.	The ester produced is methyl ethanoate.

2. Determine the condensed formula of methyl ethanoate by writing the condensed formulas of the reaction between ethanoic acid and methanol.	$CH_3COOH + CH_3OH \rightarrow CH_3COOCH_3 + H_2O$
3. Write the molecular formula of methyl ethanoate.	$C_3H_6O_2$
d. Isomers have the same molecular formula.	Of those listed, (b) ethyl methanoate and (c) methyl ethanoate are isomers. Molecular formula is $C_3H_6O_2$.

PRACTICE PROBLEM 10

Name the esters produced from the following alcohols and carboxylic acids and write their molecular formula.

a. Propanoic acid and methanol
b. Methanoic acid and propan-1-ol
c. Ethanoic acid and ethanol
d. Which esters from those listed are isomers?

Properties, uses and reactions of esters

Esters have relatively low boiling points because the intermolecular attraction is dipole–dipole (and dispersion forces) instead of the hydrogen bonding that occurs in carboxylic acids. Table 8.5 demonstrates this for a similarly sized ester and carboxylic acid.

TABLE 8.5 Ester and carboxylic acid boiling points

Name	Molecular formula	Condensed formula	Type	Boiling point (°C)
Ethyl ethanoate	$C_4H_8O_2$	$CH_3COOCH_2CH_3$	Ester	77.1
Butanoic acid	$C_4H_8O_2$	$CH_3CH_2CH_2COOH$	Carboxylic acid	164

Although esters cannot form hydrogen bonds with themselves, they can with water. This leads to the same solubility trend that is observed for alcohols and carboxylic acids. Smaller esters are soluble but, as the number of carbons increases, they become less and less soluble (see table 8.6).

TABLE 8.6 The trend in ester solubility with increasing molecular size.

Name	Formula	Solubility in water (g/100 g)
Ethyl methanoate	$HCOOCH_2CH_3$	10.5
Ethyl ethanoate	$CH_3COOCH_2CH_3$	8.8
Ethyl propanoate	$CH_3CH_2COOCH_2CH_3$	1.7

The formation of an ester can be reversed by heating the ester with a dilute acid such as sulfuric acid or a dilute base such as sodium hydroxide. This is called **hydrolysis**. As expected, this regenerates the original alcohol and carboxylic acid.

Esters are important biologically. The important class of biomolecules called lipids (which includes oils and fats) are esters formed between an alcohol called glycerol and long chain carboxylic acids (often referred to as fatty acids).

TABLE 8.7 Some esters and their uses

Name	Common name	Uses
Methyl salicylate		Liniment (treatment for sprains and bruises)
Ethyl ethanoate	Ethyl acetate	Nail polish remover, glue
Ethyl butanoate	Ethyl butyrate	Banana flavouring
Propyl ethanoate	Propyl acetate	Pear flavouring
Ethyl hexanoate		Pineapple flavouring
Acetylsalicylic acid	Aspirin	Pain relief, blood thinner

Esters are also becoming popular as a fuel, where they are being used to make *biodiesel*. In this production, naturally occurring plant and animal esters (oils and fats) are reacted with methanol in a reaction where one type of ester is transformed into another to make a fuel that can be used in place of diesel. This is examined in unit 3.

on Resources

Digital documents	Experiment 8.1 Properties of carboxylic acids (doc-30849)
	Experiment 8.2 Formation of esters (doc-30850)
Teacher-led video	Experiment 8.1 Properties of carboxylic acids (tlvd-0624)

8.3 EXERCISE

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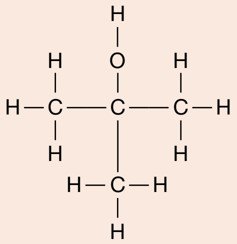
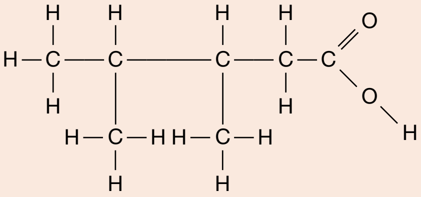
1. Write the formula for an alcohol that contains seven carbon atoms in its molecules.
2. Write the formula for a carboxylic acid that contains six carbon atoms in its molecules.
3. Give a definition of a functional group. Give two examples.
4. (a) Name the following molecule.



(b) Write the condensed structural formula for this molecule.
(c) Explain why a number is not required to indicate the CH_3 group in the name from part a, but is required to indicate the —OH group.

5. (a) Name the following carboxylic acid.



   (b) Write the condensed formula for this compound.

6. (a) What is a homologous series?
   (b) Explain why alcohols are an example of a homologous series. Support your answer with some appropriate formulas.
   (c) Explain why carboxylic acids are an example of a homologous series. Support your answer with some appropriate formulas.
   (d) Explain why alkanes are an example of a homologous series. Support your answer with some appropriate formulas.

7. Write molecular, structural and condensed formulas for the following.
   (a) Butan-1-ol
   (b) Butan-2-ol

8. Write molecular, structural and condensed formulas for the following.
   (a) Ethanoic acid
   (b) Propanoic acid
   (c) Butanoic acid

9. (a) Name the esters produced from the following alcohols and carboxylic acids.
      i. Butanoic acid and methanol
      ii. Propanoic acid and ethanol
      iii. Methanoic acid and methanol
   (b) Which esters from (i) to (iii) are isomers? What is their molecular formula?

10. Draw the structural formula for nonanoic acid and write its condensed formula.

11. Draw the structural formula for ethyl pentanoate and write its condensed formula.

12. Pentanoic acid and ethyl propanoate both have the formula $C_5H_{10}O_{2.}$ Their boiling points are 185 °C and 99 °C respectively.
    With reference to their structures and bonding, explain why this difference occurs.

13. Name the alcohol and carboxylic acid used to make the following.
    (a) Propyl ethanoate
    (b) Ethyl propanoate
    (c) Pentyl decanoate

14. Methyl ethanoate is heated with dilute sulfuric acid and a reaction takes place. What are the products of this reaction?

15. Glycerol, $C_3H_8O_3$, is an important naturally occurring alcohol that is present in our bodies. It contains three hydroxy groups per molecule. Each hydroxy group is attached to a separate carbon atom. Also important in our bodies are long chain carboxylic acids called fatty acids. These react with glycerol to make esters that are commonly called oils or fats. One example of such a fatty acid is stearic acid, $C_{18}H_{36}O_2$.
    (a) Write the condensed formula for glycerol.
    (b) Write the condensed formula for stearic acid.

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8.4 The origin and use of crude oil

KEY CONCEPT

- The origin of crude oil and its use as a source of hydrocarbon raw materials

8.4.1 Formation

Crude oil is also known as petroleum. The word ‘petroleum’ is derived from two Latin words, *petro* meaning ‘rock’ and *oleum* meaning ‘oil’. So it is literally ‘rockoil’. It was given this name because it was originally found seeping through rocks to the Earth’s surface. Petroleum takes many thousands of years to form. Millions of years ago, warm seas covered much of the area that is now land. These areas contained plant and animal life that died, fell to the sea floor and became covered with sediment. Increased pressure and temperature over a long period of time caused these substances to change into petroleum. Later, movements of the Earth’s crust pushed these deposits closer to the surface and trapped pockets of natural gas, crude oil and salt water.

Hydrocarbons are found in crude oil, and processing the oil yields many useful products. The hydrocarbons described so far in this topic are largely synthesised or produced from fossil fuels. Fossil fuels also form the major part of our fuel resources. As the name suggests, they are derived from the fossil remains of living organisms, which have been altered by heat and pressure.

8.4.2 Refining

The process of obtaining useful products from crude oil is called refining. It is carried out on a large scale in complexes called oil refineries. At the time of writing, Australia has operating oil refineries in Victoria, Queensland, South Australia and Western Australia. Oil refining involves several processes. However, two of the most important are fractional distillation (also known as fractionating) and cracking.

Fractional distillation

Many mixtures of liquids, including crude oil, may be separated into its components, called fractions, using the process known as **fractional distillation**. This process is used to separate one liquid from a mixture of liquids with different boiling temperatures. In industry, this process is used on a large scale to separate crude oil into a number of fractions based on boiling temperatures. This is the first stage of refining. It represents a ‘rough sorting out’ of the thousands of different compounds that are in the crude oil.

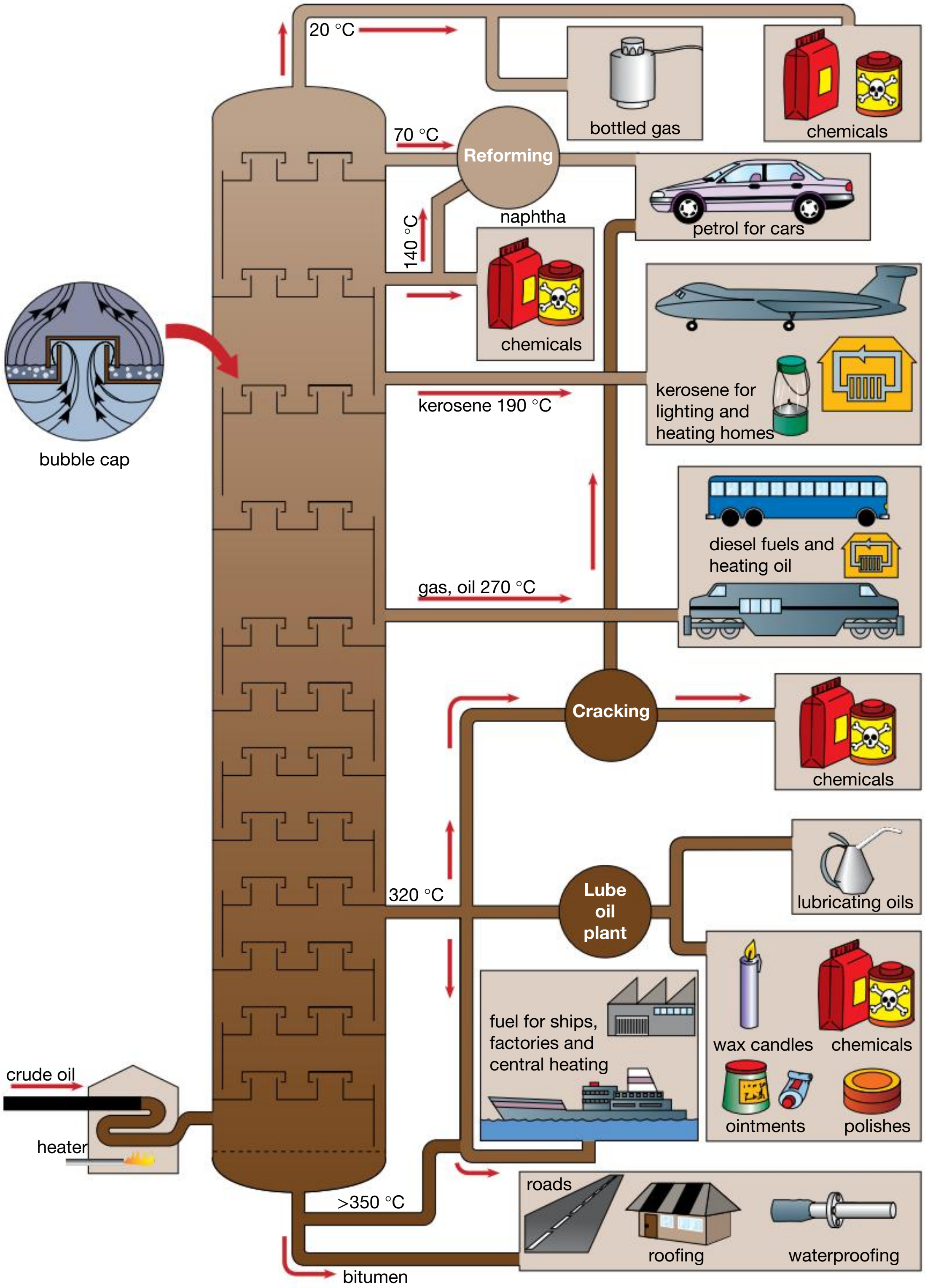
In this process, hot crude oil (which is essentially a very complicated mixture of alkanes) enters the base of the fractionating tower and almost all of it is immediately vaporised. The small portion that remains liquid (due to the high boiling temperature caused by its large molecules) is removed from the base to make products such as bitumen for roads. As the vapours rise up the tower, they cool and re-liquefy at different heights depending on their boiling (or condensation) temperatures. The lighter fractions have the smallest

TABLE 8.8 Products from the fractional distillation of crude oil

Product	Number of carbon atoms	Boiling point (°C)	Use
Gas	1–4	<40	Bottled gas, plastics, chemicals
Petrol	4–12	40–75	Fuel for vehicles, chemicals
Kerosene	9–16	150–240	Jet fuel, chemicals
Diesel oil	15–25	220–250	Fuel
Lubrication oil	20–70	250–350	Lubricants, waxes
Bitumen residue	>70	>350	Road surfaces

molecules and, therefore, the lowest condensation temperatures; they are collected near the top of the tower. Moving down the tower, the fractions collected become ‘heavier’; that is, they contain larger and larger molecules. Special arrangements inside the tower called bubble caps let vapours rise while allowing any condensed liquids to be collected on special trays.

FIGURE 8.29 A fractionating tower showing oil products and their uses



SAMPLE PROBLEM 11

During the fractional distillation of crude oil, explain why the molecules for wax are collected from a lower level than the molecules for diesel.

Teacher-led video: SP11 (tlvd-0549)

THINK	WRITE
Recall the following: • Only dispersion forces are present between the hydrocarbons in crude oil. • Increasing molecular size causes the strength of dispersion forces to increase and, hence, boiling temperatures as well. • Temperature decreases as vapours rise up the distillation column.	The molecules making up wax are larger than those making up diesel fuel. Because dispersion forces are the only intermolecular attraction between all these molecules, a stronger force of attraction will exist between wax molecules compared to diesel molecules. Wax will, therefore, have the higher boiling point (also higher condensation point) and so will be collected lower in the column.

PRACTICE PROBLEM 11

Explain why the molecules for bottled gas are collected from a higher level during the fractional distillation of crude oil than the molecules for petrol.

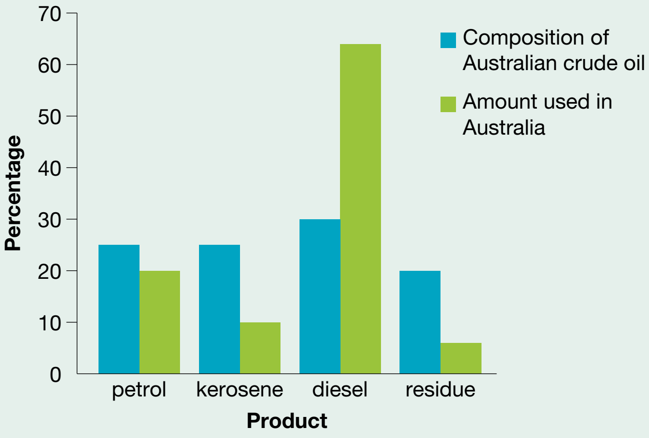
CRACKING TO PRODUCE DIESEL FUEL

We need far more diesel fuel than the amount obtained from the fractional distillation of Australian crude oil. This process also produces far more kerosene than we need.

In order to obtain more diesel and other smaller molecules, a process has been developed whereby some of the excess larger hydrocarbons are broken into smaller ones — the ones that, generally speaking, are more useful. This process is called 'cracking'. This can be done in two main ways. The first uses high temperatures and is termed **thermal cracking**. The second uses a catalyst at much lower temperatures to produce the same effect. This is called **catalytic cracking**. Both methods have their own advantages and disadvantages, and both are used at a number of locations around Australia.

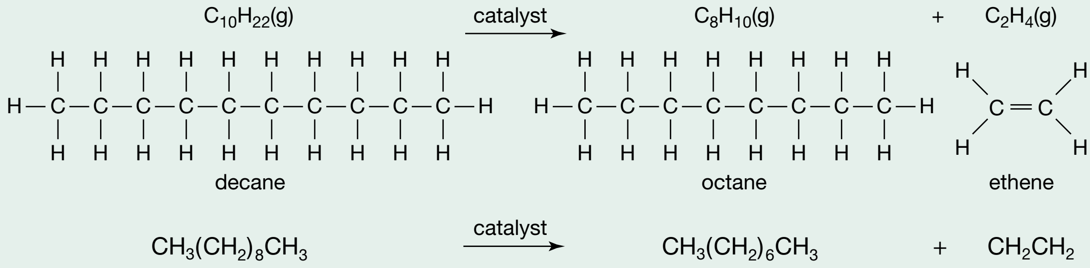
An important by-product of cracking is the industrially important class of hydrocarbons called alkenes — the raw materials of the plastics industry.

FIGURE 8.30 A number of different fractions can be obtained from crude oil. Unfortunately, demand does not equal the output of the various fractions.



An example of cracking is the breakdown of decane (a large hydrocarbon) into octane (used in motor fuel) and ethene (used for the manufacture of plastic).

FIGURE 8.31 One possible way in which decane could crack. It could also crack at other points to produce different molecules.



$C_{10}H_{22}(g) \xrightarrow{\text{catalyst}} C_8H_{18}(g) + C_2H_4(g)$

$CH_3(CH_2)_8CH_3 \xrightarrow{\text{catalyst}} CH_3(CH_2)_6CH_3 + CH_2CH_2$

8.4 EXERCISE

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1. Describe how crude oil (petroleum) is formed.
2. Explain why the molecules for petrol are collected from a higher level during the fractional distillation of crude oil than the molecules for kerosene.
3. What property of the lighter fractions of crude oil makes them burn more easily than the heavier ones?
4. Viscosity refers to how well a liquid flows; honey is described as a viscous liquid. Why do you think the fractions that are recovered lower in the fractional distillation tower have increasing viscosity?
5. Write a chemical equation to show one way in which dodecane, $C_{12}H_{26}$, could crack when subjected to catalytic cracking. Compare your answers with others in the class. Who is correct?

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8.5 Determining the empirical and molecular formulas of organic compounds

KEY CONCEPT

- Determination of empirical and molecular formulas of organic compounds

8.5.1 Empirical formulas

In topic 5, we saw how to calculate the empirical formula of an ionic compound. The empirical formula of a molecular compound can be calculated the same way.

Recall the steps involved in finding an empirical formula:

1. Write down the *symbols* of the elements present.
2. Assume that the *mass* of the sample is 100 g and all percentages become grams.
3. Convert masses to *moles*.
4. Find the simplest *ratio* of the atoms by dividing all numbers of moles by the smallest number of moles.
5. If necessary, multiply by a *factor* to convert all numbers to whole numbers.

SAMPLE PROBLEM 12

A hydrocarbon is burnt in excess oxygen to produce 2.64 g of carbon dioxide and 1.26 g of water. Calculate its empirical formula.

Teacher-led video: SP12 (tlvd-0550)

THINK

WRITE

1. Hydrocarbons burn in excess oxygen to produce carbon dioxide and water. We can use the given mass of the products to determine the mass of the carbon and the hydrogen in the reactants. All the carbon ends up in CO_2; all the hydrogen ends up in H_2O.

   Follow the steps listed in section 8.5.1 to find the empirical formula.

2. Step 1: The empirical formula will contain C and H.

   Step 2: To determine the mass of carbon in 2.64 g CO_2, find the mass ratio of C in CO_2 and multiply by total mass of CO_2.

   To determine the mass of hydrogen in 1.26 g H_2O, find the mass ratio of H in H_2O and multiply by total mass of H_2O.

   Step 3: Convert masses to moles of each

   Step 4: Find simplest ratio by dividing all numbers of moles by the smallest number of moles.

   Step 5: Convert to whole numbers

Step 1	**Symbols**	**C**	**H**
Step 2	**Masses**	$\left(\frac{12.0}{44.0}\right) \times 2.64$ $= 0.720\text{ g}$	$\left(\frac{2.0}{18.0}\right) \times 1.26$ $= 0.140\text{ g}$
Step 3	**Moles**	$\frac{m}{M}$ $= \frac{0.720}{12.0}$ $= 0.0600\text{ mol}$	$\frac{m}{M}$ $= \frac{0.140}{1.0}$ $= 0.140\text{ mol}$
Step 4	**Simplest ratio**	$\frac{0.0600}{0.0600} = 1$	$\frac{0.140}{0.0600} = 2.33$
Step 5	**Whole numbers**	$1 \times 3 = 3$	$2.33 \times 3 = 7$

3. Write the empirical formula

C_3H_7

PRACTICE PROBLEM 12

A compound containing carbon, hydrogen and oxygen is found to contain 0.120 g of carbon, 0.025 g of hydrogen and 0.080 g of oxygen. Calculate its empirical formula.

8.5.2 Molecular formulas

The molecular formula of a compound represents the actual composition of a compound that is made up of molecules.

A molecular formula may be the same as its experimentally determined empirical formula or be a whole-number multiple of it. The molecular formula of a compound may be determined from its empirical formula only if its molar mass is also known.

Molecular formula of a compound:

$$\textbf{molecular formula} = n \times (\textbf{empirical formula})$$

where *n* represents a whole number.

Although compounds may have the same empirical formula, compounds with different molecular formulas may have very different structures and, as a result, very different properties. Consider the empirical formula CH_2O. Methanal, also known as formaldehyde, has the molecular formula CH_2O, with a boiling point of –21 °C. It is commonly used as a preservative of biological samples and in the manufacture of pressed wood products. A multiple of two of the same empirical formula gives ethanoic acid, $C_2H_4O_2$, also known as acetic acid, but commonly recognised as vinegar. It has a boiling point of 117 °C. A multiple of six of the same empirical formula gives a variety of compounds, including many of simple sugars, such as glucose, which can have a ring or straight chain structure. A comparison of empirical formulas and molecular formulas of some organic molecules can be found in table 8.9.

TABLE 8.9 Comparison of empirical and molecular formulas

Name of molecule	Empirical formula (simplest ratio of atoms in molecule)	Molecular formula (actual number of atoms in molecule)
Ethyne (acetylene)	CH	C_2H_2
Benzene	CH	C_6H_6
Formaldehyde	CH_2O	CH_2O
Ethanoic acid (acetic acid)	CH_2O	$C_2H_4O_2$
Glucose	CH_2O	$C_6H_{12}O_6$

SAMPLE PROBLEM 13

Benzene has the empirical formula CH and its molar mass is 78.0 g mol^{-1}. Find the molecular formula.

Teacher-led video: SP13 (tlvd-0551)

THINK

Determine the empirical formula mass. If the molar mass is 78.0 g mol^{-1}, its relative molecular mass (the sum of the relative atomic masses according to the molecular formula) is also 78.0. If the empirical formula is written as C_xH_y, then the molecular formula is $(C_xH_y)_n$, where x, y and n are whole numbers. The value of n may be determined by comparing the relative molecular mass and the empirical formula mass.

WRITE

The empirical formula is CH.

$$\begin{aligned} M(\text{empirical}) &= A_r\,(\text{C}) + A_r\,(\text{H}) \\ &= 12.0 + 1.0 \\ &= 13.0 \\ \frac{M(\text{actual})}{M(\text{empirical})} &= \frac{78.0}{13.0} \\ \text{ratio} &= 6 \\ \therefore \text{molecular formula} &= 6 \times \text{empirical formula} \\ &= 6 \times \text{CH} \\ &= \text{C}_6\text{H}_6 \end{aligned}$$

PRACTICE PROBLEM 13

Calculate the molecular formula of the hydrocarbon from Sample problem 12, given that its molar mass is 86.0 g mol^{-1}.

SAMPLE PROBLEM 14

A compound with the properties of an acid is analysed and found to contain 40.0% carbon, 6.7% hydrogen and 53.3% oxygen. It is found to have a molar mass of 60.0 g mol^{-1}.

a. **Calculate its empirical formula.**

b. **Use its empirical formula to determine its molecular formula.**

c. **Write the structural and condensed formulas for this compound.**

Teacher-led video: SP14 (tlvd-0552)

THINK

a. 1. To determine the empirical formula, follow the steps outlined in section 8.5.1.

Step 1: The compound contains C, H and O.

Step 2: Convert percentages to g by assuming a 100 g sample.

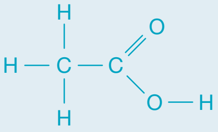
Step 3: Convert masses to moles of each by using the formula $n = \frac{m}{M}$.

TIP: The formula can be found in Table 3 of the VCE Chemistry Data book.

Step 4: Find the simplest ratio of the atoms.

WRITE

Step 1	Symbols	C	H	O
Step 2	Masses	40 g	6.7 g	53.3 g
Step 3	Moles	$\frac{40}{12.0}$ $= 3.33$	$\frac{53.3}{16.0}$ $= 3.33$	$\frac{53.3}{16.0}$ $= 3.33$
Step 4	Simplest ratio	$\frac{3.3}{3.3}$ $= 1$	$\frac{6.7}{3.3}$ $= 2$	$\frac{3.3}{3.3}$ $= 1$

2. Write the empirical formula	Empirical formula is CH_2O
b. 1. Calculate the molar mass that the empirical formula, CH_2O, predicts.	$M(\text{empirical formula}) = C + 2 \times H + O$ $= 12.0 + (1.0 \times 2) + 16.0$ $= 30\ \text{g mol}^{-1}$
2. Compare the mass of the empirical formula to given molar mass and find the ratio between the two to obtain molecular formula.	$M(\text{actual}) = 60\ \text{g mol}^{-1}$ $\frac{M(\text{actual})}{M(\text{empirical formula})} = \frac{60}{30}$ $\text{Ratio} = 2$ $\therefore M(\text{actual}) = 2 \times M(\text{empirical formula})$
3. Write the molecular formula.	Molecular formula $= 2 \times CH_2O = C_2H_4O_2$
c. Given the compound has the properties of an acid, draw out a structure containing carboxylic acid functional group (—COOH) and add remaining atoms. Write condensed formula.	 CH_3COOH

PRACTICE PROBLEM 14

A compound with the properties of an acid is analysed and found to contain 26.1% carbon, 4.3% hydrogen and 69.6% oxygen. It is found to have a molar mass of 46 g mol^{-1}.

a. Calculate its empirical formula.

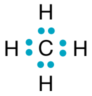
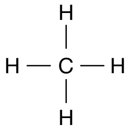
b. Use its empirical formula to determine its molecular formula.

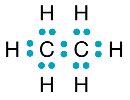
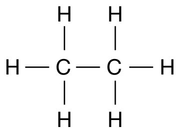
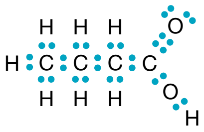
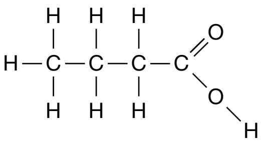
c. Write the structural and condensed formulas for this compound.

8.5.3 Summary of formula types

It may now be seen that a molecular substance can have many different types of formula. Each of these different types shows different amounts of information in different ways. Which type is chosen will depend on the information that is being conveyed. Tables 8.3 (section 8.2.8) and 8.10 show the different types of formulas that may be written for a molecular compound.

TABLE 8.10 Comparison of formula types for three different organic compounds.

Name	Empirical formula	Molecular formula	Electron dot (Lewis) structure	Structural formula	Condensed formula
Methane	CH_4	CH_4			CH_4

Name	Empirical formula	Molecular formula	Electron dot (Lewis) structure	Structural formula	Condensed formula
Ethane	CH_3	C_2H_6			CH_3CH_3
Butanoic acid	C_2H_4O	$C_4H_8O_2$			$CH_3CH_2CH_2COOH$

8.5 EXERCISE

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1. Which of the following formulas are empirical formulas?
   (a) CH_4
   (b) CH_4O
   (c) C_2H_3
   (d) C_2H_6
   (e) C_2H_6O
   (f) C_2H_2
   (g) $C_3H_4O_2$
   (h) C_3H_8O
   (i) $C_4H_6O_2$
2. Explain why the molecular formula of an alcohol that contains only one hydroxy group in its molecules is always also its empirical formula.
3. Ethane-1,2-diol (ethylene glycol) is used in antifreeze. It has the empirical formula CH_3O and its molar mass is 62.0 g mol^{-1}. Calculate its molecular formula.
4. Calculate the molecular formulas for the compounds labelled a. to g. from the information in the following table.

Compound	Empirical formula	Molar mass (g mol^{-1})
a.	C_5H_{12}	72.0
b.	C_3H_5	82.0
c.	C_3H_4	80.0
d.	CH_2	84.0
e.	CH_2	140.0
f.	CH_2O	180.0
g.	CH_2Cl	99.0

5. Nicotine, the main active chemical in tobacco, has the empirical formula C_5H_7N and a relative molecular mass of 162.0. Determine the molecular formula of nicotine.
6. Caffeine is a stimulant that is found naturally in coffee, tea and chocolate. Analysis of caffeine shows that it contains 49.5% carbon, 28.9% nitrogen, 16.5% oxygen and 5.1% hydrogen by mass. Determine the molecular formula of caffeine given that its molar mass is 194.2 g mol^{-1}.

7. The compound methyl butanoate smells like apples. Its percentage composition is 58.8% C, 9.8% H and 31.3% O. If its molar mass is 102.0 g mol^{-1}, what is its molecular formula?
8. The taste of sour milk is due to lactic acid. The percentage composition of lactic acid by mass is 40.00% carbon, 6.71% hydrogen and 53.29% oxygen, and the molar mass is 90.0 g mol^{-1}. Find the empirical formula and molecular formula of lactic acid.
9. A compound contains 12.8% carbon and 2.13% hydrogen, the rest being bromine. The relative molecular mass of the compound is 188.0. Calculate the empirical formula and the molecular formula of the compound.
10. An ester that is used in making nail polish remover is manufactured using ethanol. It has the following composition: 54.5% carbon, 9.15% hydrogen and 36.3% oxygen. Its molar mass is 88.0 g mol^{-1}. Calculate the empirical formula and molecular formula of the ester and draw its structure.
11. A 5.00 g sample of a hydrocarbon contains 0.908 g of hydrogen. Its molar mass is 44.0 g mol^{-1}.
    (a) Find its molecular formula.
    (b) Draw its structure.
    (c) State the homologous series to which it belongs.
12. The combustion of a certain hydrocarbon in excess oxygen yields 0.341 g of carbon dioxide and 0.134 g of water. Its molar mass is discovered to be 70 g mol^{-1}.
    Calculate its molecular formula.
13. Analysis of a certain compound reveals that it contains only carbon, hydrogen and oxygen.
    (a) State two homologous series that this compound might belong to.
    (b) Describe a simple test which *might* differentiate between the answers from part a.
    (c) The complete combustion of 6.998 g of this compound yielded 9.636 g of carbon dioxide and 7.866 g of water.
    Calculate its empirical formula.
    (d) In a further experiment, its molar mass was found to be 32 g mol^{-1}.
    Calculate its molecular formula.
14. A compound has the empirical formula C_2H_4O. When magnesium is added to it, hydrogen gas is evolved in a 1:1 ratio of compound to hydrogen.
    Name this compound and write its structural formula.

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8.6 Review

8.6.1 Summary

Grouping hydrocarbons into families

- Organic chemistry is the study of the compounds of carbon. The oxides of carbon, carbonates, carbides and cyanides, and compounds of the remaining elements, are considered to be inorganic substances.
- Carbon in organic compounds always forms four covalent bonds.
- Hydrocarbons are organic compounds made up of only carbon and hydrogen atoms. Dispersion forces are involved in intermolecular bonding in hydrocarbons. This means that boiling temperatures are low. Hydrocarbon molecules are non-polar so they are insoluble in water.
- Saturated hydrocarbons contain only single bonds; unsaturated hydrocarbons contain double or triple bonds.
- Alkanes are hydrocarbons containing only single bonds between carbon atoms. Alkanes can undergo oxidation (combustion) and substitution reactions. They burn in a plentiful supply of oxygen to produce carbon dioxide and water.
- Haloalkanes contain one or more halogen atoms.

- Alkenes are hydrocarbons containing one double bond between two carbon atoms. Alkenes can undergo combustion, addition and polymerisation reactions. These reactions involve, respectively, reacting with oxygen to form carbon dioxide and water; reacting with hydrogen (or the halogens) so that the double bond between the two carbon atoms is broken; and forming long chains called polymers as a result of alkene molecules joining together.
- Alkynes are hydrocarbons containing one triple bond between two carbon atoms. Alkynes can undergo combustion and addition reactions.
- In a homologous series, successive members differ by a $—CH_2—$ group.
- Isomers of a compound have the same molecular formula but different structural formulas.
- Every organic compound can be named according to its structure using a set of rules devised by the IUPAC. However, many compounds have older names that are still in common use.
- Condensed formulas (also called semi-structural formulas) are often used in organic chemistry to denote the composition and structure of organic molecules. Lewis structures (electron dot diagrams) can also be used but are not as common.

Functional groups

- Isomers of a compound have the same molecular formula but different structural formulas.
- A functional group is the bond, atom or group of atoms that gives a group of molecules its specific properties.
- Alcohols contain one or more hydroxy ($—OH$) groups. Due to the presence of hydrogen bonding, alcohols have a higher boiling temperature than the corresponding alkane. Smaller alcohols are soluble in water.
- Carboxylic acids contain a carboxyl ($—COOH$) group at the end of a chain. Due to the presence of hydrogen bonding, carboxylic acids have a higher boiling temperature than the corresponding alkane. Smaller carboxylic acids are soluble in water.
- Esters form when alcohols react with carboxylic acids. Esters are named by changing the name of the alcohol to the alkyl form and replacing the ending of the name of the carboxylic acid with *-oate*. For example, the ester formed from propanol and ethanoic acid is propyl ethanoate.

The origin and use of crude oil

- Crude oil is formed from the remains of small marine organisms that existed millions of years ago. These died and sank to the bottom where they became covered with sediment. Increased temperature and pressure over a long time turned these remains into crude oil.
- The three main groups of fossil fuels are natural gas, petroleum and coal.
- Hydrocarbons are obtained from petroleum (crude oil).
- Petroleum is separated using fractional distillation. Fractional distillation separates components according to boiling points. In a fractionating tower, components with lower boiling points are drawn up closer to the top.

Determining the empirical and molecular formulas of organic compounds

- The empirical formula of a compound is the simplest whole number ratio of atoms that are present.
- The molecular formula of a compound may be determined from its empirical formula and its molar mass. The molecular formula will always be some whole number (including one) times the empirical formula.
- An empirical formula (and molecular formula if appropriate) uses either experimentally determined elemental masses or percentages as the basis for its calculation.

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8.6.2 Key terms

addition reaction a reaction in which one molecule bonds covalently with another molecule without losing any other atoms
alcohol compound containing at least one hydroxy (—OH) group
alkane a simple, saturated binary compound of carbon and hydrogen atoms with single bonds between the carbon atoms; the general formula for all members of this homologous series is C_nH_{2n+2}
alkene a hydrocarbon that contains one carbon–carbon double bond
alkyne a hydrocarbon that contains one carbon–carbon triple bond
carboxylic acid compound containing at least one carboxy (—COOH) group
catalytic cracking the use of a catalyst to break bonds in long molecules to produce shorter molecules
condensed formula formula derived from the structural formula that shows the atoms present in a structure carbon by carbon and is written on a single line; also called a semi-structural formula
crude oil naturally occurring mixture of compounds that are mainly hydrocarbons; also known as petroleum
cyclic compound compound containing a ring structure, such as benzene
double bond strong bond between two atoms formed by two pairs of electrons that are shared by the two nuclei
empirical formula formula that shows the simplest numerical ratio in which atoms are combined
ester organic compound formed from a condensation reaction between an alcohol and a carboxylic acid
fractional distillation method of separating the components of a liquid mixture that depends on the ease of vaporisation of the components
functional group group of atoms attached to or part of a hydrocarbon chain that influence the physical and chemical properties of the molecule
haloalkane compound containing a halogen and an alkane
homologous series series of organic compounds with the same structure but in which the formula of one molecule differs from the next by a —CH_2 group
hydrocarbon compound containing only carbon and hydrogen
hydrolysis reaction involving the addition of water
Lewis structure structure that shows how outer shell electrons are arranged in a molecule; also known as electron dot structure
molecular formula actual number of atoms of each element in a molecule of a substance
organic chemistry the study of carbon-containing compounds and their properties
polymer large molecule formed by the joining together of many smaller molecules
saturated hydrocarbon compound composed of carbon and hydrogen and containing only single covalent bonds
single bond bond formed by one pair of electrons shared between two atoms
skeletal formula structural representation where a molecule is represented by vertices; carbons and hydrogens are assumed to be at each vertex and at the ends unless otherwise indicated
structural formula diagrammatic representation of a molecule showing every bond
structural isomers compounds that have the same molecular formula but different arrangements for the atoms in that formula
substitution reaction reaction in which an atom (or group of atoms) is removed and replaced by a different atom (or group of atoms)
thermal cracking the use of heat to break bonds in long molecules to produce shorter molecules
triple bond strong bond between two atoms formed by three pairs of electrons that are shared by the two nuclei
unsaturated hydrocarbon compound containing carbon and hydrogen but possessing one or more double or triple bonds

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Digital document Key terms glossary — Topic 8 (doc-30952)

8.6.3 Practical work and experiments

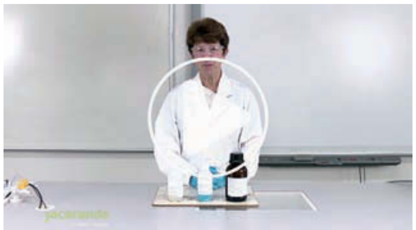
Experiment 8.1

Properties of carboxylic acids

Aim: To examine a range of properties of carboxylic acids.

Digital document: doc-30849

Teacher-led video: tlvd-0624



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Digital documents Practical investigation logbook (doc-30953)
Experiment 8.2 Formation of esters (doc-30850)

8.6 Exercises

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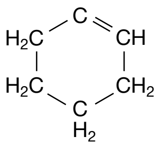
8.6 Exercise 1: Multiple choice questions

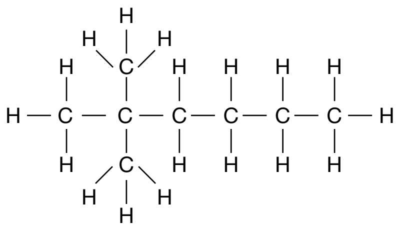
1. Which of the following is *not* the formula of an alkane?
   A. C_4H_{10}
   B. C_7H_{16}
   C. $C_{11}H_{22}$
   D. $C_{14}H_{30}$
2. Which of the following is the formula for an alcohol?
   A. $CH_3CH_2CH_2COOH$
   B. $CH_3CH_2CH_2OH$
   C. $CH_3CH_2CH_2COH$
   D. $CH_3CH_2CH_2CH_3$
3. Which of the following is the formula for a carboxylic acid?
   A. $CH_3CH_2CH_2COH$
   B. $CH_3CH_2CH_2OH$
   C. $CH_3CH_2CH_2COOH$
   D. $CH_3CH_2CH_2CH_3$
4. In which of the following groups are all the compounds members of the same homologous series?
   A. CH_4; C_2H_4; C_3H_4
   B. CH_3OH; CH_3CH_2OH; $CH_3CH_2CH_2OH$
   C. CH_4; CH_3OH; $HCOOH$
   D. CH_3OCH_3; CH_3CH_2OH; C_2H_5OH

5. Which one or more of the following statements is *not* correct?
   - A. A functional group remains unchanged in most chemical reactions.
   - B. A functional group is common to all members of the particular homologous series.
   - C. A functional group may be regarded as being responsible for a characteristic set of chemical reactions for a given homologous series.
   - D. A functional group varies from member to member by —CH_2.
6. What is an alkyl group?
   - A. A strongly basic group in an organic molecule
   - B. A group of organic compounds related to the alkanes
   - C. A group equivalent to an alkane molecule less one hydrogen atom
   - D. A group of organic compounds with the general formula C_nH_{2n+2}
7. What is the process of changing large hydrocarbon molecules into smaller hydrocarbon molecules called?
   - A. Distillation
   - B. Polymerisation
   - C. Cracking
   - D. Esterification
8. Which of the following statements best describes a saturated compound?
   - A. A saturated compound, when used as a solvent, does not dissolve any more solute at a particular temperature.
   - B. A saturated compound contains carbon and hydrogen atoms only.
   - C. A saturated compound has only single carbon–carbon bonds.
   - D. A saturated compound does not react with chlorine.
9. Which one or more of the following compounds is saturated?
   - A. $H_2C = CH_2$
   - B. $HC \equiv CH$
   - C.

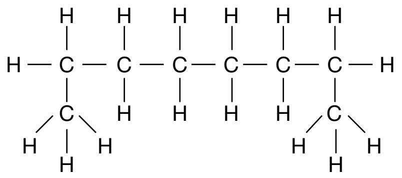


   - D. $CH_3CH_2CH_3$
10. The structure that is *not* an isomer of C_8H_{18} is:

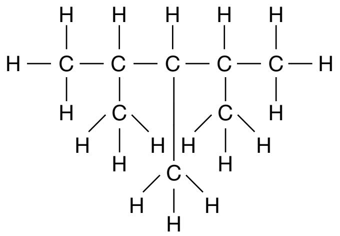
A.



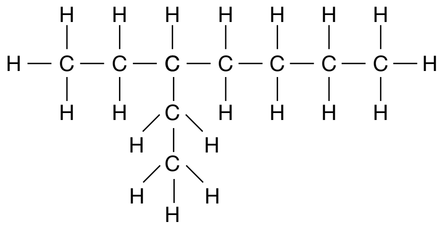
B.



C.



D.



11. Which of the following condensed formulas is an *incorrect* representation of the molecules shown in question 10?

A. $CH_3C(CH_3)_2CH_2CH_2CH_2CH_3$
B. $CH_3(CH_2)_6CH_3$
C. $CH_3CH(CH_3)CH(CH_3)CH(CH_3)CH_3$
D. $CH_3CH_2CHCH_2(CH_3)CH_2CH_2CH_2CH_3$

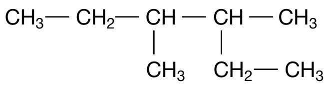
12. What is 2-methylbutan-2-ol an isomer of?

A. Butan-2-ol
B. Pentanoic acid
C. Pentan-1-ol
D. 2, 3-dimethyl-butan-2-ol

13. Which organic compound in the following group has the highest boiling point?

A. Acetylene (ethyne)
B. Ethane
C. 1,1,1-trichloroethane
D. Ethan-2-ol

14. An organic compound has the following structure.

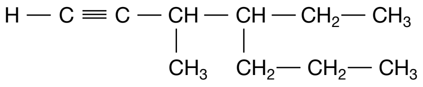


$CH_3—CH_2—CH(CH_3)—CH(CH_2—CH_3)—CH_3$

What is the name of the compound?

A. 2-ethyl-3-methylpentane
B. Octane
C. 3,4-dimethylhexane
D. 3-methyl-4-ethylpentane

15. An organic compound has the following structure.



$H—C\equiv C—CH(CH_3)—CH(CH_2—CH_2—CH_3)—CH_2—CH_3$

What is the name of this compound?

A. 3-methyl-4-ethylhept-1-yne
B. Decyne
C. 4-ethyl-3-methylhept-1-ene
D. 3-methyl-4-propylhex-1-ene

8.6 Exercise 2: Short answer questions

1. a. What is meant by the term ‘hydrocarbon’?
b. Give the names and formulas of three hydrocarbons each containing three carbon atoms.

2. Explain why the compounds methene and methyne do not exist.

3. a. How are alkanes ‘produced’?
b. What are the major uses of alkanes?

4. Define the term ‘isomer’. Illustrate your answer with the isomers of pentane and butene.

5. a. Distinguish between saturated and unsaturated hydrocarbons.
b. Which is more reactive?

6. Write a formula for each of the following.
a. An alkane with 22 carbon atoms
b. An alkene with 17 carbon atoms
c. An alkyne with 13 carbon atoms

7. The alkanols form a homologous series. What does this mean?

8. Briefly explain the meanings of the following terms. Use examples in your answers.
   a. Substitution reactions
   b. Addition reactions
   c. Functional group
   d. Fractional distillation
   e. Catalytic cracking

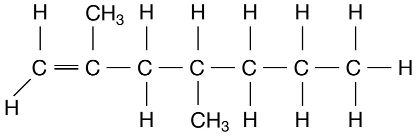
9. Name the esters formed from the following reactions.
   a. Methanol and butanoic acid
   b. Propan-1-ol and propanoic acid
   c. Ethanol and methanoic acid

10. Write formulas for the following esters.
    a. Ethyl butanoate
    b. Propyl ethanoate
    c. Butyl propanoate

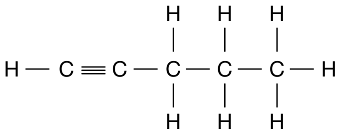
11. A molecular compound contains four carbon atoms. Write a condensed formula for the compound if it is the following.
    a. An alkene
    b. An alcohol
    c. A carboxylic acid
    d. An ester

12. Write condensed (semi-structural) formulas for each of the following.

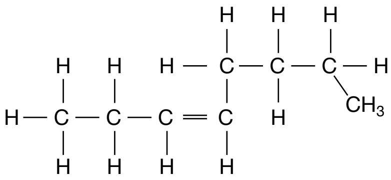
a.



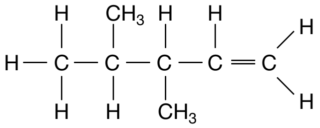
b.



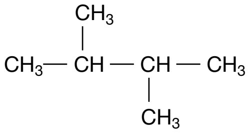
c.



d.

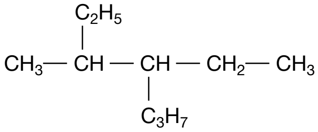


e.



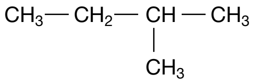
$CH_3-\overset{CH_3}{\overset{|}{CH}}-\underset{\underset{CH_3}{|}}{CH}-CH_3$

f.

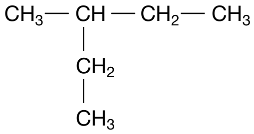


$CH_3-\overset{C_2H_5}{\overset{|}{CH}}-\underset{\underset{C_3H_7}{|}}{CH}-CH_2-CH_3$

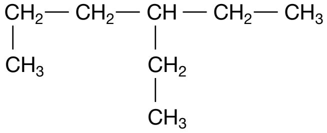
13. Name the following organic compounds.
(*Note:* the longest continuous chain may not be written in a straight line.)

a. 

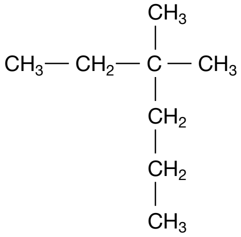
$CH_3-CH_2-\underset{\underset{\displaystyle CH_3}{|}}{CH}-CH_3$

b. 

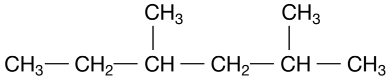
$CH_3-\underset{\underset{\underset{\underset{\displaystyle CH_3}{|}}{CH_2}}{|}}{CH}-CH_2-CH_3$

c. 

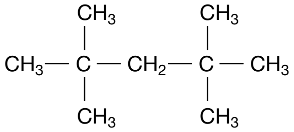
$\underset{\underset{\displaystyle CH_3}{|}}{CH_2}-CH_2-\underset{\underset{\underset{\underset{\displaystyle CH_3}{|}}{CH_2}}{|}}{CH}-CH_2-CH_3$

d. 

$CH_3-CH_2-\overset{\overset{\displaystyle CH_3}{|}}{\underset{\underset{\underset{\underset{\underset{\underset{\displaystyle CH_3}{|}}{CH_2}}{|}}{CH_2}}{|}}{C}}-CH_3$

e. 

$CH_3-CH_2-\overset{\overset{\displaystyle CH_3}{|}}{CH}-CH_2-\overset{\overset{\displaystyle CH_3}{|}}{CH}-CH_3$

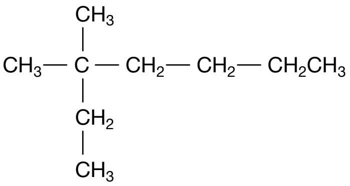
f. 

$CH_3-\overset{\overset{\displaystyle CH_3}{|}}{\underset{\underset{\displaystyle CH_3}{|}}{C}}-CH_2-\overset{\overset{\displaystyle CH_3}{|}}{\underset{\underset{\displaystyle CH_3}{|}}{C}}-CH_3$

14. Draw structural formulas for the following.

a. 2-methylheptane
b. 2, 3-dimethylbut-2-ene
c. Pent-2-yne
d. 2, 6-dimethylhept-3-ene
e. Oct-3-ene
f. 4-ethyl-3-methylhex-2-ene
g. 3, 4-diethyl-3-methylhexane
h. 3-ethyl-4,5-dipropyloctane

15. Jock isolated the following compound and named it 2-ethyl-2, 4-dimethylhexane. Lillian said that the compound was incorrectly named. Explain why Lillian is right and give the correct name for the compound.



$CH_3-\overset{\overset{\displaystyle CH_3}{|}}{\underset{\underset{\underset{\underset{\displaystyle CH_3}{|}}{CH_2}}{|}}{C}}-CH_2-CH_2-CH_2CH_3$

16. Draw structural formulas for the following compounds.

a. The paint remover dichloromethane
b. The water repellent used in Scotchgard, 1, 1, 1-trichloroethane

17. Use a labelled diagram of a methanol molecule and a water molecule to explain why methanol is soluble in water.

18. An alcohol containing only the elements carbon, hydrogen and oxygen was shown by analysis to contain 60.0% carbon and 13.3% hydrogen. Its molar mass is 60.0 g mol^{-1}. Identify the alcohol and draw possible structures.

19. In a chemical laboratory, the labels have fallen off bottles of two colourless liquids stored close together in the flammable liquids section.

    In an effort to identify them, the following data was obtained through experiment.
    Both react with ethanoic acid to produce pleasant smelling products.
    Both have the same elemental analysis: carbon (60%), hydrogen (13.3%) and oxygen (26.7%).
    Both have a molar mass of 60 g mol^{-1}.
    They have different boiling points.

    a. Calculate the empirical formulas of the two liquids.
    b. Calculate the molecular formulas of both liquids.
    c. To what class of organic compounds do these two liquids belong? Explain.
    d. From your answer to parts b and c, draw and name all the possible structural formulas.
    e. How would you use the experimental information given to distinguish between the possibilities that have been identified in part d?

8.6 Exercise 3: Exam practice questions

Question 1 (6 marks)

a. In terms of bonding, explain why heavier hydrocarbons such as diesel oil (those with a high number of carbon atoms) have higher boiling temperatures than lighter hydrocarbons such as propane (those with fewer carbon atoms). **2 marks**

b. In terms of bonding, explain why lighter hydrocarbons such as petrol are more flammable than heavier ones such as paraffin wax. **2 marks**

c. Propane and ethanol contain molecules that are nearly the same size. Explain why ethanol is a liquid at room temperature and propane is a gas. **2 mark**

Question 2 (5 marks)

Analysis of a certain compound containing carbon, hydrogen and oxygen revealed that it contained 40% carbon and 6.7% hydrogen.

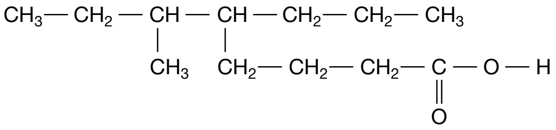
a. Calculate the empirical formula of the compound. **2 marks**

b. Calculate its molecular formula given that its molar mass is 60g mol^{-1}. **1 mark**

c. Draw a possible structural formula for this compound. **2 marks**

Question 3 (5 marks)

An organic compound has the following formula.



$$\begin{array}{ccccccccccccccccc}
CH_3 & - & CH_2 & - & CH & - & CH & - & CH_2 & - & CH_2 & - & CH_3 & & & & \\
 & & & & | & & | & & & & & & & & & & \\
 & & & & CH_3 & & CH_2 & - & CH_2 & - & CH_2 & - & C & - & O & - & H \\
 & & & & & & & & & & & & \| & & & & \\
 & & & & & & & & & & & & O & & & &
\end{array}$$

a. What is the name of the homologous series to which this compound belongs? **1 mark**

b. Name this compound, using IUPAC nomeclature. **2 marks**

c. Draw the structure of one isomer of this compound. **2 marks**

Question 4 (4 marks)

If a mixture of lubricating oil, diesel oil, petrol, paraffin and kerosene were placed in a fractional distillation apparatus, answer the following.

a. Which fraction would be collected first? **1 mark**

b. Which would have the lowest boiling temperature? **1 mark**

c. Which would have the highest boiling temperature? **1 mark**

d. What could be done if the fractions collected were found to still be mixtures? **1 mark**

Question 5 (9 marks)

Analysis of a certain compound reveals that it contains only carbon, hydrogen and oxygen.

a. A small amount of crushed calcium carbonate is added to a sample of the compound. A gas is observed to form. Name the homologous series to which this compound belongs. **1 mark**

b. If you had not done the test described in part a, name two other homologous series that this compound might belong to. **2 marks**

c. A 9.217 g sample of this compound was combusted to yield 19.873 g of carbon dioxide and 8.127 g of water. Complete the following.
   i. Calculate the mass of carbon in the compound. **1 mark**
   ii. Calculate the mass of hydrogen in the compound. **1 mark**
   iii. Calculate the mass of oxygen in the compound. **1 mark**

d. From your findings, calculate the compound's empirical formula. **2 marks**

e. In a further experiment, the compund's molar mass was found to be 102 g mol^{-1}. Calculate its molecular formula. **1 mark**

8.6 Exercise 6: studyON Topic Test online only

Fully worked solutions and sample responses are available in your digital formats.

teachon

Test maker

Create unique tests and exams from our extensive range of questions, including practice exam questions.
Access the Assignments section in learnON to begin creating and assigning assessments to students.

AREA OF STUDY 2
HOW CAN THE VERSATILITY OF NON-METALS BE EXPLAINED?

9 Polymers

9.1 Overview

Numerous **videos** and **interactivities** are available just where you need them, at the point of learning, in your digital formats, learnON and eBookPLUS at www.jacplus.com.au.

9.1.1 Introduction

Originally, toothbrushes were made from the neck hairs of wild hogs, and these hairs were attached to either a bone or bamboo handle. But, thankfully, with the advent of modern plastics, getting a new toothbrush usually requires only a short trip to the local supermarket. Toothbrushes today are typically constructed with nylon bristles fixed to a polypropene (or, more recently, a bamboo) handle. Toothbrushes were one of the first commercial uses of nylon. Each of these materials has specific properties that make it suitable for this purpose. Unlike hog's hair, which is a natural polymer, both nylon and polypropene are synthetic polymers. What are polymers, and how are they customised and produced?

FIGURE 9.1 Toothbrushes are made of polymers.

This topic introduces the polymerisation processes by which complex organic molecules, called polymers, may be formed from simpler molecules. The different methods of polymer production are described, as is how polymer properties can be altered by different processes and additives. The advantages and disadvantages of the uses of polymers are discussed, as well as which plastics can be recycled to form other useful materials.

This topic builds on prior knowledge of hydrocarbons, especially alkenes, and the effects of molecular bonding — both the intramolecular bonding that exists within molecules and the types and properties of intermolecular bonding that exists between molecules.

9.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- the formation of polymers from monomers, including addition polymerisation of alkenes
- the distinction between linear (thermoplastic) and cross-linked (thermosetting) polymers, with reference to structure, bonding and properties, including capacity to be recycled
- the features of linear polymers designed for a particular purpose, including the selection of a suitable monomer (structure and properties), chain length, degree of branching, percentage crystalline areas and addition of plasticisers
- the advantages and disadvantages of the use of polymer materials.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of learning and assessment. Experiments and investigations, supported by a **Practical investigation logbook** and **Teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

on Resources

Digital documents Key science skills (doc-30903)
Key terms glossary — Topic 9 (doc-30955)
Practical investigation logbook (doc-30956)

studyon

To access key concept summaries and practice exam questions download and print the **studyON: Revision and practice exam question booklet** (doc-30957).

9.2 Formation of polymers

KEY CONCEPT

- The formation of polymers from monomers, including addition polymerisation of alkenes

9.2.1 Plastics and polymers

It is difficult to imagine, now that we have grown so used to plastics, what life was like without them. Many of the items we use every day, from objects as insignificant as pens and zippers to larger items such as furniture, refrigerators and the interiors of cars, are made from plastics. Even our clothes can be made from synthetic fibres. Many modern products are produced from a selection of plastics, each one chosen for its specific properties. The term ‘plastic’ refers to something that is pliable and/or able to be moulded, either in its final state or at some stage in its production. Many of what are called **plastic** and **synthetic fibres** in everyday life are what chemists call **polymers**.

The name polymer comes from the Greek *polymeres* (‘of many parts’). A polymer molecule is made up of thousands of units strung together into very long chains. The simple molecules that are strung together are called **monomers**. Monomers link together to form a polymer chain in a process called **polymerisation**. The monomers used to make the polymer can be the same or different. **Copolymers** are formed when two or more different monomers are used. Most polymerisation reactions require a catalyst. Polymers may be either natural or synthetic.

Wool, cotton, linen, hair, skin, nails, rubber and flesh are all naturally occurring polymers. Most natural polymers are made of proteins or cellulose. Cotton is nearly all cellulose and hair is a protein polymer.

Synthetic polymers are made by either of two main methods: **addition polymerisation** or **condensation polymerisation**. Some examples of these, along with their method of production, are given in table 9.1.

9.2.2 Addition polymers

For addition polymerisation to occur, the monomers must contain a carbon to carbon double bond. That is, the monomer must be an alkene. In the process of joining the monomers together, one of these bonds is broken and the electrons that are freed up are used to form a new bond to the next monomer. This links the monomers together and forms the long chains. Alkenes, and especially the smaller members such as ethene and propene, are thus ideal candidates to act as monomers.

TABLE 9.1 Common polymers and their uses

Polymer	Type of polymerisation	Use
Polyethene	Addition	Cling film, squeeze bottles, milk crates
Polypropene (polypropylene)	Addition	Mixing bowls, ice-cream containers, moulded chairs
Polyvinyl chloride (PVC)	Addition	Water piping, upholstery covering, toys, flexible laboratory tubing, suitcases, packaged meats, juice bottles, credit cards, shampoo bottles
Nylon	Condensation	Hosiery, carpets, clothing, bearings, fishing lines, tennis racquet strings, toothbrush bristles, mascara brush bristles
Polystyrene	Addition	Foam for insulation, plates, cups, trays, packaging, surfboards, beanbag filling, refrigerator and washing machine components
Polytetrafluoroethene (PTFE, Teflon®)	Addition	Washers, coating on frying pans, razor blades, skis
Polymethyl methacrylate (Perspex®)	Addition	Advertising signs, telephones, light fittings, aeroplane windows
Polyethene terephthalate (PET)	Condensation	Soft drink bottles, other clear plastic bottles, hard wearing carpets, as a polyester filling (for pillows, weather jackets, etc.)
Urea formaldehyde or phenol formaldehyde (Bakelite®)	Condensation	Backs of television sets, ash trays, vacuum jugs
Epoxy resin (Araldite®)	Condensation	Handles of screwdrivers
Melamine formaldehyde (Formica®)	Condensation	Kitchen cupboards and benches
Polyurethane	Condensation	Foam rubber, pillow filling, packaging

Polymers are named by placing the prefix ‘poly’ in front of the monomer. In many (but not all) cases, the older or trivial name is used for the monomer instead of the IUPAC name.

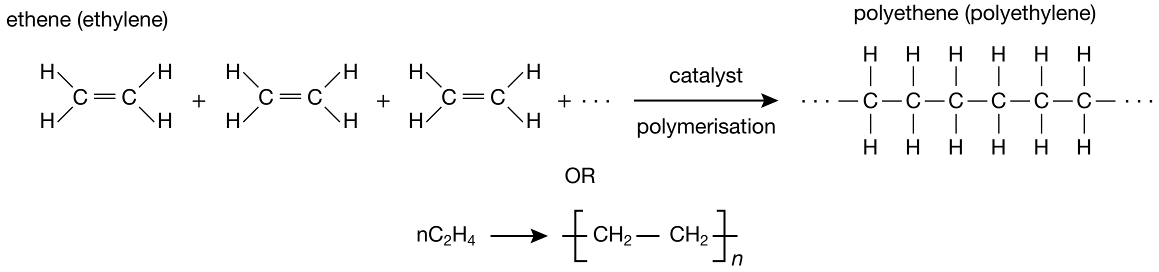
The simplest monomer that can undergo addition polymerisation is ethene, which results in the polymer polyethene.

When ethene is subjected to high pressure, it changes from a gas to a liquid. If the liquid ethene, still under pressure, is heated in the presence of a catalyst (a small quantity of oxygen), an **addition reaction** takes place in which the ethene molecules join together and form a long chain (**polyethene** or poly-ethylene). The length of these chains can vary from 4000 to 20 000 carbon atoms, depending on the temperature and pressure used.

Addition polymers are formed from monomers, which contain a carbon–carbon double bond. The bond is broken and a new bond forms with the next monomer, creating very long polymer chains.

FIGURE 9.2 Milk bottles are made of high-density polyethylene (HDPE).

FIGURE 9.3 The formation of the polymer polyethene. The double bond between the two carbon atoms in the monomer breaks, allowing long chains to form.



The second representation of the reaction provided in figure 9.3 is called the condensed or bracketed form. It can be seen that the empirical formula of the polymer is the same as the monomer.

It can be seen that polymers are very large, chain-like molecules. Due to their large size, they will have very high molar masses. It is not uncommon for these values to be in the region of 500 000 g mol^{-1}.

Many other addition polymers exist — all of which are formed from monomers containing a carbon–carbon double bond. These are dealt with later in this topic.

SAMPLE PROBLEM 1

Calculate the approximate molar mass of a polyethene sample, the average chain length of which is 35 000 carbon atoms.

Teacher-led video: SP1 (tlvd-0553)

THINK	WRITE
1. Recall that in addition polymerisation, all the atoms in the original monomers are used in the production of the resulting polymer (nothing is lost).	Carbon atoms = 35 000
2. Recall that the molecular formula of ethene is C_2H_4. Hence, two carbon atoms exist per monomer of ethene.	Number of ethene monomers $= \frac{35\,000}{2}$ $= 17\,500$
3. Determine the approximate molar mass of the polyethene sample by multiplying the number of ethene monomers by the molar mass of ethene.	Molar mass ethene $= 2 \times \text{C} + 4 \times \text{H}$ $= 2 \times 12.0 + 4 \times 1.0$ $= 28.0 \text{ g mol}^{-1}$ Approximate molar mass $= 17\,500 \times 28.0$ $= 490\,000 \text{ g mol}^{-1}$

PRACTICE PROBLEM 1

Calculate the approximate molar mass of a polyethene sample, the average chain length of which is 5000 carbon atoms.

on Resources

Video eLesson Addition polymerisation (eles-2481)

CONDENSATION POLYMERS

Condensation polymers form from monomers that have *two functional groups per molecule*. Such molecules are said to be **difunctional**. Polymerisation occurs when one functional group of one monomer reacts with one functional group on another. A small molecule (which is often, but not always, water) is removed in this process. Because each monomer has one more functional group on it (they are difunctional), this process can be repeated on and on to produce the long polymer chains.

Many natural polymers such as proteins and polysaccharides are examples of condensation polymerisation. Common synthetic polymers produced using this method include nylon, polyester and polyethene terephthalate (PET) plastic. As you might expect, many of these reactions show similarities to the esters introduced in topic 8.

Nylon

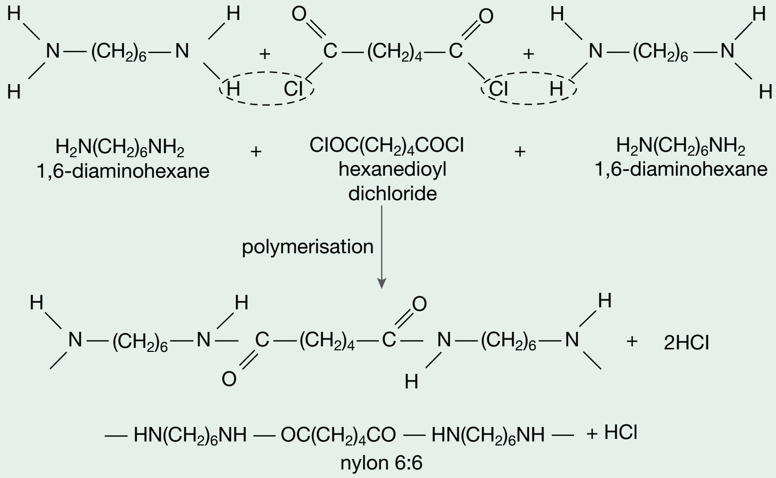
Nylon can be extruded when molten to form fibres or sheets of strong, durable and elastic material. It is used to make fabrics and other products, such as the bristles of brushes and velcro. Its invention had a great impact on the textile and clothing industry.

Nylon 6:6 is one type of copolymer formed by the condensation polymerisation of two monomers: hexanedioyl (adipyl) chloride and 1,6-diaminohexane. An amine group in 1,6-diaminohexane reacts with a chloride group in adipyl chloride, forming a bond between the two molecules and eliminating a molecule of hydrochloric acid. Further molecules add on in the same way to build up a long chain.

Nylon is a linear chain containing up to 100 repeated units. The name ‘nylon 6:6’ refers to the existence of 6 carbon atoms on each of the units. You will notice that a repeated linkage pattern exists consisting of —NHOC—units. This is called the amide group and results when the small molecule HCl is eliminated from the two functional groups during the condensation reaction. Nylon is classed as a polyamide fibre. This same linking group is found in the protein polymer chain that makes up wool, although there it is referred to as a peptide link.

FIGURE 9.4 Electron micrograph of velcro, which is made of nylon. Velcro was inspired by burrs on plants that attach themselves to passing objects through thousands of tiny hooks.

FIGURE 9.5 Nylon 6:6



A modification of this manufacturing process uses haxanedicarboxylic acid, $HOOC(CH_2)_4COOH$, in place of hexanedioyl dichloride. This produces exactly the same nylon product, but has the advantage that the small molecule removed is water.

PET (or PETE) plastic

Plastic soft-drink bottles are a good example of an article made from **PET** polymer. Its full name is polyethene terephthalate. It has also been used to make fibres and is often included in carpets designed for high traffic areas. It is an example of a **polyester**.

To make PET, terephthalic acid (also known as benzene-1,4-dicarboxylic acid) and ethylene glycol (also known as ethane diol) are reacted together. The structures of these two compounds are shown in figure 9.6.

The reaction is shown in figure 9.7

FIGURE 9.6 The structures of terephthalic acid and ethylene glycol

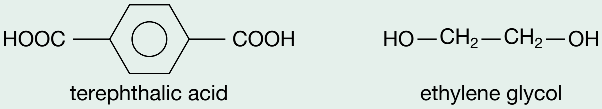
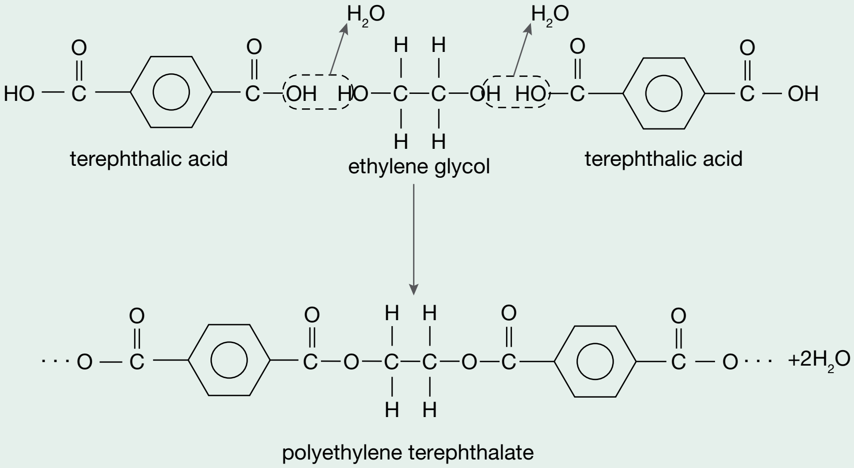


FIGURE 9.7 The reaction between terephthalic acid and ethylene glycol



Note the removal of the small molecule (water), which makes this another example of a condensation polymerisation.

on Resources

Weblink Nobel Prize in Chemistry for conductive polymers

9.2 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Explain why alkenes are able to act as monomers for addition polymerisation, but alkanes are not.
2. Give examples of the following.
   (a) Three natural polymers
   (b) Three synthetic polymers
3. Calculate the approximate molar mass of a polyethene sample, the average chain length of which is 50 000 carbon atoms.
4. Ethene is a gas that can be stored for long periods of time. Explain why it can be stored as a gas without it polymerising into polyethene.
5. A sample of polyethene is found to have an approximate molar mass of 450 000 units.
   (a) Calculate the approximate number of monomer units that are joined to make the molecules in this sample.
   (b) What is the approximate chain length in terms of number of carbon atoms?
6. Are alkynes able to act as monomers for addition polymerisation? Explain.

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9.3 Linear and cross-linked polymers

KEY CONCEPT

- The distinction between linear (thermoplastic) and cross-linked (thermosetting) polymers, with reference to structure, bonding and properties, including capacity to be recycled

9.3.1 Thermoplastic and thermosetting polymers

Plastics may be classified in a number of ways. One of these, as we have already seen, is on the mechanism of their manufacture; that is, whether they are addition polymers or condensation polymers. Another method is on the basis of their thermal behaviour. This method provides insight into their uses and applications as well as their ability to be recycled. Plastics that may be repeatedly melted, reshaped and hardened by cooling are called **thermosoftening** plastics or **thermoplastics**. An example of a thermoplastic polymer is polystyrene. Plastics that do not melt but char when heated are called **thermosetting** plastics. These plastics must be moulded or shaped during their manufacture. Bakelite is an example of a thermosetting plastic material. Table 9.2 lists a range of common polymers classified as to their thermal behaviour.

TABLE 9.2 The reaction of some common polymers to heat

Polymer	Reaction to heat	Classification
Polyethene	Melts/softens	Thermoplastic
Polypropene	Melts/softens	Thermoplastic
Polyvinyl chloride	Melts/softens	Thermoplastic
Nylon	Melts/softens	Thermoplastic
Polystyrene	Melts/softens	Thermoplastic
Polytetrafluoroethene (PTFE, Teflon®)	Melts/softens	Thermoplastic
Polymethyl methacrylate (Perspex®)	Melts/softens	Thermoplastic
Polyethene terephthalate (PET)	Melts/softens	Thermoplastic
Urea formaldehyde or phenol formaldehyde (Bakelite®)	Chars	Thermosetting
Epoxy resin (Araldite®)	Chars	Thermosetting
Melamine formaldehyde (Formica®) — condensation	Chars	Thermosetting
Polyurethane	Chars	Thermosetting

Thermoplastic and thermosetting polymers obviously have very different properties. These can be explained if we look closer at the long polymer chains themselves and the forces and other physical factors that operate between them.

THE FIRST PLASTICS

FIGURE 9.8 The casings and dials of early telephones, radios and many other appliances were frequently produced from Bakelite.

The first plastics were developed in the middle of the nineteenth century, following a competition set up by a manufacturer of billiard balls. In those days, billiard balls were made of ivory and the manufacturer, Phelan and Collander, offered a substantial prize for the discovery of a satisfactory alternative. One of the entrants in the competition was John Wesley Hyatt, who developed a substance called celluloid. He did not win the prize, as celluloid is volatile, and the billiard balls would have exploded on impact. However, celluloid was the first thermoplastic — a substance moulded using heat and pressure that retains its shape once cooled — and celluloid became a widely used polymer. Its applications included dental plates and men's collars, but it is most identified with its role in the early photographic and motion picture industries, where it was used to make film.

In 1907 the first completely synthetic polymer, called Bakelite, was produced by Belgian-American chemist Leo H. Baekeland (1863–1944). Bakelite, unlike celluloid-based plastics, is a thermosetting polymer and cannot be softened by heat once it has set. It is also resistant to common acids and solvents. These properties made it virtually indestructible, and Bakelite was a landmark in the history of plastics. Its discovery started a large plastics industry and the age of plastics had begun.

9.3.2 Linear polymers

Linear polymers are thermoplastic polymers. As their name suggests, linear polymers are polymers that form as long chains. In any sample of a linear polymer, huge numbers of such chains will be present and these, due to their length, will invariably become tangled. As we have also seen in previous topics, intermolecular forces operate between these chains because each chain is, in effect, a very large molecule. These forces between the chains, along with their entanglement, will give a degree of rigidity to the overall polymer. The stronger the forces and the higher the entanglement, the harder the polymer will be. The weaker the forces and the lesser the degree of entanglement, the softer and more flexible it will be.

When heat is applied to such polymers, the weak intermolecular forces between the chains are, at least partially, overcome, and the chains are able to move relative to one another. This results in the overall polymer becoming softer and more liquid-like. Hence it will be able to be moulded into new shapes, which will once again 'set' upon cooling as the intermolecular forces between the chains once again take effect. This process of melting and softening is additionally assisted by the increased thermal vibration of the long chain polymer molecules with increasing temperature. This also assists them to slide past each other and form new shapes.

Therefore, the observed properties of thermoplastic polymers — that is, their ability to be melted and softened and then moulded into new shapes, and the variety that they display in terms of flexibility and rigidity — can all be explained in terms of their long linear polymer chains and how these chains interact with one another.

- **Linear polymers are thermoplastic polymers.**
- **Thermoplastic polymers may be repeatedly melted and reshaped and are hardened by cooling.**
- **Heating of thermoplastic polymers overcomes the intermolecular forces between the long polymer chains.**
- **The length and degree of entanglement of the chains affects their properties.**

9.3.3 Cross-linked polymers

Polymers can also be produced where the long chains are formed first and then deliberately connected together by strong covalent bonds. This process is called **cross-linking**. This means that the forces that operate between the chains are now strong covalent forces rather than the weak intermolecular forces seen in linear polymers.

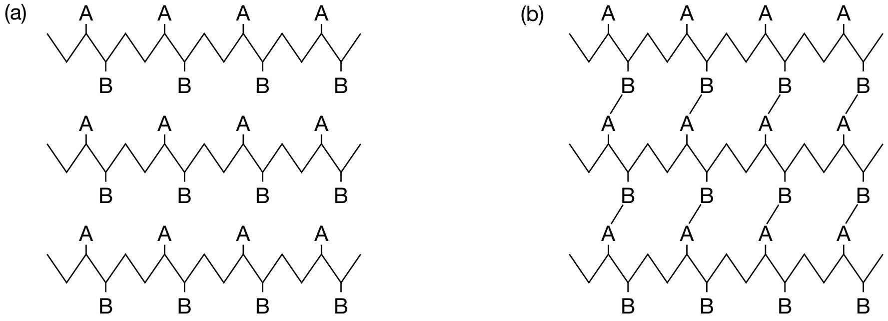
If only a small amount of crosslinking is introduced, an **elastomer** results. As the name suggests, elastomers can be stretched because the chains can still move, but in a restricted manner, past each other. Once the stretching force is removed, however, the cross-links pull the polymer back into its original shape. Examples of elastomers include vulcanised rubber, where a small amount of cross-linking using sulfur produces disulfide bonds between the chains.

Thermosetting polymers

If the degree of cross-linking within a polymer is large, then, as expected, the entire arrangement will be very rigid and strong due to the covalent bonding that now exists in all dimensions. Such polymers will not soften or melt when heated but can char at high temperatures. These are thermosetting polymers, and can be referred to as thermosets.

To make a thermosetting polymer, long linear chains with functional groups attached to them are produced first. The cross-linking is then brought about either by using heat or by adding a chemical to react between the lateral functional groups linking the chains together.

FIGURE 9.9 (a) A collection of linear polymers with lateral functional groups A and B. (b) When the A and B groups are linked, a rigid cross-linked polymer results.



For example, a particular article may be made by placing the preliminary linear polymer into a mould and then heating it. Cross-links form, producing a rigid thermosetting article in the mould. Note that thermosetting polymers may still be classed as plastics because, in the early stages of their manufacture (before the cross-linking is introduced), they can still be moulded.

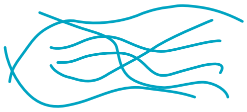
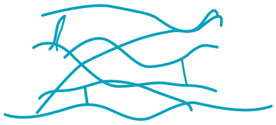
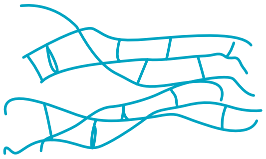
A familiar example of the second method of producing cross-links is when two-part glues are mixed. One tube contains the preliminary linear polymer with the lateral functional groups. The other tube contains the 'hardener'. When mixed, this chemical reacts with the lateral functional groups, bonding them together and linking the chains together with strong covalent cross-links.

- **Cross-linked polymers are thermosetting polymers.**
- **Cross-linked polymers are created by long polymer chains with functional groups subsequently joined by strong covalent bonds.**
- **Large amounts of cross-linking create rigid, strong compounds due to covalent forces.**

9.3.4 Summary of thermoplastics, elastomers and thermosets

Table 9.3 provides a summary of the different properties of thermoplastics, elastomers and thermosets.

TABLE 9.3 Comparing thermoplastics, elastomers and thermosets

Property	Thermoplastics	Elastomers	Thermosets
Effect of heat	Softens/melts	Varies	Does not soften/melt; chars
Number of cross-links	Nearly zero	A small number	A much larger number
Forces between chains	Weak intermolecular forces; often dispersion forces and/or hydrogen bonding	Weak intermolecular forces and some strong covalent bonds	Strong covalent bonds
Hardness	Variable	Varies according to degree of cross-linking	Usually hard
Capacity for recycling	A number are ideally suited	Limited	Very limited
Diagrammatic representation	 thermoplastic	 elastomer	 thermoset

9.3.5 Recycling polymers

The different structures of thermoplastic polymers and thermosetting polymers explain their different properties. A further consequence of this is their ability to be recycled. As a general rule, thermosetting polymers are currently not able to be recycled, whereas some (but not all) thermoplastic polymers can be.

FIGURE 9.10 Sorting plastic by type allows recycling of many plastics.

Several environmental and ethical factors need to be considered in the production of polymers. These issues include their production from finite resources, particularly crude oil, and their long-lasting and typically non-biodegradable nature. The latter issue means polymers are a severe litter and pollution problem, particularly in marine environments and waterways. Furthermore, despite their durability, they are typically single-use products. The ability to make polymers cheaply, however, and their function as hygienic, light-weight and strong materials has ensured our ongoing reliance on polymers.

These significant concerns have led to ongoing research into ways and methods to improve the recyclability of both classes of polymers.

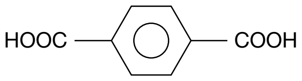
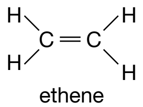
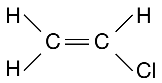
The ability of thermoplastic polymers to be softened and melted by heat and then remoulded is ideal for recycling purposes. However, some polymers such as PVC produce toxic chemicals when heated and this limits their ability to be recycled.

At the time of writing, three main approaches to the recycling question are used. The first approach is to sort the plastic by type and then recycle it in the same form. Many plastics, including two forms of polyethene (low-density polyethene, LDPE, and high-density polyethene, HDPE) and PET are recycled this way.

Once the plastics have been separated, they are shredded into flakes and washed to remove contaminants. The flakes are melted and extruded and then shredded again into flakes, which can then be remoulded into new products.

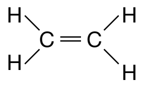
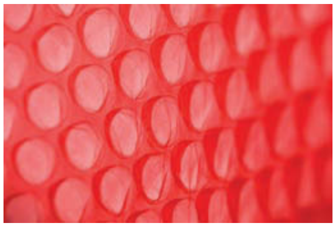
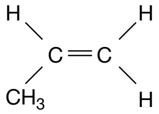
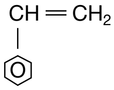
Recycled HDPE is used for compost bins, detergent bottles, pipes, plumbing fittings, household bags and irrigation pipes. Recycled LDPE is used for carry bags, packaging film and household bags.

TABLE 9.4 Recyclable plastics

Symbol and name	Monomer	Examples
♳ PETE Polyethene terephthalate	 terephthalic acid HO—CH_2—CH_2—OH ethylene glycol	Soft drink and water bottles 
♴ HDPE High-density polyethene	 ethene	Milk, juice and water bottles 
♵ V PVC	 chloroethene vinyl chloride	Juice bottles, detergent bottles, credit cards, PVC piping 

(Continued)

TABLE 9.4 Recyclable plastics *(Continued)*

Symbol and name	Monomer	Examples
4 LDPE Low-density polyethene	 ethene	Frozen food bags, bin liners, squeezable bottles, flexible container lids, cling film, bubble wrap 
5 PP Polypropene	 propene	Reusable microwave containers, margarine tubs, kitchen ware, yoghurt containers, disposable cups and plates 
6 PS Polystyrene	 styrene	Egg cartons, packing 'peanuts', disposable cups, plates, trays and cutlery, disposable takeaway containers 
7 OTHER Other		Beverage bottles, baby milk bottles, electronic casing 

The second approach to recycling is the co-mingled process. This type of recycling produces products from co-mingled polymers that do not have to be sorted as rigorously as in the first method, and may also include small amounts of polymers that cannot be recycled by the first method. Co-mingled recycling is used to make plastic posts, poles and other articles that can be used instead of wood.

Another approach being developed is to depolymerise the polymer. This process can either regenerate the original monomers (which can subsequently be repolymerised) or can produce other products, some of which can be used as fuels. To date, this method is energy intensive and is not widely used.

on Resources

Weblink Government action on plastics and packaging

9.3 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Explain why the bonding between chains is stronger in thermosetting polymers than in thermoplastic polymers.
2. Describe the difference(s) between thermoplastic polymers, elastomers and thermosetting polymers in terms of the following.
   (a) Bonding
   (b) Properties
3. (a) In general terms, describe the steps in the production of an object made from a thermosetting polymer.
   (b) Why are thermosetting polymers still described as plastics, even though they cannot be melted and remoulded?
4. A number of common thermoplastics are recycled in large amounts, whereas articles made from thermosetting polymers are rarely recycled. Explain why this is so.
5. Two linear polymers, *A* and *B*, are of roughly equal molar mass. Hydrogen bonding exists between the chains in A whereas dispersion forces are the only forces operating between the chains in *B*.
   How will the properties of polymer *A* compare to polymer *B*?

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9.4 Polymer selection

KEY CONCEPTS

- The features of linear polymers designed for a particular purpose, including the selection of a suitable monomer (structure and properties), chain length, degree of branching, percentage crystalline areas and addition of plasticisers
- The advantages and disadvantages of the use of polymer materials

9.4.1 Features of linear polymers

In 1963, the Nobel Prize–winning chemist Giulio Natta said, ‘A chemist setting out to build a giant molecule is in the same position as an architect designing a building. He has a number of building blocks of certain shapes and sizes, and his task is to put them together in a structure to serve a particular purpose.' Today we can take this statement to mean that, due to their versatility, polymers can be produced for almost any imagined purpose.

The versatility of polymers is due to their many different properties. They can be hard or soft, flexible or rigid, transparent or coloured, brittle or able to be stretched. These properties of polymers are affected by the:

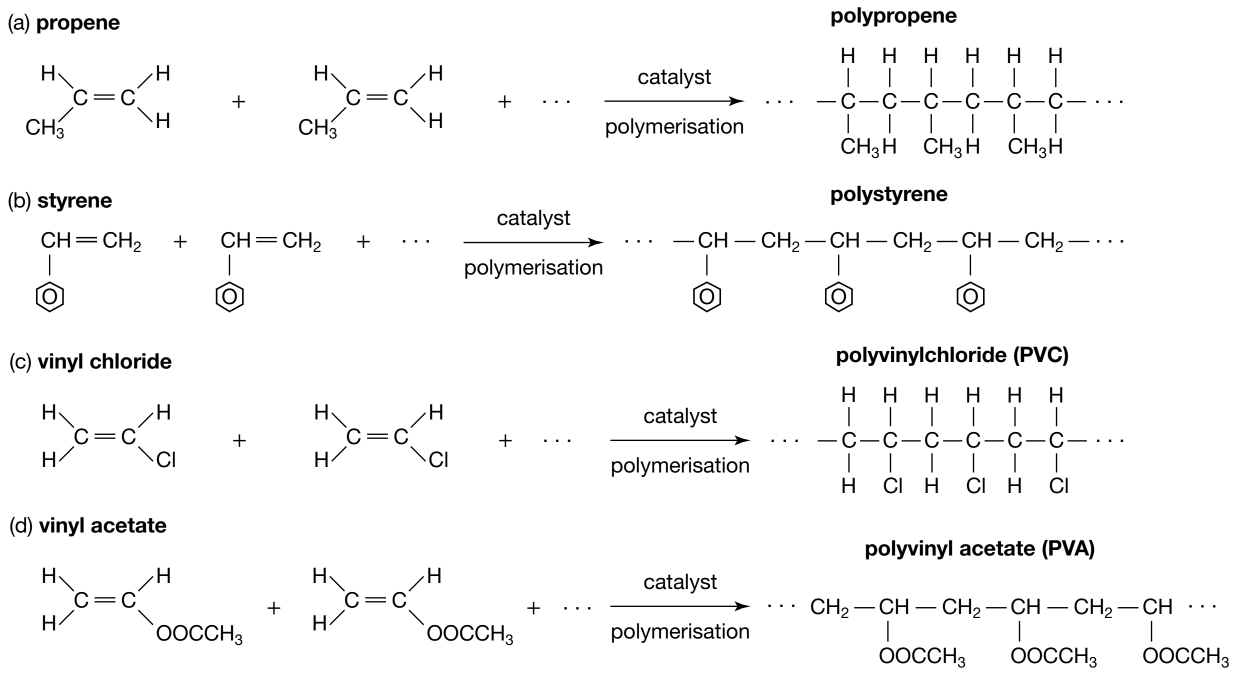
- choice of monomer
- degree of branching
- degree of crystallinity
- length of the polymer chains
- addition of plasticisers
- additives.

Choice of monomer

As seen earlier in this topic, ethene can be polymerised to produce polyethene. The critical feature that enables this to occur is the presence of the carbon–carbon double bond in the monomer, ethene. By using ethene and substituting other atoms or groups of atoms in place of hydrogen atoms, molecules can be produced in which the double bond is still intact. These molecules will, therefore, be able to act as monomers by which a further range of polymers with different properties may be produced. Figure 9.11 shows this process for a number of commonly used polymers. You will note that, as is the case for ethene, the method of polymerisation is addition polymerisation and that all monomers contain a carbon–carbon double bond.

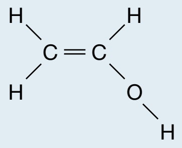
To create an addition polymer, all monomers must contain a carbon–carbon double bond.

FIGURE 9.11 The formation of (a) polypropene from propene, (b) polystyrene from styrene, (c) polyvinyl chloride (PVC) from vinyl chloride and (d) polyvinyl acetate (PVA) from vinyl acetate



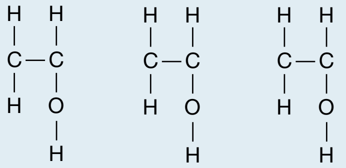
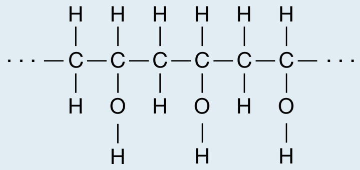
SAMPLE PROBLEM 2

Polyvinyl alcohol has many uses. Play 'slime' and the water-soluble plastic used in laundry and dishwashing pods are just two of these. It is made from vinyl alcohol, the structure of which is shown.



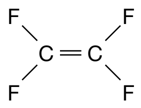
Draw a section of the polymer that is produced when this alcohol is polymerised to make polyvinyl alcohol.

Teacher-led video: SP2 (tlvd-0554)

THINK	**WRITE**
1. Draw a number of vinyl alcohol molecules next to each other, but with only a single bond between the carbons.	
2. Place a new bond between the molecules to link them together. Show three dots at each end to signify only a section is shown.	 $\begin{array}{cccccccc} & H & H & H & H & H & H & \\ & \vert & \vert & \vert & \vert & \vert & \vert & \\ \cdots - & C - & C - & C - & C - & C - & C & - \cdots \\ & \vert & \vert & \vert & \vert & \vert & \vert & \\ & H & O & H & O & H & O & \\ & & \vert & & \vert & & \vert & \\ & & H & & H & & H & \end{array}$
Alternatively, you can show the answer in condensed form.	$+CH_2 - CHOH+_n$

PRACTICE PROBLEM 2

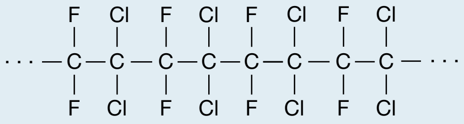
Teflon® is made from polytetrafluoroethene (PTFE). The monomer used to make this polymer is shown.



Draw a section of the polymer that is produced when this monomer is polymerised to make PTFE.

SAMPLE PROBLEM 3

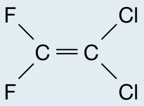
A section of a polymer produced at a research laboratory has the structure shown.



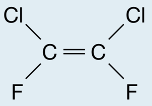
$$\begin{array}{cccccccccc} & F & Cl & F & Cl & F & Cl & F & Cl & \\ & \vert & \vert & \vert & \vert & \vert & \vert & \vert & \vert & \\ \cdots - & C - & C - & C - & C - & C - & C - & C - & C & - \cdots \\ & \vert & \vert & \vert & \vert & \vert & \vert & \vert & \vert & \\ & F & Cl & F & Cl & F & Cl & F & Cl & \end{array}$$

Which of the following three possible monomers is it made from?

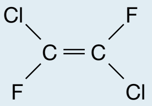
(a)



(b)



(c)



Teacher-led video: SP3 (tlvd-0555)

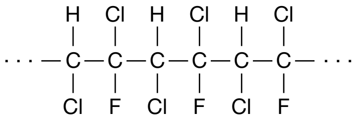
THINK	**WRITE**
1. The polymer shows alternating F and Cl atoms above and below the carbon backbone. It, therefore, cannot be (b).	

2. Each F is attached to the same carbon, as is each Cl. It cannot be (c).
Structure (a) is the only monomer that will produce the required pattern when polymerised.

Monomer (a) is the only monomer that will produce a polymer with the fluorine and chlorine atoms in the correct pattern.

PRACTICE PROBLEM 3

Draw the monomer that the following polymer is made from.



$$\begin{array}{cccccccccccccc} & & H & & Cl & & H & & Cl & & H & & Cl & \\ & & | & & | & & | & & | & & | & & | & \\ \cdots & - & C & - & C & - & C & - & C & - & C & - & C & - \cdots \\ & & | & & | & & | & & | & & | & & | & \\ & & Cl & & F & & Cl & & F & & Cl & & F & \end{array}$$

TIP: When drawing the structural formula for monomers and polymers, remember it is important to retain the position of the functional groups above or below the carbon backbone.

on Resources

Interactivity	Making polymers (int-3849)
Digital document	Experiment 9.1 Cross-linking an addition polymer to make slime (doc-30851)
Teacher-led video	Experiment 9.1 Cross-linking an addition polymer to make slime (tlvd-0626)

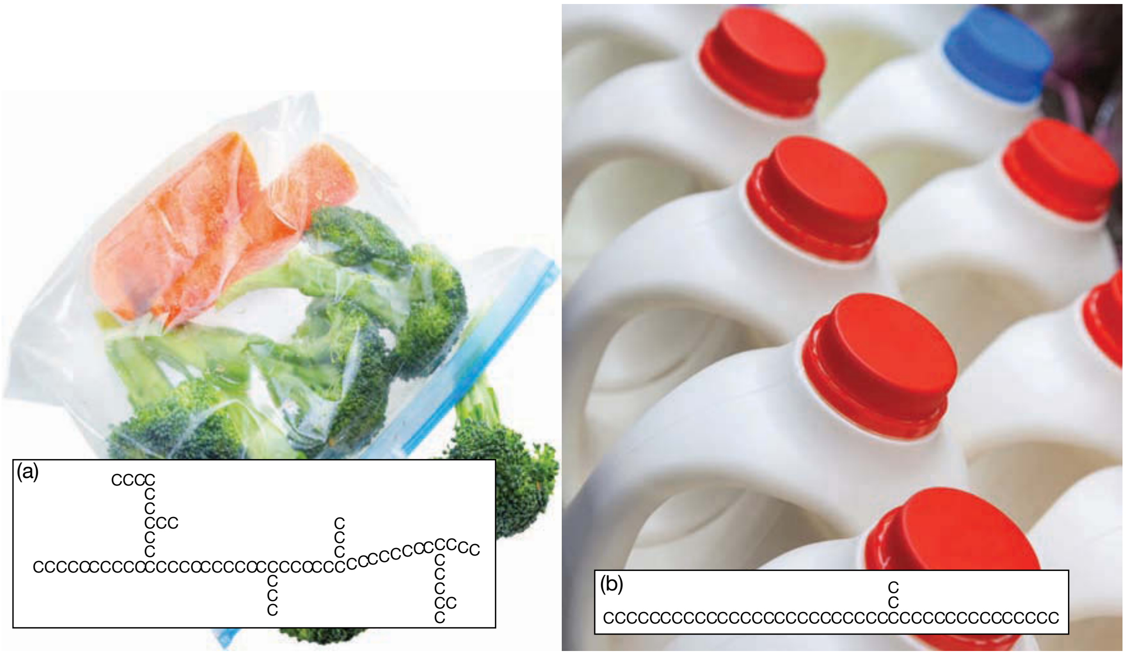
Degree of branching

In reality, polymerisation does not produce purely linear chains. For various reasons, branching occurs that causes side chains of various lengths to form. A linear polymer may, therefore, be more accurately thought of as branches on a tree. As a general rule, the more extreme the conditions required for polymerisation, the greater the number and length of such side branches. This phenomenon is illustrated by polyethene.

Ethene can be polymerised to produce both low- and high-density polyethene. Polyethene was discovered in 1933. Scientists had been experimenting with the effect of heat and very high pressures on ethene. It was not until 1939 that full commercial production of polyethene began. The first polyethene produced was called **low-density polyethene (LDPE)**, because its polymer chains support a large number of long side branches, producing a low-density substance (see figure 9.12a). Since the only forces causing these polymer chains to attract each other are dispersion forces, the effect of the branches is to keep the chains apart. Because the attraction becomes weaker as the chains are further apart, the density of the resultant compound is low, and LDPE is soft, flexible and translucent, with a waxy surface that repels water. Among its uses are plastic bags.

In the early 1950s, ethene was polymerised using lower pressure and lower temperatures. The result was a polymer of ethene with very few branches, and any branches that developed were short. Dispersion forces act more effectively on these chains because they can pack more closely together, and this type of polyethene is more rigid, stronger and more opaque than LDPE. It is slightly flexible and also has a waxy surface that repels water. It is called **high-density polyethene (HDPE)** and can be used to make bottles (see figure 9.12b).

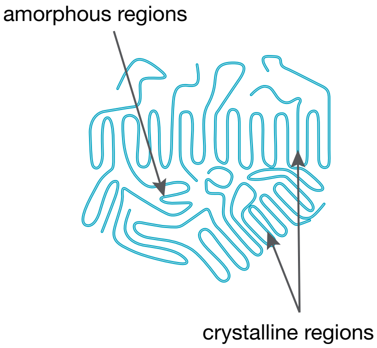
FIGURE 9.12 The carbon backbone in a portion of a polymer chain in (a) LDPE and (b) HDPE.



Degree of crystallinity

Polymer chains can be arranged in two ways. They can be crystalline, in that they are regularly organised into lines, or randomly packed with no particular order. Polymers can be partially crystalline or almost totally amorphous. The percentage of regularly ordered chains usually ranges between 10% and 80%. The more crystalline a structure is, the greater its hardness, tensile strength and opacity. This is because the chains are packed closer together in the crystalline regions, resulting in the dispersion forces that operate between then being more effective. This close packing also means that light waves are scattered rather than transmitted, resulting in opacity. Amorphous polymers are more easily deformed and often transparent. The factors that affect crystallinity include chain length, branching and interchain bonding, as seen previously for LDPE and HDPE. HDPE is more likely to conform to a crystalline structure because of its unbranched carbon chains.

FIGURE 9.13 The strength of a polymer is increased if more crystalline areas are present.



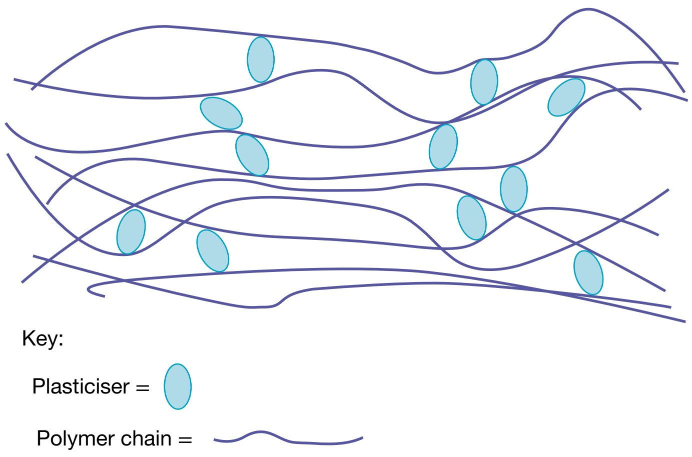
Length of polymer chain

Longer polymer chains are, in fact, larger molecules. This means that the dispersion forces between chains are stronger. Additionally, the longer polymer chains are, the more they become tangled and are less likely to slide over each other than smaller chains. Both of these features mean that it is harder for the chains to move with respect to one another. Consequently, polymers with longer chains tend to be harder/stronger and less flexible, and are able to withstand higher temperatures before they soften or melt.

Addition of plasticisers

Plasticisers are major components added to polymers such as PVC during production. They are small molecules that cause the polymer chains to move slightly further apart, resulting in a softer and more flexible polymer. Many different plasticisers are available but the most commonly used are phthalates.

FIGURE 9.14 Plasticisers are added to polymers to increase flexibility.



Other additives

Additives in polymers can have many different uses. Examples include the following:

- UV stabilisers, which absorb UV rays to prevent the polymer breaking down
- flame retardants to reduce the tendency of the polymer to burn
- dyes to add colour or provide patterns.

9.4.2 Examples of properties related to uses

The versatility of polymers results from the factors mentioned in the previous section, and from the degree of cross-linking that may or may not be present. This means a polymer may be chosen (or made) to suit nearly any purpose, based upon its properties. Table 9.5 gives just a few examples of how a polymer's properties make it useful for a particular application.

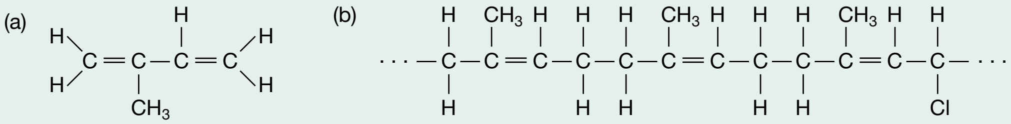
FIGURE 9.18 Many items, such as this in-line skate, are fabricated from a diverse range of plastic components. Each of these is chosen for its particular properties.

MODIFYING A NATURAL POLYMER — VULCANISING RUBBER

Natural **rubber** is produced from latex, a milky white addition polymer that can be harvested from rubber trees by making a cut in the bark and collecting the sap as it runs out. A well-managed plant can yield approximately 1.8 kg of dry crude rubber per year.

Rubber is a completely amorphous polymer. The monomer in natural rubber is **isoprene** or 2-methylbuta-1,3-diene. Isoprene polymerises to form long chains, as shown in figure 9.15, and the molecular formula is written as $(C_5H_8)_n$.

FIGURE 9.15 Structural formulas of (a) the monomer isoprene and (b) the polymer for raw (natural) rubber.



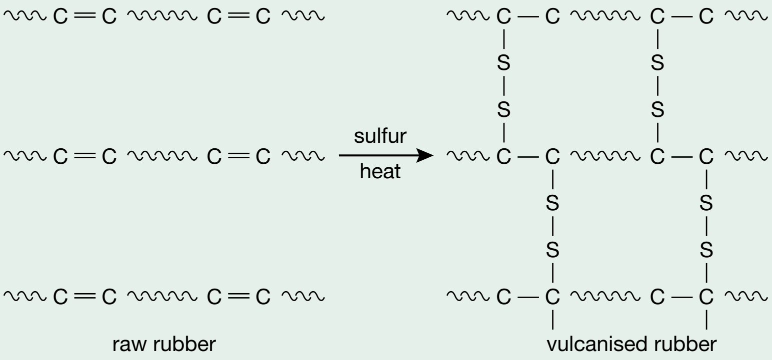
This chain is similar to that of polyethene (polyethylene) but with an important difference: rubber still contains double bonds that can be attacked by oxygen and, unlike polyethene, rubber can perish. Another disadvantage of natural rubber is that it is not elastic. When stretched, the long, tangled chains straighten out and remain this way, with little tendency for them to return to their original shape. Natural rubber is also susceptible to temperature changes, becoming very brittle when cold and sticky when hot.

An American inventor, Charles Goodyear, experimented for many years to find a way of countering these tendencies. In 1839, Goodyear accidentally dropped a piece of rubber that had been treated with sulfur onto a hot stove. He later found that this sample had improved elasticity and greater resistance to temperature change. His discovery formed the basis of the process of **vulcanisation**, which is still widely used in the rubber manufacturing industry today to improve the durability and elasticity of natural rubber.

FIGURE 9.16 Harvesting of natural rubber

Natural rubber is vulcanised in an industrial process, where it is mixed with sulfur and heated. The sulfur atoms form cross-links between chains of rubber molecules, reducing the number of double bonds, as shown in figure 9.17.

FIGURE 9.17 Formation of vulcanised rubber



When vulcanised rubber is stretched, the sulfur linkages stop the chains slipping past one another and the rubber returns to its original shape when the stretching force is removed. Vulcanised rubber is an example of a cross-linked polymer where the sulfur atoms link straight chains together. Rubber is used for tyres, carpet backing, tyre tubing and the soles of sports shoes.

TABLE 9.5 Some examples of polymer properties and related uses

Polymer	Properties	Use
High-density polyethene (HDPE)	Semi-rigid, chemically inert	Milk and juice bottles, water pipes, toys, wheelie bins
Low-density polyethene (LDPE)	Flexible, water repellent, chemically inert, electrical insulator	Plastic shopping bags, food wrap, electrical wire covering
Polyvinyl chloride	Flexible*, electrical insulator, water repellent	Garden hoses, raincoats, electrical wire covering, spouting, downpipes, non-food containers
Polypropene	Rigid, water repellent, strong, chemically inert	Plastic furniture, plastic buckets, food containers, car parts
Polystyrene; polystyrene (expanded)	Tough, rigid; lightweight, thermal insulator	Plastic cutlery, CD cases, laboratory ware; packaging, insulation, coffee cups
Polytetrafluoroethene (PTFE) (Teflon®)	Lowcoefficient of friction, chemically inert	Non-stick cookware, low-friction bearings, plumbing tape
Polyethene terephthalate (PET/PETE)#	Tough, flexible, clear, lightweight, shatterproof, inert	Soft drink bottles, water bottles, moulded packaging, carpet, clothing

* Varies and depends on use of plasticisers; can range from stiff to flexible.

\# PET is a condensation polymer. All others shown are addition polymers.

9.4 EXERCISE

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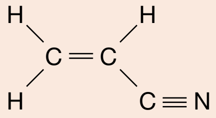
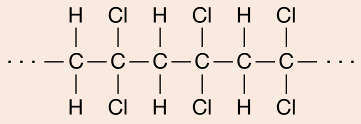
1. The structure of acrylonitrile is shown.

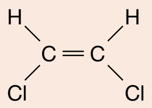


Draw a section of the polymer that results when acrylonitrile polymerises to form polyacrylonitrile.

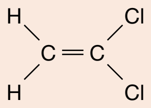
2. The structure of polyvinylidiene chloride is shown. The structure of two possible monomers from which it might be made is also shown.



Monomer I



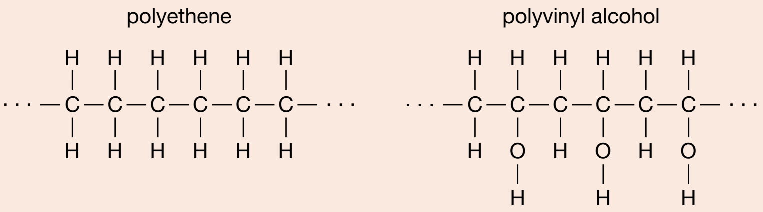
Monomer II



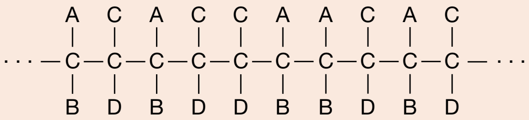
   (a) Which is the correct monomer and why?
   (b) Draw the polymer produced if the other monomer is polymerised.

3. Explain how chain length and the degree of branching affect the properties of a linear polymer.

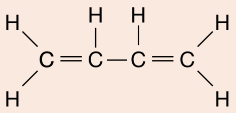
4. Explain why plasticisers can make a polymer softer and more flexible.
5. With reference to the structures of polyethene and polyvinyl alcohol, shown, explain why one is water repellent and the other is water soluble.



6. Explain why polymers with regularly spaced side groups along their carbon backbones are more likely to show a higher degree of crystallinity than those where the side groups are more randomly orientated.
7. Polypropene, polystyrene and polytetrafluoroethene are all water repellent. Explain with reference to their structure and bonding.
8. An application requires a polymer that can soften under moderate heat but set again when it cools. When set, it needs to be rigid and water repellent.
   Discuss how you would design a polymer to meet these specifications.
9. The structure of a section of hypothetical polymer is as follows.



   Draw the structure of the monomer from which it is produced.
10. The first synthetic rubber was made from but-1,3-diene.



   This was polymerised to make poly(but-1,3-diene), also known as polybutadiene.
   Draw the structure of this polymer.
11. Modern cars contain many parts that are made of plastic; for example, bumper bars and front side panels. Give two advantages and one disadvantage of using polymers in this way.
12. Give one advantage for the use of polymers in place of each of the following materials.
    (a) Metal
    (b) Wood
    (c) Glass
    (d) Wool
    (e) Surface coating
13. PVC is the third most produced polymer after polyethene and polypropene. However, a much greater amount of polyethene and polypropene are recycled than is PVC. Give possible reasons for this discrepancy.

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9.4.3 Advantages and disadvantages of using polymers

Advantages of using polymers

A quick look around the room will show the versatility of polymers and how dependent our society is on them. Plastics contribute to transport, construction, entertainment, packaging, clothing and other everyday items such as furniture, toys, bags and pens. Using plastics has many advantages, but advantages such as durability are also a disadvantage when it comes to recyclability.

Advantages of plastics include:

- can be moulded into any shape
- don't corrode
- are cheaper raw materials than alternative materials
- are less dense so lighter
- are good insulators of heat (home insulation)
- are good insulators of electricity (wire covers)
- are water resistant
- some are recyclable
- most are chemically inert
- are strong
- are relatively cheap to produce
- are useful for surfaces for construction purposes
- reduce the need to use limited natural materials such as timber.

Disadvantages of using polymers

Disadvantages of plastics include:

- use up a non-renewable/finite resource
- take a long time to decompose
- take up space in landfill
- common source of pollution, especially in oceans
- produce toxic gases on combustion
- some cause health problems.

9.5 Review

9.5.1 Summary

Formation of polymers

- A polymer is a long chain of units called monomers that are linked together in a process called polymerisation.
- Natural polymers include wool, cotton and hair. Synthetic polymers are commonly called plastics or synthetic fibres.
- Polymers are usually made by either addition or condensation polymerisation. Addition polymers are formed when an addition reaction causes monomers containing carbon–carbon double bonds (alkenes) to link together.
- Ethene polymerises to form polyethene (also known as polyethylene).

Linear and cross-linked polymers

- Thermoplastic (or thermosoftening) materials soften when heated and can be remoulded. An early example is celluloid.
- Thermosetting materials do not melt when heated.
- Thermoplastic materials are made from linear polymer chains.
- Thermosetting materials are made from chains that have been linked together with strong covalent bonds, and are called cross-linked polymers.

- Thermosetting polymers are difficult to recycle. A number of thermoplastic polymers, however, are recycled on a large scale.
- The presence of cross-linking, or bonding between the chains, affects the elasticity and rigidity of the polymer. A small amount of cross-linking produces an elastomer, which is relatively elastic, and a large amount produces a rigid polymer, because atoms are strongly bonded in all three dimensions.
- Many plastics can be recycled. Discarded plastics can be sorted according to standard recycling codes. The different groups of plastics can then be shredded, washed and melted down to produce chips to be remoulded.

Polymer selection

- Ethene can be modified by substituting different functional groups for hydrogen atoms to produce other monomers such as are styrene, vinyl chloride, tetrafluorene and propene. These can be polymerised to produce polymers with different properties and many uses.
- The structure of a linear polymer determines its properties, as follows:
  - *Extent of branching:* Low-density polyethene (LDPE) and high-density polyethene (HDPE) are two important polymers produced from ethene. The two substances have very different properties due to the different structures of their polymer chains.
    - LDPE chains support a large number of long side branches, which keep the chains apart. As the attraction by dispersion forces is weaker when the chains are further apart, LDPE is a soft, flexible and translucent material with a waxy surface that can be used for packaging, flexible industrial piping, and wire and cable insulation sheathing.
    - HDPE is produced using lower pressure and temperatures than LDPE and the resulting chains have very few and short side branches, which are held closer together by dispersion forces, making the material less flexible, stronger and opaque. HDPE is used to make tough plastic bags, kitchen containers, car parts, rigid pipe and fibres.
  - *Cross-linking of polymer chains:* The presence of cross-linking, or bonding between the chains, affects the elasticity and rigidity of the polymer. A small amount of cross-linking produces an elastomer, which is relatively elastic, and a large amount produces a rigid polymer, because atoms are strongly bonded in all three dimensions
  - *Degree of crystallinity:* More regions of regular arrangement increase strength.
  - *Length of the polymer chains:* The longer the chain, the stronger the polymer.
  - *Addition of plasticisers:* Plasticisers are small molecules in polymers that separate the polymer chains and make them more flexible.
  - *Additives:* These contribute to the appearance and stability of the polymer.
- The many ways that polymers may be made or modified means that a polymers may be made or chosen for a wide variety of applications.
- Using plastics has many advantages; however, some of the advantages, such as their durability, are also a disadvantage when it comes to recyclability.

Resources

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9.5.2 Key terms

addition polymerisation the process in which monomers with at least one double bond react together to form a polymer by addition reactions

addition reaction reaction in which one molecule bonds covalently with another molecule without losing any other atoms

condensation polymerisation polymerisation in which two monomers combine and a smaller molecule is eliminated
copolymer polymer formed from the polymerisation of two monomers
cross-linking the bonding between two polymers
difunctional refers to a molecule that contains two functional groups
elastomer polymer that can be stretched and return to its original shape
high-density polyethene (HDPE) thermoplastic polymer with very few side branches produced from the monomer ethene
isoprene volatile liquid hydrocarbon obtained from petroleum, the molecule of which forms the basic structural unit of natural and synthetic rubbers
linear polymer polymer that does not have side chains
low-density polyethene (LDPE) thermoplastic polymer with a large number of side branches produced from the monomer ethene
monomer molecule that links together to form a polymer
PET thermoplastic polymer produced from the monomer ethylene terephthalate; typically used as synthetic fibres (polyester) and packaging (PET); commonly recycled
plastic polymer that can be moulded when hot and retains its shape when cooled
plasticisers small molecules that are added to some polymers to improve their flexibility
polyester synthetic polymer in which the structural units are linked by ester bonds
polyethene polymer consisting of monomers of ethene
polymers molecule made up of thousands of units strung together into very long chains
polymerisation the formation of giant molecules by repeated monomers that are joined by covalent bonds
rubber an elastic polymer formed from the latex of tropical rubber trees and plants
synthetic fibre a polymer that is stretched to make a fibre instead of being moulded
thermoplastic describes polymers that soften on heating; also known as thermosoftening
thermosetting describes polymers that do not soften on heating and char if heated strongly
vulcanisation the process in which sulfur is added to rubber and heated to cause cross-linking of polymer chains, increasing the strength of the rubber

on Resources

Digital document Key terms glossary — Topic 9 (doc-30955)

9.5.3 Practical investigations

Experiment 9.1

Cross-linking an addition polymer to make slime

Aim: To investigate how the properties of a linear polymer may be altered by the introduction of weak cross-linking between its chains

Digital document: doc-30851

Teacher-led video: tlvd-0626

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Digital documents Practical investigation logbook (doc-30956)

9.5 Exercises

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9.5 Exercise 1: Multiple choice questions

1. Which one of the following is a natural polymer?
- **A.** Cellulose
- **B.** Polypropene
- **C.** Teflon
- **D.** Kevlar

2. Large molecules can be built up by the combination of smaller molecules. What are these smaller molecules called?
- **A.** Polymers
- **B.** Isomers
- **C.** Monomers
- **D.** Allotropes

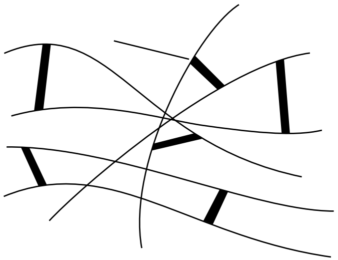
3. What structural characteristics are present in the monomers used to produce addition polymers?
- **A.** They must contain only carbon and hydrogen.
- **B.** They must all be the same.
- **C.** They must contain a carbon–carbon double bond.
- **D.** They must have a low molecular mass.

4. Which of the following best describes polymerisation?
- **A.** Small molecules decompose to form a new substance.
- **B.** Two or more chemicals react together.
- **C.** Small molecules react to form very long molecules.
- **D.** Small molecules react to form a thicker substance.

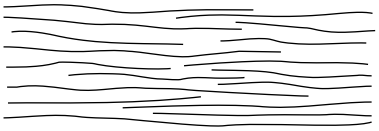
5. Which one of the following is a characteristic of thermoplastic?
- **A.** Cross-links between chains
- **B.** Can only be moulded during manufacture
- **C.** Chars when heated
- **D.** Has weak dispersion forces between chains

6. Which of the following diagrams depicts a thermosetting polymer?

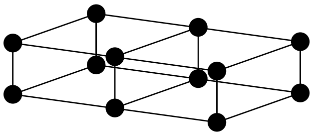
A.



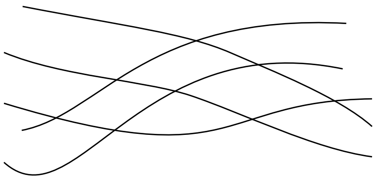
B.



C.



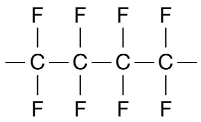
D.



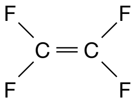
7. An elastomer is a linear polymer
- **A.** with a large number of cross-links between its chains.
- **B.** with a small number of cross-links between its chains.
- **C.** with no cross-links between its chains.
- **D.** without cross-links that has tangled chains.

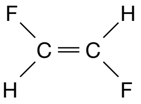
8. Why are thermoplastic polymers more suited to recycling than thermosetting polymers?
   - **A.** They are lighter and easier to collect in roadside collection services.
   - **B.** Most have short-term uses and quickly become waste.
   - **C.** They have cross-links between their polymer chains that are easily broken by heat.
   - **D.** Their properties mean that they can be re-melted and remoulded into new articles.

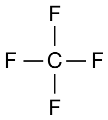
9. Teflon is made from the polymer polytetrafluoroethylene (PTFE). It has the following structure.

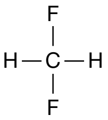


The monomer for this polymer is:

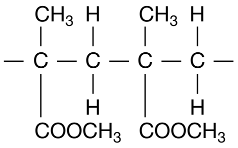
A.


B.


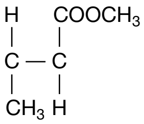
C.


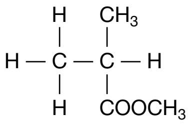
D.


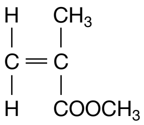
10. Polymerisation of methyl methacrylate produces ‘perspex’. The structure of perspex is as follows.

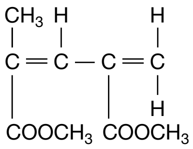


What is the structure of methyl methacrylate?

A.


B.


C.


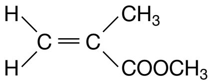
D.


11. When the chain length of a polymer increases
   - **A.** the melting temperature is lowered.
   - **B.** the substance becomes more viscous.
   - **C.** solubility in non-polar solvents increases.
   - **D.** density decreases.

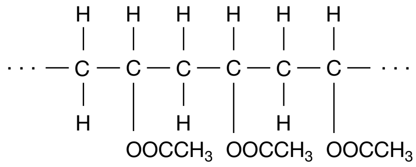
12. What is the difference in strength between low-density polyethene (LDPE) and high-density polyethene (HDPE) mostly due to?
   - **A.** The degree of cross-linking
   - **B.** The degree of side branching
   - **C.** A difference between the two in relative molecular mass
   - **D.** The orientation of hydrogen atoms along the carbon backbone

9.5 Exercise 2: Short answer questions

1. What are monomers and polymers?
2. List three advantages and three disadvantages of using polymers.
3. Which of the following molecules could possibly act as monomers in a polymerisation reaction? In each case, explain why or why not.
   a. C_4H_{10} b. C_3H_6 c. C_2H_6 d. C_4H_6 e. C_2H_4 f. C_2H_2 g. C_3H_8
4. Distinguish between thermosetting and thermoplastic polymers and give three examples of each.
5. In what ways is an elastomer similar to
   a. a thermoplastic substance?
   b. a thermosetting substance?
6. Briefly explain why a thermosetting polymer is more rigid than a thermoplastic one.
7. LDPE is soft and flexible whereas HDPE is stronger and more rigid. Explain.
8. Explain why the recycling of plastics is desirable.
9. Describe why the coding system used for plastics is beneficial.
10. Perspex is an addition polymer that has the appearance of glass. It is made from the monomer.



   a. Draw part of the polymer of perspex.
   b. In what situations would the use of perspex be superior to the use of glass?
   c. In what situations would the use of glass be superior to the use of perspex?
11. A polymer used in the manufacture of wood glue is shown.



   a. Draw the structure of the monomer from which the polymer is derived.
   b. Describe the forces that would hold these chains together.
   c. Show the structure in the condensed (bracketed) form.
12. Transparent film for packaging can be made from LDPE but would not be made from HDPE, which is more opaque. Explain this difference in terms of the molecular structure of the two forms of polyethene.
13. A feature of polymer materials is that they can be designed for a particular purpose. Accordingly, research into new polymers often focuses on means by which a polymer's properties may be influenced. Some of the factors that influence these properties are
      i. choice of monomer
      ii. length of polymer chains
      iii. degree of branching present in the chains
      iv. degree of crystallinity
      v. presence of plasticers.

   What factors would be important in a polymer designed for use in the following situations?
   a. A dissolvable wrapping for a dishwasher tablet
   b. A child's ride-on toy
   c. Plastic bags for fruit and vegetables at the supermarket
   d. Garden hose

14. Polymers are sometimes called giant molecules because they contain so many atoms. As polymer chains are twisted together, they form fibres. Diamond and graphite are also classed as giant molecules. A list of the types of bonds studied in this course follows. These bonds may exist either within a chain or molecule (intramolecular bonds) or between chains or molecules (intermolecular bonds). Complete the table that follows, using the bond types.

Bond types: covalent, dispersion forces, hydrogen bonding, disulfide

	Bond type	
Polymer	**Intramolecular**	**Intermolecular**
Diamond		
Graphite		
Polyethene		
Vulcanised rubber		

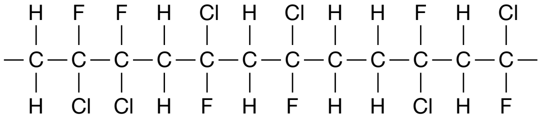
9.5 Exercise 3: Exam practice questions

Question 1 (5 marks)

a. Draw the full structural formula of the unsaturated molecule represented by the semi-structural formula CHClCHF. **1 mark**

b. Draw a section of the polymer produced when the monomer from part a. undergoes polymerisation. **2 marks**

c. A section of an addition polymer is as follows.



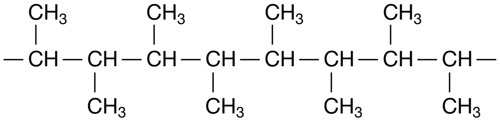
$$\begin{array}{ccccccccccccccccccccccccc} & H & & F & & F & & H & & \mathrm{Cl} & & H & & \mathrm{Cl} & & H & & H & & F & & H & & \mathrm{Cl} & \\ & | & & | & & | & & | & & | & & | & & | & & | & & | & & | & & | & & | & \\ - & C & - & C & - & C & - & C & - & C & - & C & - & C & - & C & - & C & - & C & - & C & - & C & - \\ & | & & | & & | & & | & & | & & | & & | & & | & & | & & | & & | & & | & \\ & H & & \mathrm{Cl} & & \mathrm{Cl} & & H & & F & & H & & F & & H & & H & & \mathrm{Cl} & & H & & F & \end{array}$$

Draw the structural formula of the monomer that it is produced from. **2 marks**

Question 2 (5 marks)

a. Draw the structural formulas of the following molecules.
   i. But-2-ene
   ii. But-1-ene **2 marks**

b. Examine the section of polymer that is shown.



$$\begin{array}{ccccccccccccccccc} & CH_3 & & & & CH_3 & & & & CH_3 & & & & CH_3 & & & \\ & | & & & & | & & & & | & & & & | & & & \\ - & CH & - & CH & - & CH & - & CH & - & CH & - & CH & - & CH & - & CH & - \\ & & & | & & & & | & & & & | & & & & | & \\ & & & CH_3 & & & & CH_3 & & & & CH_3 & & & & CH_3 & \end{array}$$

From which of the substances in part a. could this polymer have been made? **1 mark**

c. Draw a section of the polymer chain that could be made from the other molecule (i.e. the one *not* used for part b.). **2 marks**

Question 3 (8 marks)

High density polyethene (HDPE) and polyvinyl chloride (PVC) are both formed with chains that have only a small amount of branching. They show a number of similarities, but also a number of differences.

a. Draw the structural formulas of the monomers that each is made from. **2 marks**

b. For samples of each in which the length of the polymer chains is about the same, what would be some of the factors that affect the difference in strength and hardness of these two polymers? Explain. **2 marks**

c. One of these polymers has plasticisers added to it to increase its versatility. Explain how the addition of plasticisers affects the properties of this polymer. **2 marks**

d. Polyethene can also be produced in a different form called low density polyethene (LDPE). Describe the main difference in molecular structure between LDPE and HDPE and the effect that this has on the polymer's properties. **2 marks**

Question 4 (6 marks)

Polyethyne (polyacetylene) is a polymer that is made from the monomer ethyne. In 1977 it was discovered that it could be modified to become an electrically conductive polymer. This led to an explosion in research and possible applications. In 2000, the Nobel prize in chemistry was awarded for work in this field.

a. Draw the structural formula for a molecule of ethyne and use this to explain why it can be polymerised. **2 marks**

b. Draw a section of polyethyne that shows at least three monomer units. **2 marks**

c. Show your answer from part **b**. using condensed (bracketed) notation. **1 mark**

d. How would you expect the reactivity of polyethyne to compare to polyethene? **1 mark**

9.6 Exercise 2: studyON Topic Test online only

Fully worked solutions and sample responses are available in your digital formats.

teachon

Test maker

Create unique tests and exams from our extensive range of questions, including practice exam questions.
Access the Assignments section in learnON to begin creating and assigning assessments to students.

AREA OF STUDY 2 How can the versatility of non-metals be explained?

OUTCOME 2

Investigate and explain the properties of carbon lattices and molecular substances with reference to their structures and bonding, use systematic nomenclature to name organic compounds, and explain how polymers can be designed for a purpose.

PRACTICE EXAMINATION

STRUCTURE OF PRACTICE EXAMINATION		
Section	**Number of questions**	**Number of marks**
A	20	20
B	7	30
	Total	**50**

Duration: 50 minutes

Information:

- This examination consists of two parts, you must answer all question sections.
- Pens, pencils, highlighters, erasers, rulers and a scientific calculator are permitted.
- You may use the VCE Chemistry Data Book for this task.

on Resources

Weblink VCE Chemistry Data Book

SECTION A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

1. Which of the following best describes a molecular compound?

A. Ions held together by the electrostatic attraction between positive and negative ions

B. Atoms containing elements that differ greatly in electronegativity

C. Atoms in which the elements share electrons

D. A lattice of non-metallic atoms held together by shared electrons

2. Both bonding and non-bonding electrons are present in the valence structures of covalent compounds. How many non-bonding electrons are present in the structure of methane, CH_4?

A. 0

B. 2

C. 4

D. 6

3. The following table shows the electronegativity values of five different elements.

Element	***W***	***X***	***Y***	***Z***
Electronegativity	0.8	2.2	3.5	4.0

Based on the information in the table, what would be expected?

A. A bond formed between *Y* and *Z* would be strongly polar due to the high electronegativity values.
B. The compound formed from elements *Z* and *W* would be ionic in nature.
C. A bond formed between *X* and *Z* would be polar with the more positive end found at *Z*.
D. The bond formed between *W* and *X* would be more polar than the one formed between *X* and *Z*.

4. The shape of a molecule can have a marked effect on its polarity. Which of the following molecules contains polar bonds but is an overall non-polar molecule?

A. The linear molecule HF.
B. The tetrahedral molecule $SiCl_4$
C. The angular molecule H_2S
D. The pyramidal molecule PBr_3

5. Which of the following pairs of molecules consist only of polar molecules?

A. H_2O and CH_4
B. NH_3 and N_2
C. HCl and NaOH
D. HBr and OBr_2

6. Covalent solids can be classified as covalent network (3D) lattices, covalent layer (2D) lattices, or molecular lattices. Which of the following has the substances correctly assigned?

	Covalent network lattice	**Covalent layer lattice**	**Molecular lattice**
A	Diamond, C	Graphite, C	Carbon dioxide, CO_2
B	Graphite, C	Ice, H_2O	Silicon carbide, SiC
C	Silicon dioxide, SiO_2	Diamond, C	Carbon dioxide, CO_2
D	Carbon dioxide, CO_2	Graphite, C	Ice, H_2O

A. Option A
B. Option B
C. Option C
D. Option D

7. A student researches the properties of three substances: magnesium sulfate, $MgSO_4$; sulfur dioxide, SO_2; and magnesium metal, Mg. Which of the following statements regarding these substances is correct?

A. Both Mg and $MgSO_4$ have high melting points but only $MgSO_4$ conducts electricity in its molten form.
B. Neither $MgSO_4$ nor SO_2 would conduct electricity in the solid state.
C. Both SO_2 and $MgSO_4$ have high melting points and conduct electricity in their liquid state.
D. Both SO_2 and Mg can conduct electricity in either the liquid or solid state.

8. Fluoromethane, CH_3F, can be solidified if cooled to −138°C. Apart from dispersion forces, what bond types would be present in a sample of frozen fluoromethane?

A. Both covalent bonds and hydrogen bonding
B. Hydrogen bonding only
C. Covalent bonding and dipole–dipole interactions
D. Covalent bonding only

9. Hydrocarbons can be classified as alkanes, alkenes or alkynes. Which of the following statements is true?

A. Due to the presence of multiple bonds, alkenes and alkynes tend to be more polar than alkanes.
B. Alkenes have a larger proportion of hydrogen than alkanes.
C. Alkynes are the only compounds that have just triple bonds between each carbon.
D. Alkanes, alkenes and alkynes have only dispersion forces for their intermolecular attraction.

10. Which of the following statements about the organic molecule 3,3-dimethylpent-2-ene are true?

I. It is an unsaturated molecule.
II. It has a total of 7 carbon atoms.
III. It has a different molecular formula from 3-ethylpent-2-ene.

A. Only statements (I) and (II) are true.
B. Only statements (I) and (III) are true.
C. Only statement (III) is true.
D. Only statement (I) is true

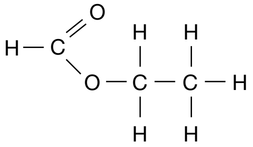
11. What is the number of hydrogen atoms present in one molecule of propanoic acid?

A. 4
B. 6
C. 8
D. 10

12. Which of the following is expected to have the *lowest* boiling point?

A. CH_3OH
B. CH_3CH_3
C. $CH_3CH_2CH_3$
D. CH_3NH_2

13. The ester shown here was formed by a condensation reaction between which reactants?



A. Ethanol and methanol
B. Methanol and ethanoic acid
C. Ethanol and methanoic acid
D. Methanoic acid and ethanoic acid

14. Which one of the following molecules *cannot* undergo an addition reaction?

A. $CH_3CH_2CH_3$
B. Propene
C. $CH_3CHCHCH_3$
D. But-1-ene

15. Which of the following describes a substitution reaction?

A. Ethene reacts with oxygen to produce carbon dioxide and water.
B. Ethane reacts with oxygen to produce carbon dioxide and water
C. Ethene reacts with HCl(g) in the presence of a catalyst to produce chloroethane.
D. Ethane reacts with chlorine gas in the presence of UV light to produce chloroethane and HCl.

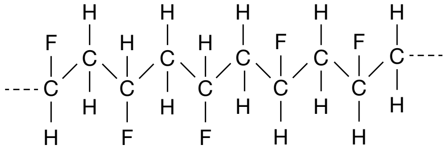
16. What is the correct systematic (IUPAC) name of $CH_3CH_2CH(CH_3)CH_2CH_2Cl$?

A. 1-chloro-3-methylpentane
B. 3-methyl-6-pentachloride
C. 1-chloro-2-methylpentane
D. 1-chloro-4-ethylbutane

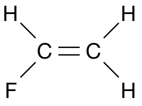
17. What is the number of different structural isomers that have the formula C_5H_{12}?

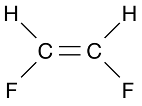
A. 5
B. 4
C. 3
D. 2

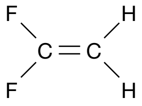
18. Consider the following polymer segment.

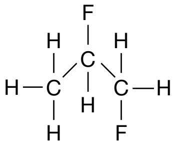


Which monomer would be used for producing this polymer?


Molecule A


Molecule B


Molecule C


Molecule D

A. Molecule A
B. Molecule B
C. Molecule C
D. Molecule D

19. What is the process of fractional distillation used to do?

A. Separate crude oil into its individual hydrocarbon compounds
B. Break down longer hydrocarbon chains into smaller, more useful, molecules
C. Remove hydrogen atoms from alkanes to make unsaturated molecules
D. Separate the hydrocarbon mixture into groups of similar boiling points

20. Which of the following statements about thermosetting and thermoplastic polymers is correctly classified?

	Thermosetting	**Thermoplastic**
A	Contains extensive cross links between polymer chains	Forms rigid materials that cannot be softened on melting
B	Chars when moderately heated	Only dispersion forces operate between polymer chains
C	Only dispersion forces operate between polymer chains	Polymer chains are kept separated by the addition of plasticisers
D	Forms rigid materials that cannot be softened on melting	Unable to be recycled

A. Option A
B. Option B
C. Option C
D. Option D

SECTION B

Question 1 (8 marks)

1. The following table lists two molecules.
Complete the following for each molecule.

a. Draw the Lewis (electron dot) formula. **1 mark for each**

b. Draw the structure giving its shape according to VSEPR theory and including any non-bonding electrons. **1 mark for each**

c. Classify each as polar or non-polar and mark in the $\delta-$ and $\delta+$ regions on opposite ends of any of the polar structures. Leave blank if non-polar. **1 mark for each**

d. Classify the shape of the molecule. **1 mark for each**

Molecule	Nitrogen tri-iodide, NI_3	Sulfur dichloride, SCl_2
Lewis (electron dot) formula		
Structure of molecule showing shape and non-bonding electrons		
Polar or non-polar		
Shape		

Question 2 (2 marks)

2. From the following list, choose the process that will break the types of bond given in the table.
Process:

- Burning methane gas
- Boiling decane
- Evaporating liquid nitrogen
- Melting tin
- Evaporating ethanol
- Melting magnesium sulfate

Bond type	Process
Covalent	
Hydrogen bonding	

Question 3 (4 marks)

3. Explain why the properties of these substances differ by referring to their bonding.

a. Diamond and graphite are allotropes of carbon. Diamond is the hardest naturally occurring substances but graphite is soft and tends to form flakes. **1 mark**

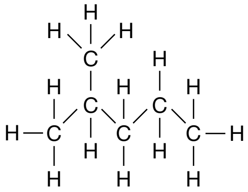
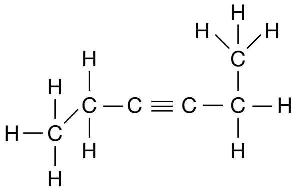
b. Butane has a lower boiling point than pentane. **1 mark**

c. Hexane and butanoic acid have a similar molar mass, but hexane has a boiling point of 68 °C while butanoic acid has a boiling point of 163 °C. **1 mark**

d. Both carbon dioxide, CO_2, and water are molecules containing polar bonds and three atoms, yet CO_2 is non-polar while water is polar. **1 mark**

Question 4 (6 marks)

4. a. Complete the following table, providing either the systematic (IUPAC) name or the structural formula as required for the compounds. **4 marks**

Structural formula	Systematic name
	
	But-1-ene
	
	Methanoic acid

b. Write the semi-structural (condensed) formula for ethylbutanoate. **1 mark**

c. Write the molecular formula for octane. **1 mark**

Question 5 (3 marks)

5. Write balanced equations for the following reactions. You may use structural formulae, condensed formulae or molecular formulae in your equations. Symbols of state are not required.

a. The reaction of propene with chlorine gas. **1 mark**

b. The reaction between ethane gas and chlorine gas under UV light. **1 mark**

c. The catalytic cracking of dodecane $C_{12}H_{26}$, to form a pair of hydrocarbon products **1 mark**

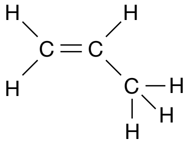
Question 6 (4 marks)

6. Draw and systematically name two *structural isomers* for a pair of alcohols that have the molecular formula C_3H_8O.

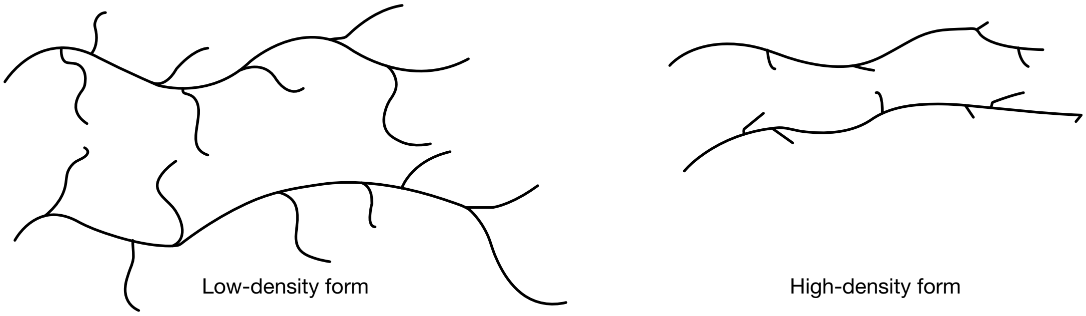
Question 7 (3 marks)

7. a. Draw a segment of polymer (at least three units) formed by addition polymerisation of the monomer shown.

1 mark



b. Depending on the reaction conditions, this polymer may be formed in a low-density form or a high-density form, as illustrated in the following diagram.



Apart from their different densities, describe two other ways that their physical properties would differ, making reference to the bonding involved.

2 marks

ASSESSMENT TASK — CREATING A BLOG

Most people receive their news about science through the popular media, rather than through academic research papers. One of the most important scientific discussions in our society today is how we choose to use our resources, including non-renewable resources such as crude oil. In this task you will create a blog on an issue related to the use of crude oil in our society.

- Part A is the creation of the blog.
- Part B is the final written response in response to a prompt. It will be written on lined paper provided to students. Students are permitted their blog notes, pens and pencils during this time.

Total time Part A: one full week to complete the preliminary investigation and blog posts.
Total time Part B: 50 minutes
Total marks: 22 marks

ISSUES AROUND CRUDE OIL

Introduction

Carbon dioxide is a greenhouse gas that traps the sun's rays, allowing fewer of them to exit the atmosphere. While greenhouse gases insulate the earth, excessive amounts of carbon dioxide in the atmosphere have resulted in the global temperature increasing rapidly over time. This is the phenomenon we know as global warming. Carbon dioxide emissions have been increasing rapidly since the Industrial Revolution due to the interplay between significant population increases and our dependence on energy and transport. Although many other gases contribute to global warming and the greenhouse effect, it is the sheer volume of carbon dioxide that is a major concern.

The burning of fossil fuels, including coal and the combustion of crude oil in engines, results in the release of carbon dioxide. Fossil fuels formed millions of years ago, capturing carbon to form carbon–carbon bonds as long-chained hydrocarbons. The combustion of hydrocarbons releases energy, water and carbon dioxide. For example, octane, a fuel found in crude oil and a component of petrol, will undergo complete combustion. The equation for this reaction is as follows:

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

For every two octane molecules, 16 carbon dioxide molecules are released into the atmosphere, contributing to the greenhouse effect.

We can use chemistry to understand our dependence on crude oil and use of carbon-based compounds to meet our energy needs.

Task

This task is broken into parts A and B. The final assessment will be in part B.

Part A: Over one week, write a minimum of five posts for a blog considering the issue: 'The origin of crude oil and its use as a source of hydrocarbon raw materials.' At least one post must be an individual post and you must comment on posts from other students with questions, reflections and comments about the statements made. The stimulus material provided below is not exclusive, and you may do additional research as long as you cite your source(s). Feedback will be provided on responses and comments for part A of the assessment task before proceeding to part B.

Part B: A timed, written response to explain the origin and use of crude oil, and intermolecular and intramolecular bonding. You will also discuss the advantages and disadvantages of the different efforts being made to combat the issues surrounding crude oil and our dependence on this as a fuel source.

The following prompting questions and links are provided to help you complete your preliminary research.

Link	Prompting questions
The Conversation 1	What are the key concerns outlined in this article? What are the factors that contribute to the length of time for something to biodegrade? How are biodegradable plastic bags made? What are the suggestions made in this article? What is the concern regarding materials released from degradable plastics?
Science Alert	What is polypropylene? What is a hydrocarbon? What is hydrothermal liquefaction? What is the end product of the process? What might be some concerns with this process?
The Scientist	What are first-, second- and third-generation fuels? How do we optimise the oil production of the algal pools? Why is it important that the oils produced are chemically similar to petroleum-based fuels? To make serious headway, how much fuel will need to be produced a day? What is the limitation with this? Why would it be better to directly consume carbon-biodegrabased compounds rather than take carbon dioxide from the air? What is the aim with genetic engineering of algal species? How do solazyme make biodiesel?
The Conversation 2	What is the issue with carbon capture? What are the two products of a complete combustion reaction that are released in the exhaust? Which product interferes with carbon capture? What industries are identified as releasing large amounts of concentrated carbon dioxide without water vapour? Why should we consider CO_2 utilisation technologies?
Truth Out	What does it mean to decarbonise our economies? Which countries are increasing in their emissions? What do you think is contributing to this increase? How has the use of renewable energy changed over the years? What does 'renewable energy' include? How does renewable energy not contribute to the increase release of carbon dioxide?

The Conversation 1:
http://theconversation.com/why-compostable-plastics-may-be-no-better-for-the-environment-100016
Science Alert:
https://www.sciencealert.com/this-is-how-we-might-turn-millions-of-tonnes-of-plastic-back-into-oil
The Scientist: https://www.the-scientist.com/uncategorized/future-oil-44421
The Conversation 2: https://theconversation.com/co-shortage-why-cant-we-just-pull-carbon-dioxide-out-of-the-air-99255
Truth Out: https://truthout.org/articles/carbon-emissions-will-reach-a-record-high-of-37-billion-tons-in-2018/

on Resources

Digital document U1AOS2 School-assessed coursework (doc-30963)

AREA OF STUDY 3
RESEARCH INVESTIGATION

10 Research investigation

10.1 Overview

Numerous **videos** and **interactivities** are available just where you need them, at the point of learning, in your digital formats, learnON and eBookPLUS at www.jacplus.com.au.

10.1.1 Introduction

Strong research skills allow us to better understand patterns and relationships in chemistry. Scientific research in chemistry is based on our fundamental understanding of chemical structure. Research discoveries deepen this understanding and allow the creation of new materials that assist us in all aspects of our life. However, for this research to be of benefit it must be effectively communicated. As part of your research investigation, you will apply creative and critical thinking skills to investigate discoveries that have allowed for the development and modification of useful materials and chemicals. You will build skills in effective scientific communication and inquiry, and gain an understanding of the nature of the evidence that we use to draw conclusions.

FIGURE 10.1 Effective research involves careful analysis of information and evidence

10.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- the 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
- the chemical concepts specific to the investigation: definitions of key terms; use of appropriate chemical terminology, conventions, units and representations
- the use of data representations, models and theories in organising and explaining observed phenomena and chemical concepts, and their limitations
- the nature of evidence and information: distinction between weak and strong evidence, and scientific and non-scientific ideas; and validity, reliability and authority of data, including sources of possible errors or bias
- the influence of social, economic, environmental and ethical factors relevant to the selected chemical investigation.

KEY SCIENCE SKILLS

- Plan and undertake investigations
- Analyse and evaluate data, methods and scientific models
- Draw evidence-based conclusions
- Communicate and explain scientific ideas

Source: VCE Chemistry Study Design (2016–2021) extracts © VCAA; reproduced by permission.

on Resources

Digital documents Key science skills (doc-30903)
Key terms glossary — Topic 10 (doc-30958)

10.2 What is a research investigation?

KEY CONCEPTS

- The 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
- The use of data representations, models and theories in organising and explaining observed phenomena and chemical concepts, and their limitations

KEY SCIENCE SKILL

- Plan and undertake investigations

10.2.1 Why do we conduct research investigations?

A research investigation is an exploration of a scientific question or questions and presentation of findings. In Outcome 3 you have the opportunity to find out more about an aspect of the chemistry of materials that you have studied in Unit 1. This will involve researching, evaluating and interpreting information, and presenting it in a cohesive, accurate and detailed report.

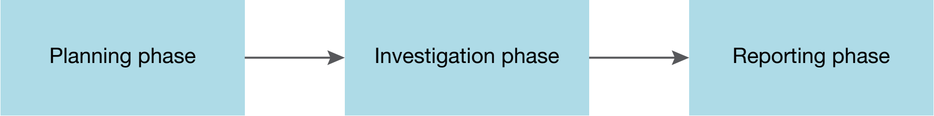
Researching allows us to:

- gain a comprehensive understanding to a topic and research question
- discover current and prior investigations and findings regarding a topic
- make informed decisions
- investigate patterns and relationships
- understand differences between strong evidence that has been peer-reviewed and weaker evidence that shows bias
- understand ways to conduct and apply our research to practical approaches and the development of new scientific understanding.

10.2.2 Planning and presenting a research investigation

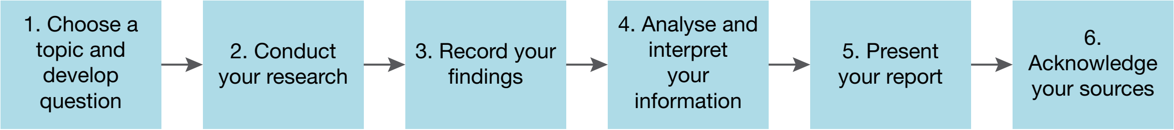
When planning and conducting your research investigation, it is helpful to follow a structure in order to make sure that your research is valuable, appropriate to the questions, and allows you to apply critical thinking, inquiry and communication skills. You will use both your prior knowledge from topics you have covered throughout the year, as well as gather new ideas and information to help draw conclusions. Your investigation will be made up of three main phases, as outlined in figure 10.2.

FIGURE 10.2 Phases of research investigations



These three phases can be further subdivided into six steps that will help you plan your investigation and your report. It is important to note that you may move back and forward between these steps as your research investigation develops.

FIGURE 10.3 Planning and presenting your research investigation



Throughout the planning and presenting of a research investigation, it is important to consider:

- the target audience results are being communicated to
- the format of communication
- if work is independent or in groups
- if research is required outside of class time
- time allocated in class
- how work will be checked and authenticated.

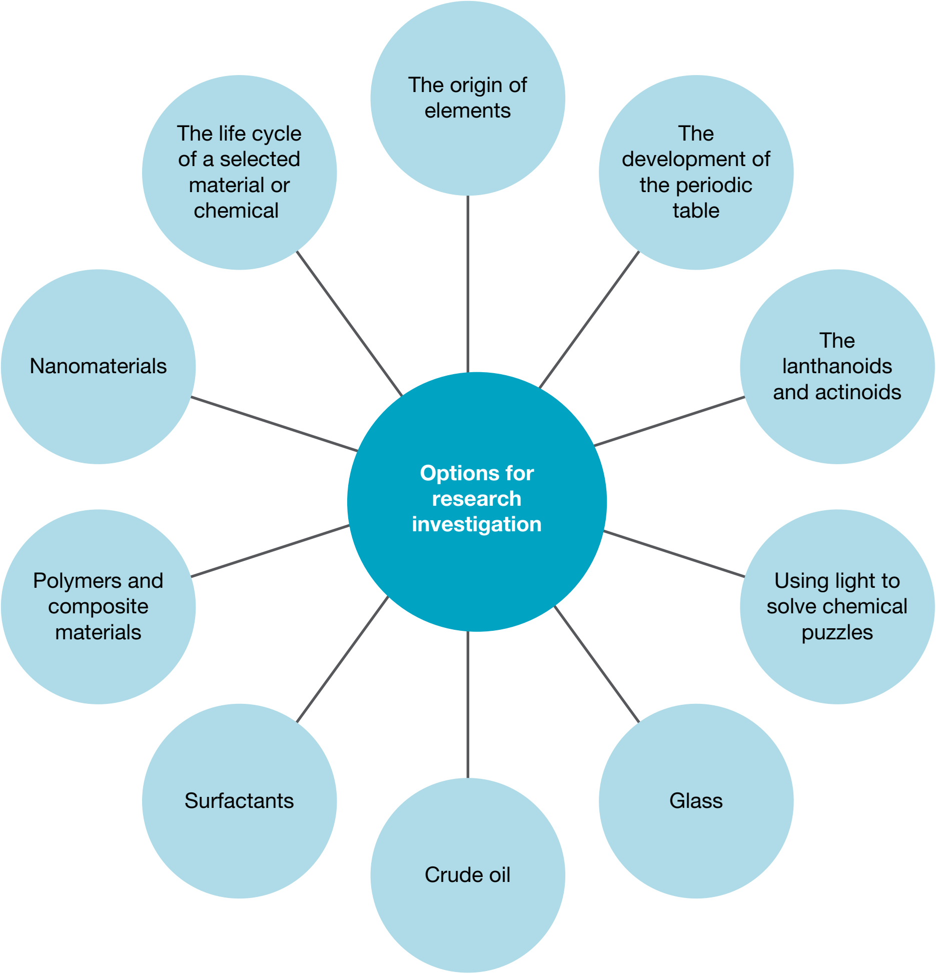
10.2.3 Choosing a topic and developing a question

Investigate a topic that interests you. You may wish to consider one of the questions provided in the options in the VCAA Chemistry Study design. Alternatively, after a discussion with your teacher, you may prefer to respond to questions of your own choosing or have a combination of both. Typical questions begin with how, why, what, when, who or where. Record the question(s) that you are going to investigate. These questions may change as you start researching.

For Unit 1, Outcome 3, your teacher may choose to conduct the task in different ways, so it is important to ask what is expected. Some of these different methods include the following:

- The teacher provides the list from the study design and students select their question and submit a research proposal.
- Groups of students investigate questions and each member of the group contributes to the assessment task in a multimedia format.
- The teacher selects questions from the study design focusing on case-study themed questions and provides case studies for students to examine.
- The teacher selects questions that have an experimental theme and students need to provide a response to these.
- Students need to come up with their own questions, which are then approved by the teacher.

FIGURE 10.4 For Unit 1 Outcome 3, you can choose from ten topics for your research investigation.



10.2.4 Conducting your research

Once the key questions have been decided, it is time to commence your research using the abundance of available sources. Later in this topic (in subtopic 10.4), we will investigate the nature of evidence and information and the type of evidence you should be using in your research investigation.

You should use multiple mechanisms for conducting your research, not just the internet. Some valuable places to conduct your research are:

- books
- newspapers
- the internet
- scientific journal articles
- scientific magazines
- podcasts
- videos
- interviews with experts in the field.

The data and research you conduct will involve the collection of a combination of **primary data** and **secondary data**. Primary data is data that is collected by a researcher from first-hand sources; examples include surveys, measurements or experiments. Secondary data is analysis of data from other people's investigations; examples include published papers or databases.

FIGURE 10.5 Finding appropriate information for your investigation takes time and research

Using a combination of primary and secondary data will provide more scope and detail for your research investigation, allowing you to focus and gain information about your chosen question.

Lots of valuable information is out there, and part of research is deciphering the right information for your investigation and identifying information that may be irrelevant, biased or misleading. Types of information and data will be explored further in subtopic 10.4.

10.2.5 Record your findings

Record your information in an organised way. This may be in a digital or a written format.

In your log book, you should carefully examine each piece of research. One way to do this might be to set up a table, as shown in table 10.1.

TABLE 10.1 Organising your research

Title of research	
Date accessed	
Summary of information (in your own words)	
Comment about the information (is it relevant?)	
Reference (for bibliography)	

Remember to note any gaps in the information or the presence of competing opinions. You can organise the information using highlighters to collate different groups of ideas or sections. Write a plan showing the flow of the report and where each piece of information will be presented.

10.2.6 Analyse and interpret your information

As part of your research, it is important to analyse and interpret the information you have — both information that is prior knowledge, and information that you have researched.

This process should involve you analysing and assessing data, exploring trends and patterns, making generalisations and drawing conclusions. Be sure to make connections between the various sources of your information and clearly link research back to the question you are investigating. Interpret your findings — what do they suggest about the topic you are exploring? Do they provide answers to your question of investigation?

Once you have recorded information from a few sources, read your notes carefully and write out information as if you were explaining it to another person, using your own words and sentence structure. You may also consider using a mind map to better understand the relationships between different pieces of evidence. It is important that you understand both the content and make links between different lines of evidence. This will help you come to a conclusion.

When drawing conclusions, you should summarise and tie together the findings of your report and reflect back on the introductory paragraph. This might provide the opportunity to suggest areas for further investigation and reflect on your learning, but it should not include any new information.

10.2.7 Present a report of your research investigation

Decide how you will present your report. The various ways of presenting a report include:

- written format
- poster
- slideshow
- video
- animation
- oral presentation
- information pamphlet
- community campaign.

Check the requirements from your teacher for this task, because they may have a set way they wish for you to present.

10.2.8 Acknowledge your information

It is important that you do not **plagiarise** your research material. Changing a few words or reorganising a few sentences does not make the work your own. If you are quoting another person, make this clear by using quotation marks, and remember to cite the author. For information on how to cite information, refer to topic 20 (section 20.11.8). It is important to make sure that when you reference your sources that you do so using a standard referencing format such as the Harvard style. This ensures that anyone reading your work can easily go back to the initial source.

Acknowledging tables and images

Tables and images can be very effective at summarising large amounts of data and providing added interest; however, they need to be relevant and referenced. Images and data may be covered by copyright. Data should always be accompanied by a source line to demonstrate where it was found. Copyright-free images may be identified by a Creative Commons icon or using 'usage rights' in an advanced Google search engine. Many websites also have Creative Commons usage rights such as Wikimedia Commons.

Many different icons in Creative Commons determine how you can use an image, as shown in Figure 10.6. Because you will be using your images in a non-commercial setting, it is often easier to get permissions, but it is important to note what is required when an image (or other work) is used.

Creative Commons requirements may include the following:

- *Attribution (BY):* Referencing the image with the copyright holder's identity is required.
- *No Derivative (ND):* Altering or remixing the image (or other work) is not permitted.
- *Non-commercial (NC):* The image (or other work) is only available for non-commercial purposes (not for financial gain).
- *Share-alike:* Works that are adaptations or derivatives of the work need to be under the same license as the original.

Some work is in the public domain and may be used freely, without any issues regarding attribution and derivatives. The image for the public domain is shown in figure 10.7. Some examples of websites that have images in the public domain include Pexels, Pixabay and Unsplash. While including attribution or crediting the owner of the images is not required, it is always good practice to acknowledge your sources.

FIGURE 10.6 Different icons involved in Creative Commons

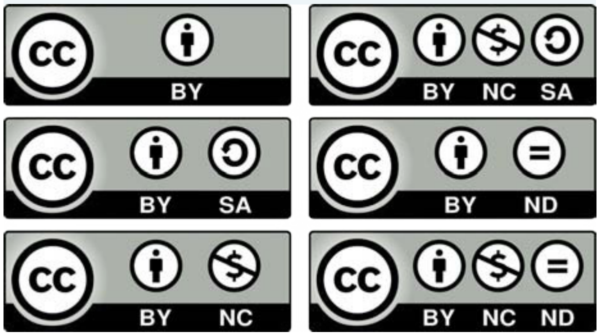


FIGURE 10.7 Images and media in the public domain can be used without copyright issues. Sometimes, only the icon with the crossed-out C is shown.

10.2 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Why is it important to conduct research investigations?
2. Explain why using a poster is a beneficial way to communicate the findings in a research investigation.

3. You find an image that you would love to include on your report. Under copyright, it has the letters 'BY' and 'NC'. What does this mean and would you be able to use the image in your report?
4. An investigation is being conducted on the use of nuclear energy instead of coal in Australia. Identify two examples of primary data and two examples of secondary data you may collect that is relevant to this investigation.

Fully worked solutions and sample responses are available in your digital formats.

10.3 Effective science communication

KEY CONCEPTS

- The 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
- The chemical concepts specific to the investigation: definitions of key terms; use of appropriate chemical terminology, conventions, units and representations
- 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, environmental and ethical factors relevant to the selected chemical investigation

KEY SCIENCE SKILLS

- Communicate and explain scientific ideas

10.3.1 Characteristics of effective science communication

Communicating effectively in science allows trends and patterns to be clearly seen, provides support for research and deepens conceptual understanding. Clear communication also provides context regarding the importance and implication of findings. Science communication is effective when others can read a report and gain clarity and understanding of the information and research undertaken.

Methods that assist in effective communication include the following:

- using correct and relevant terminology and explaining all scientific terms
- dividing the report into paragraphs and, where necessary, using subheadings to separate ideas
- taking care with punctuation, spelling and grammar
- having accurate chemical information, including using correct formulas
- making sure that all information has context and is important to your findings
- being concise and sticking to provided word limits — it is better to have a succinct 100-word summary than 500 words of irrelevant information where your key ideas are hard to distinguish
- offering clear explanations of chemical concepts, ideas and models found in your research and your report
- considering the audience you are writing for when establishing the ways you should communicate your findings.

Using chemical concepts specific to investigations

In a specialised field such as chemistry, communication can be improved by providing definitions of key terms, and using appropriate chemical terminology, conventions, units, representations, models and theories. This helps prevent misunderstandings in your communication and allows ideas to be presented clearly. Chemistry frequently uses capital letters, superscripts and subscripts to convey information and these should be applied with care. For example, CO is carbon monoxide, while Co is the element cobalt. Chemical formulas should always include the use of subscripts and superscripts where appropriate; for example, the formula for water should be written as H_2O, not H2O.

Each of these ideas will be further detailed in topic 20, but are outlined briefly in table 10.2.

TABLE 10.2 Chemical concepts specific to investigations

Chemical concept	Use	Limitations
Definition of key terms	Allows for a clear understanding of the meaning of terminology in the context of the investigation	Uncertainty exists about which specific words need to be defined.
Chemical terminology	Allows for individuals reading a report to contextualise the information to chemistry concepts	Similar words can be used to describe an identical concept or a phenomenon, which leads to confusion.
Conventions	Provides a set of standards followed in reporting and investigations (e.g. including an introduction, aim and conclusion)	Some conventions are unclear and so are changed, leading to confusion when different conventions are used.
Units	Provides an international set of units (SI) for measurements, allowing for easy comparisons between different research investigations	Units may need to be converted (i.e. from kilograms to grams); care should be taken to avoid a misinterpretation of data.
Representations	Shows certain chemistry aspects using symbols and models that make communication easier and clearer	This can lead to oversimplification of concepts by only representing them rather than showing all detail.
Models	Provides an explanatory framework to assist with understanding observed phenomena and abstract concepts	This oversimplifies concepts because they cannot be true representations of real-world conditions or show all details.
Theories	Outlines a well-supported explanation of phenomena, based on investigation, research and observations	Can rely on specific conditions being met which are not always realistic. Phenomena often need to be described with multiple theories.

10.3.2 The influence of social, economic, environmental and ethical factors

Many factors can have a great impact of the information obtained and the presentation of your research investigation, other than just scientific factors.

It is important to incorporate these factors in your investigation and understand the ways they affect findings and results.

Social factors

In your investigation, you need to have an understanding of factors that underpin the development, use and modification of chemicals and materials. One of the largest facets to consider is social factors. Social factors are those that shape the behaviours and thoughts of individuals. This involves the historical, cultural and political aspects of a population. It is vital to understand how new discoveries or changes to current discoveries affect both individuals and society as a whole.

When creating new chemicals and materials, researchers need to understand society in order to enhance marketability of the product. They need to make sure that the product will have long- and short-term benefits for individuals. This may be done through the use of surveys, focus groups and opinion polls.

Social factors also affect ideas such as education of individuals and their understanding and interpretation of information being presented to them. When researching, it is vital to gain an understanding of how society will be involved with the use of the product, because this in turn will affect the supply and demand of a new development.

An example of this may be in the creation of a new material that can be used as a contraceptive. This would affect societies in various ways — some individuals would be opposed to the use of contraception for cultural or religious reasons, or political interference may prevent the new material being imported into their country. It may be that the contraceptive was designed to better protect against HIV, which is prevalent in different levels in various populations, so may have very different demands based on societal requirements. Understanding social factors is vital to understanding the viability of the manufacturing and altering of chemicals or materials, because the behaviours and thoughts of society can be a huge barrier.

Economic factors

Economic factors are a large part of research. These factors relate to the cost of not only conducting investigations but also acquiring talent, identifying, applying and developing chemicals, and analysing scientific information to draw conclusions.

Research is very reliant on economic factors, in the form of grants to conduct investigations. It is important to be aware of the economic situation of where you are getting your information from, because research may be incomplete or rushed due to the loss of funding.

It is also important to consider **bias** in research due to financial incentives and funding from external companies. Sometimes, data is skewed or misrepresented due to financial reasons — it may be that research is being funded by a source that wants data represented in a way that is favourable for their intentions. This greatly affects the validity of data collected and research conducted. When looking at research, it is important to check that no conflict of interest exists that may lead to bias.

Another example of economic factors in chemistry research is the access to different research based on cost (some academic journals are quite expensive to download, while others are free). This can affect the availability of information to different socio-economic groups.

Finally, in investigating your topic, it is vital to understand the economic viability of the development of new materials and chemicals — will they provide significant benefit to others to an extent that they are profitable?

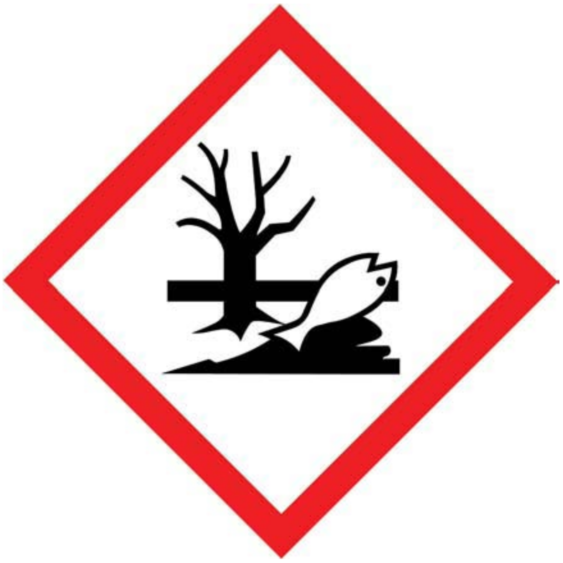
Environmental factors

Considering environmental factors is important in many areas of your chemistry representation. When synthesising and developing new chemicals and materials, the impact of the environment must be considered. This is evident in our search for new sources of fuel that are renewable; and in creating new materials that are biodegradable and do not build up in the environment and cause harm.

Symbols are often used on chemical products to alert the user to potential hazards. For example, the symbol seen in figure 10.8 means that the chemical is harmful to aquatic life. This is particularly important with the disposal of chemicals and how they can cause harm to the environment.

An example of a topic that focuses on environmental factors is in the creation of biodegradable surfactants. By creating biodegradable resources, the environment is less likely to be harmed during the manufacture, use and disposal of new materials and chemicals.

FIGURE 10.8 Symbol showing a chemical is hazardous to the aquatic environment



Sometimes, environmental issues are unintentional, in which chemicals or materials designed for beneficial purpose are found to have drastic effects in the long-term because they are retained in the environment. This is important to consider in investigations, because often long-term effects are difficult to study and so can be neglected.

Ethical factors

When conducting research, ethics are vital to consider; that is, the morals around what actions are correct or incorrect. In the development of new materials and chemicals, and the adaptations of current materials and chemicals, it is important to take into consideration ethical implications of both the results and the method of the research being conducted.

This includes the process of testing and checking that informed consent and confidentiality were provided to those involved in the study. It is important that the risks involved are made exceptionally clear to individuals taking part in studies.

FIGURE 10.9 The impact of chemicals on the environment must be considered at all stages, from manufacture to disposal.

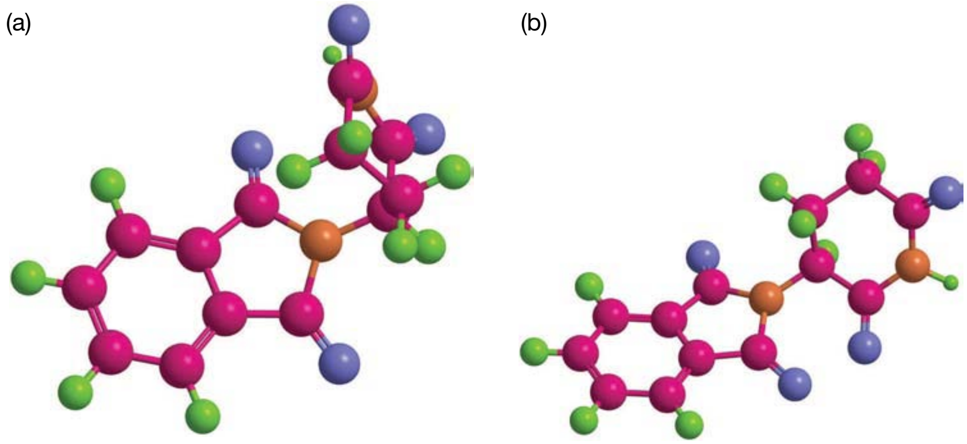
Ethics are also involved in the data and research presented. Authors of research may choose to select the best information to include in their data, leading to a misrepresentation of the information and conclusions drawn. It is important to be able to carefully analyse the research you are investigating for bias (both economic and ethical) that may lead to a lack of validity of results. Research that has been conducted unethically should not be used for your investigation.

An example of this was in the MMR (mumps, measles and rubella) vaccine and autism controversy, in which Andrew Wakefield published a paper arguing a link existed between the MMR vaccination and autism. In this paper (which has since been redacted), data was falsified to show that the MMR vaccination caused autism. A significant bias existed in the individuals chosen to examine as well as issues relating to the limited sample size (only 12 individuals were used in the sample). However, as a result of this paper, vaccination rates began to drop. Despite the paper being redacted and many elements shown to be false, the impact of this unethical paper is still felt today. This shows how important it is to carefully read not just the conclusions drawn in a paper, but the way the data was collected and research carried out.

Use of the chemical thalodomide demonstrates very serious ethical issues in research. In the 1950s, thalidomide was used to treat morning sickness in tens of thousands of pregnant woman. Soon after it was introduced, it was found to cause significant birth defects in babies whose mothers took thalidomide during certain stages of pregnancy. This tragedy highlights the very serious ethical concerns of releasing an inadequately tested drug to the general public. The impact of this drug completely changed the ways new drugs are tested, particularly the ethical considerations in drug trials. It has inadvertedly created another ethical concern; because of this disaster, more animal testing is now conducted.

FIGURE 10.10 An X-ray showing a child whose mother took thalidomide during key stages of the child's development

FIGURE 10.11 Thalidomide exists in two mirror image forms: (a) the R-isomer and (b) S-isomer. One has sedative actions and the other causes birth defects. However, in the body, the two isomers interconvert.



10.3 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Describe why effective science communication is required for your research investigation.
2. Why is important to use standard units in scientific reporting?
3. Explain the difference between models and theories.
4. An investigation is conducted into the synthesis of a new drug used to treat cystic fibrosis. For this research investigation, identify one of each of the following factors: social, economic, environmental and ethical factors.

Fully worked solutions and sample responses are available in your digital formats.

10.4 The nature of evidence and information

KEY CONCEPT

- The nature of evidence and information: distinction between weak and strong evidence, and scientific and non-scientific ideas; and validity, reliability and authority of data, including sources of possible errors or bias

KEY SCIENCE SKILLS

- Analyse and evaluate data, methods and scientific models
- Draw evidence-based conclusions

10.4.1 Weak and strong evidence

In your research investigation, it is important that you can recognise the difference between strong and weak evidence to gain the best possible information for your report.

Research and investigations that have strong evidence include the following features:

- a basis in facts derived from studies with high validity and minimal bias
- statistical evidence to support conclusions
- a clear distinction between **correlation** and **causation** — two variables may often have some correlation (they both increase, for example), but have no causation (one variable does not cause the change in value in the other)

- data from investigations that have a reproducible and reliable method — reliable methods include those that are **double-blind, randomised** and have a **placebo** or a **control**
- peer-reviewed research formed from scientific ideas.

FIGURE 10.12 Strong evidence involves many facets

Weak evidence usually cannot be supported by facts and data. It may be more subject to bias and misinterpretations, may not clearly address the question being asked or may draw conclusions that are inappropriate to the research and data obtained.

10.4.2 Scientific and non-scientific ideas

For your investigation, it is important to distinguish between scientific and non-scientific ideas. Think about the components of scientific reporting — clear evidence; acknowledgement of limitations, uncertainties and errors; a concise investigation including data and an analysis of results; and conclusions relating to a research question.

Many sources, especially on the internet, are not scientific in nature and do not have a high level of validity to use for research. It is important to gain an understanding of the reliability and authority of sources you are using.

10.4.3 Reliability and authority of sources

Finding reliable sources with a high level of authority is vital to allow for the most valid conclusions to be drawn in your investigation.

Government websites and those from educational institutions and established organisations are usually reliable. When using other websites, it is a good idea to look for dates and research the authors to see if the information is verifiable and current. Evaluate your information by comparing it with other websites or other sources. It is important to check for conflicts of interest when examining authors of a site, because this can greatly influence the data shown and the conclusions drawn. You should also check carefully if a website can easily be altered by the public (such as social media), because this is far less reliable than published works. A clue to the source, and potentially the reliability, of a website can be found in the generic top-level domain name. These are the abbreviations found after the 'dot' in a website address. Table 10.3 provides some examples of common generic top-level domain names. Remember that a website address ending in '.au' means that it is registered in Australia, but some non-governmental Australian websites are registered in the United States and will not include the '.au'.

TABLE 10.3 Examples of generic top-level domain names

Domain name contains	Source
.gov.au	Australian federal, state, territory and local government entities; cannot be edited by the general public
.gov	United States government entities; cannot be edited by the general public
.edu.au	Limited to Australian educational institutions including schools and universities; cannot be edited by the general public
.edu	Largely limited to American higher educational institutions (universities and colleges; not schools); cannot be edited by the general public
.org	Intended for use by not-for-profit organisations; some webpages can be edited by the public (for example, Wikipedia.org)
.com.au	Intended for commercial use, but now used within Australia across a range of sectors including businesses, not-for-profits, schools, individual people
.com	Intended for commercial use, but now used within the United States across a range of sectors including businesses, not-for-profits, schools, individual people
.net	Intended for networking technology organisations such as internet service providers, but now used across a range of sectors including businesses, not-for-profits, schools, individual people

It is important to search not just on the web, but also podcasts, journals, videos and books. When using online resources, make use of an advanced search to find a key phrase, recent document or a particular file type such as a PowerPoint (.pptx) or a PDF (.pdf).

Peer-reviewed work

A large amount of information is available to assist with your research investigation.

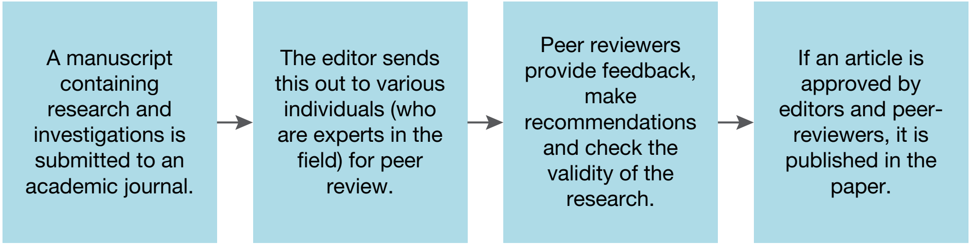
For stronger evidence, it is important to use work that has been **peer-reviewed.** In science, this is work that has been reviewed for quality of research, experimental reproducibility, accuracy and validity, and adheres to the required standards of a journal. Peer-reviewed work is checked for bias before publication is submitted, usually by a board of reviewers or independent experts who have no conflict of interest with the work being examined. The process for peer reviewing is outlined in figure 10.13.

Peer-reviewed work provides stronger evidence, because it has been accepted to be scientifically valid, based on facts, and to draw appropriate conclusions.

Unfortunately, while peer-reviewed articles are useful, they are not perfect. Usually, during reviewing, recommendations are made for required improvements and further explorations. In the case of articles in which in-depth experimental work has been completed, it is often not possible for a reviewer to replicate this work to test its validity and, therefore, it can be hard to determine the accuracy of the outlined conclusions.

It is good to find multiple peer-reviewed journals that show similar ideas and draw similar conclusions. Research articles may also cite other peer-reviewed articles, which may be useful as part of your research investigation.

FIGURE 10.13 Process for articles to be peer-reviewed



Some examples of peer-reviewed chemistry journals include:

- *Accounts of Chemical Research*
- *Acta Crysallographica*
- *Advanced Materials*
- *Angewandte Chemie*
- *Australian Journal of Publishing*
- *Bulletin of the Chemical Society of Japan*
- *Canadian Journal of Chemistry*
- *Chemical Communications*
- *ChemBioChem*
- *Chemistry: A European Journal*
- *Chemical Reviews*
- *Chemistry Letters*
- *Helvetica Chimica Acta*
- *Journal of the American Chemical* Society (JACS)
- *Journal of Chemical Sciences*
- *New Journal of Chemistry*
- *Nature Chemistry*
- *Open Chemistry*.

FIGURE 10.14 Three peer-reviewed academic journals

Another way to access strong sources of evidence is conducting web searches through Google Scholar. Rather than searching for information across all websites, Google Scholar only searches through from academic literature. This may include a variety of information sources from academic publishers, universities and government websites.

10.4.4 Sources of error and bias

Many sources of error and bias exist, and these need to be taken in to consideration when conducting research. No experiment or investigation is perfect; **human error**, **random error** or **systematic errors** can affect the collection and interpretation of data. Errors may include uncalibrated machines being used (leading to results that are constantly inaccurate). Therefore, it is important to not have all your research based on one piece of information, but rather allow for a broader sample to reduce the effect of errors that may have occurred. Errors are further investigated in topic 20 (in section 20.7.4).

Bias also has a great impact on research, in which results are influenced by external factors. Bias may occur for a variety of reasons. Researchers may have a conflict of interest in the investigation, in which specific results are desired. This may be for financial gain, career aspirations or personal reasons. It is important to ensure that minimal bias exists in investigations. Bias is further examined in topic 20 (section 20.5.5.)

Some examples of bias are outlined in table 10.4

TABLE 10.4 Types of bias

Type of bias	Outcome	Example
Measurement	Manipulating results and measurements to get a desired outcome	Changing the results of individuals on a placebo versus an experimental drug
Selection	Not randomly assigning individuals	Deliberately assigning family members or those with a milder form of an illness to the test group
Sampling	Subjects not being representative of the population	Selecting only individuals under 25 to test a new drug designed for woman of all ages
Response	Only certain individuals choosing to be part of a trial, which may not be representative of everyone	Only individuals in a high socioeconomic area chosen to take part

10.4.5 Research investigation checklist

Before completing your research investigation, use the following as your final checklist.

Have you:

☐ understood your topic sufficiently
☐ used your own words
☐ included an introduction with an outline of what the topic is about
☐ communicated the information clearly, logically and accurately
☐ used correct chemical units and terminology
☐ explained important scientific terms
☐ referred to relevant chemical theories or models
☐ checked whether your information is reliable and trustworthy
☐ considered related ethical, environmental and social impacts where necessary
☐ referenced all diagrams and tables
☐ presented the report in an appropriate and succinct form that would be understood by a fellow student at this level
☐ written a concluding paragraph that relates back to the introduction
☐ provided a detailed bibliography?

10.4 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Describe three aspects found in strong evidence as opposed to weak evidence.
2. Describe the difference between scientific and non-scientific ideas.
3. Why is it useful to use journals and articles that are peer-reviewed?
4. A researcher wants to show that a new drug decreases cancer growth by 5%. He selects those with early stage tumours to test the drug on and those with late stage tumours as the control group who are treated with a placebo. Identify and describe what type of bias this is.

Fully worked solutions and sample responses are available in your digital formats.

10.5 Review

10.5.1 Summary

What is a research investigation?

- A research investigation is an exploration of a scientific question or questions and presentation of findings.
- Research allows for the evaluation, interpretation and presentation of information to show trends and patterns, and to draw conclusions.
- An investigation involves the process of selecting a topic and developing a question, conducting research and recording findings, analysing and interpreting information, presenting a report and acknowledging sources.

Effective science communication

- Effective science communication involves accuracy, contextualisation of findings to the research questions, clarity in explaining chemical concepts, ideas and models, conciseness and coherence.
- It is important to be able to clearly incorporate concepts specific to the investigation through the use of definitions, terminology, conventions, units, representations, models and theories.
- It is vital to consider a variety of factors that influence both researched information and the presentation of information. These include social, economic, environmental and ethical factors.

The nature of evidence and information

- It is important to use strong evidence to better back up your statements in your research.
- Strong evidence should use statistical evidence, distinguish between correlation and causation, have high validity and minimal bias and be clearly reproducible.
- The use of peer-reviewed journal articles is a clear example of strong evidence, in which work has been carefully reviewed for accuracy by a panel of subject-matter experts and academics.
- It is important to be aware of the impact of bias and error when using evidence and information, and understand how this affects results and conclusions drawn.

on Resources

Digital document Topic summary — Topic 10 (doc-30961)

10.5.2 Key terms

bias the intentional or unintentional influence on a research investigation

causation when one factor or variables directly influences the results of another factor or variable

control a group that is not affected by the independent variables and is used as a baseline for comparison

correlation measure of a relationship between two or more variables

double-blind type of trial in which neither the participant nor the researcher is aware of which participants are in a control or experimental group

human error mistakes made by individuals conducting an investigation, such as mistakes in entering data or the use of the wrong type of chemical in a reaction

peer-reviewed work that has been reviewed for quality of research, experimental reproducibility, accuracy and validity

placebo substance or treatment that is not designed to affect an individual (used as a control)

plagiarise to copy or use information from a source without referencing it and pass it off as your own

primary data direct or firsthand evidence about some phenomenon

random error chance variations in measurements

randomized making sure that the assigning of individuals to an experimental or a control group is random and is not influenced by external means

secondary data comments on or summaries and interpretations of primary data

systematic error errors that affect the accuracy of a measurement that cannot be improved by repeating an experiment

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Digital documents Key terms glossary — Topic 10 (doc-30958)

UNIT 2

WHAT MAKES WATER SUCH A UNIQUE CHEMICAL?

More than 70% of the Earth's surface is covered by water. Through the water cycle, water is recycled in the ocean, the atmosphere and on the landmass (in rivers, lakes and sub-surface water). The human body also consists largely of water. What is so special about water?

Water's many special properties are related to its structure and bonding. In this unit, we will examine the importance of these properties, and how they shape the role of water as the 'universal solvent' and the reactions of water in selected contexts.

Given the importance of water, we will examine how water is sampled and analysed for salts, organic compounds and contaminants. As the population of the Earth increases, and the effects of climate change alter water's abundance and distribution, water insecurity will increase around the world. An understanding of our most precious resource is fundamental to all life.

AREA OF STUDY	OUTCOME	TOPICS
1. How do substances interact with water?	Relate the properties of water to its structure and bonding, and explain the importance of the properties and reactions of water in selected contexts.	11. Properties of water 12. Water as a solvent 13. Acid–base reactions in water 14. Redox reactions in water
2. How are substances in water measured and analysed?	Measure amounts of dissolved substances in water and analyse water samples for salts, organic compounds, and acids and bases.	15. Water sample analysis 16. Measurement of solubility and concentration 17. Analysis for salts in water 18. Analysis for organic compounds in water 19. Analysis for acids and bases in water
3. Practical investigation	Design and undertake a quantitative laboratory investigation related to water quality, and draw conclusions based on evidence from collected data.	20. Practical investigation

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11 Properties of water

11.1 Overview

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11.1.1 Introduction

If you have ever put a full bottle of water in the freezer and forgotten about it, you no doubt remember what happened. You may have been surprised to find that the bottle had shattered when you removed it. This is due to one of water's unusual properties: it expands on freezing. Water has many interesting properties; for example, it has relatively high melting and boiling temperatures compared to other hydrides of other group 16 elements. Water can absorb a relatively large amount of heat with a correspondingly low temperature rise, which is why it is used as a coolant in car radiators, and why the evaporation of sweat cools the body and contributes to regulating body temperature.

FIGURE 11.1 The evaporation of sweat helps control body temperature.

More than 70% of the Earth's surface is covered by water but only a small percentage is fit to drink. Through the water cycle, water is recycled in the ocean and the atmosphere, and on the landmass (in rivers, lakes and sub-surface water). Water is essential for all life. This small molecule consists of two hydrogen atoms covalently bonded to one oxygen atom. At 25 °C, water is a colourless, transparent and odourless liquid. It freezes at 0 °C and boils at 100 °C. The properties of water that make it crucial for life are generally due to the polar nature and hydrogen bonding between the water molecules. Both intermolecular covalent bonding and intramolecular hydrogen bonding were discussed in topic 6 and are relevant in this discussion of the properties of water. In this topic, these properties are explained in terms of water's structure and bonding. The significance of these concepts for processes in living systems and the environment is examined.

11.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- trends in the melting and boiling points of group 16 hydrides, with reference to the nature and relative strengths of their intermolecular forces, and to account for the exceptional values for water
- specific heat capacity and latent heat including units and symbols, with reference to hydrogen bonding, to account for the relatively high specific heat capacity of liquid water, and significance for organisms and water supplies of the relatively high latent heat of vaporisation of water.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of learning and assessment. Experiments and investigations, supported by a **Practical investigation logbook** and **Teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

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Digital documents Key science skills (doc-30903)
Key terms glossary — Topic 11 (doc-30906)
Practical investigation logbook (doc-30907)

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11.2 Intermolecular forces in water

KEY CONCEPT

- Trends in the melting and boiling points of group 16 hydrides, with reference to the nature and relative strengths of their intermolecular forces, and to account for the exceptional values for water

11.2.1 Hydrogen bonding in water

FIGURE 11.2 A polar water molecule showing polar bonds

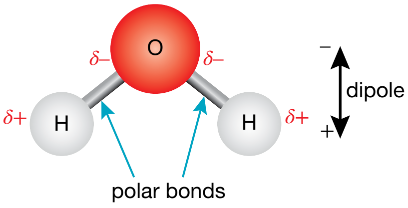
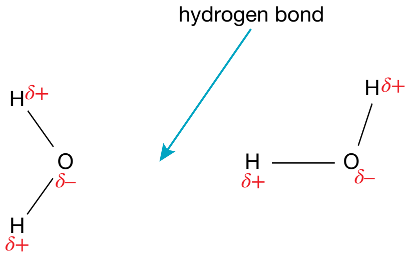


FIGURE 11.3 Hydrogen bond between two water molecules



Water is a small molecule that contains only three atoms. It would, therefore, be expected that the **dispersion forces** between its molecules are quite weak. While this is true, another type of force also exists between water molecules. This force is a **hydrogen bond** and, although it is certainly a lot weaker than the **covalent bonds** that hold the molecule together, it is nevertheless much stronger than the dispersion forces already mentioned. It is an **intramolecular** bond.

Hydrogen bonding is a stronger form of **dipole–dipole bonding** described in detail in topic 6. When a molecule consists of atoms with different electronegativities, the bonding electrons are attracted toward the atom with the stronger **electronegativity**, and this uneven distribution of charge is called a **dipole**. The atom gaining a greater share of the bonding electrons becomes slightly more negatively charged (represented by the symbol $\delta-$, 'delta negative') than the atom that had the lower electronegativity, which then becomes slightly positively charged (represented by the symbol $\delta+$, 'delta positive').

Differences in charges at either end of a molecule produce a **polar** molecule. These slight negative and positive regions cause the molecules to attract other polar molecules. The stronger this **electrostatic** attraction, the closer the molecules become, and the stronger the resulting **intermolecular bond**.

When the atoms making up a polar molecule consist of hydrogen bonded to oxygen, fluorine or nitrogen, a special case exists. The high electronegativity difference causes the hydrogen's electron to move toward the more electronegative atom, resulting in a polar molecule. When this molecule approaches another polar molecule, the slightly positive hydrogen end moves slightly toward the negative end of the other molecule. Hydrogen has only one electron, and when this electron has moved toward the more electronegative atom making up the molecule, no other electrons are left to shield the hydrogen nucleus from the negative region it is approaching. Such molecules can, therefore, attract other molecules very closely, resulting in a stronger bond than would be the case if hydrogen were not involved. This extra-strong intermolecular force is called a hydrogen bond. Hydrogen bonding is fundamental for life processes; these interactions hold the two strands together in the double helix that forms our DNA, for example. Hydrogen bonding significantly affects the melting and boiling points and other physical properties of many compounds, including water.

Water molecules are not only polar, but also contain hydrogen–oxygen bonds, creating hydrogen bonding between the water molecules. This bonding is the reason for the important properties of water.

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Video eLesson Hydrogen bonding (eles-2483)

11.2.2 Melting and boiling temperatures of water

The melting and boiling temperatures of a compound depend on the strength of the intermolecular bonding between the molecules. Water has relatively high melting and boiling points compared with similarly sized molecular substances. More heat is required to enable the molecules to gain sufficient kinetic energy to break free of the hydrogen bonds, which are stronger than dispersion forces alone.

The **boiling temperature** of a substance is the temperature at which it changes from a liquid to a gas. The input of energy is required to cause this change of state. Energy must be provided in this **endothermic** process to enable the molecules to gain sufficient energy to break the intermolecular bonds. The stronger the intermolecular bonds, the higher the boiling temperature. It is well known that the boiling temperature of pure water when measured at sea level is 100 °C. Water boils at lower temperatures at higher altitudes and if not pure; if the water contains dissolved salt, for example, it has a higher boiling temperature.

The **melting temperature** of ice is 0 °C but the presence of hydrogen bonding between water molecules significantly affects what happens at low temperatures. Most substances become denser as the temperature decreases because the particles move closer together. Water, however, is quite different.

Water molecules in liquid form have more energy and move more quickly than water molecules in ice. To change state from liquid to solid, energy must be removed; this is, therefore, an **exothermic** process (figure 11.4). As the temperature is decreased, the molecules start to move more slowly, forming a **crystalline lattice**. The open, crystalline lattice of ice places the water molecules further apart than the random arrangement that occurs in the liquid state. Each water molecule is linked to four other water molecules by hydrogen bonding in a tetrahedral arrangement (see figures 11.5 and 11.6). Because water expands on freezing, it is less dense as a solid. Ice, therefore, floats on water.

FIGURE 11.4 Changing state from a liquid to solid water is an exothermic reaction process because energy is given out.

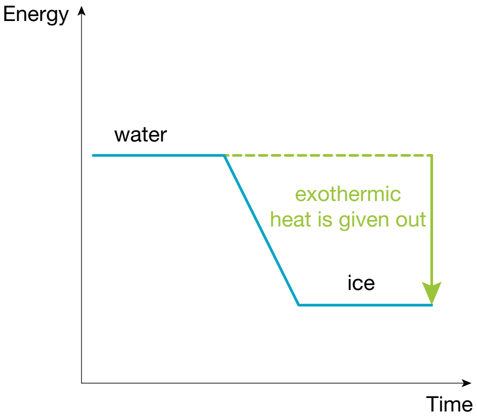


FIGURE 11.5 Arrangement of water molecules in its three different states

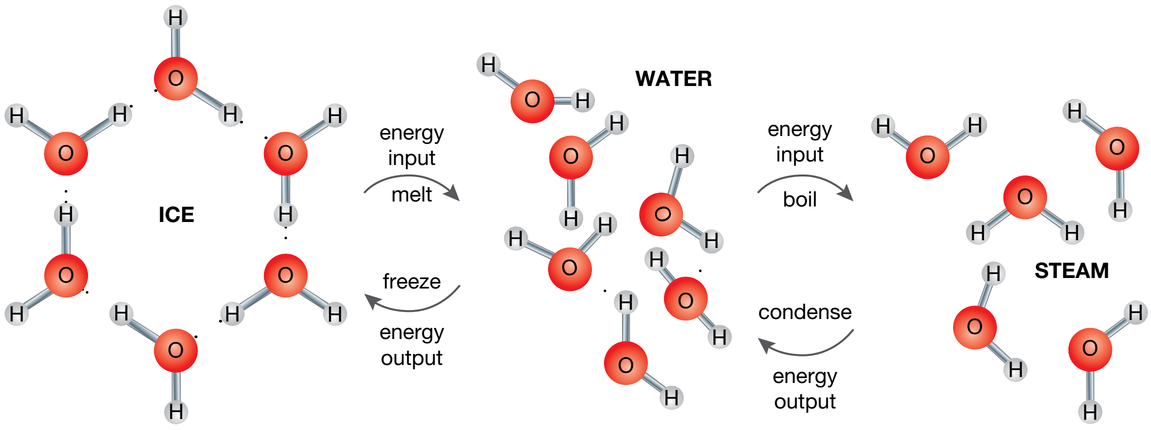
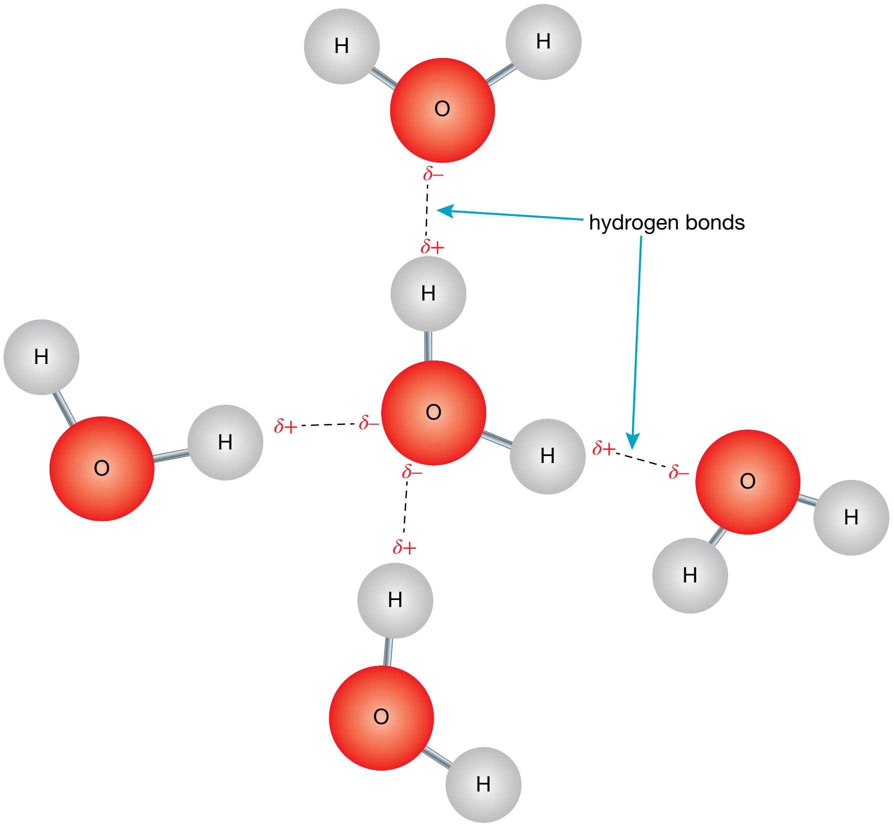


FIGURE 11.6 Each water molecule forms tetrahedral hydrogen bonds to four other water molecules.



Density of water and ice at different temperatures

The **density** of a substance is a measure of the amount of mass that is contained in a certain volume of that substance. It is calculated by using the formula $d = \frac{m}{V}$ (where m is the mass of the substance and V is its volume). Typical units are g mL^{-1} and kg m^{-3}.

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Weblink Water and Ice

SAMPLE PROBLEM 1

If the density of water at 25.0 °C is 0.997 g mL^{-1}, calculate the mass, in grams, of 200.0 mL of water.

THINK	**WRITE**
1. State the given information and identify the quantity required.	$d(H_2O) = 0.997\ g\ mL^{-1}$ $V(H_2O) = 200.0\ mL$ $m(H_2O) = ?$
2. Recall the formula to calculate density is $d = \frac{m}{V}$ and rearrange the formula to find the mass.	$d = \frac{m}{V}$ $m = dV$
3. Substitute known values to calculate mass. **TIP:** Always ensure the answer is given to the least number of significant figures and the units are correct.	$m\,(H_2O) = 0.997 \times 200.0$ $= 199\ g$

PRACTICE PROBLEM 1

Calculate the mass, in grams, of 150 mL water at 25.0 °C given that the density of water is 0.997 g mL^{-1}.

The density of water varies across its three states due to expansion and contraction, and this depends on the water's temperature. Table 11.1 shows the density of ice and water at selected temperatures, and this is demonstrated graphically in figure 11.7.

TABLE 11.1 Density of ice and water at selected temperatures

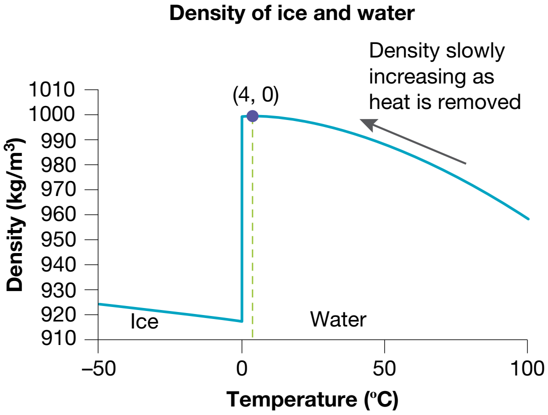
Ice		Liquid water	
Temperature (°C)	Density (ρ) (g cm^{-3})	Temperature (°C)	Density (ρ) (g cm^{-3})
−50	0.922	0	0.999 84
−40	0.921	2	0.999 94
−30	0.920	4*	0.999 98
−20	0.919	6	0.999 94
−10	0.919	8	0.999 85
0	0.916	10	0.999 7
		20	0.998 2
		25	0.997 1
		30	0.995 7

* Water has its highest density at 3.98 °C.

An inspection of table 11.1 and figure 11.7 reveals two interesting facts about water. Firstly, water's maximum density is not at its freezing point but at 4 °C. This means that water that is approaching 0 °C, and is therefore close to freezing, rises to the top because it is less dense than the surrounding water. The second interesting point is that ice has a density that is significantly less than water at the temperatures shown. This is the reason that ice floats on water.

Water has an unusual property: its solid form, ice, is less dense than its liquid form. The molecules in the solid form are arranged in a more structured hexagonal way. This means that ice floats on water and that water freezes from the surface down.

FIGURE 11.7 The density of water changes as it approaches freezing point.



The fact that water freezes with a layer of ice across the top is attributable to both these properties. The water that is about to freeze is at the surface and, when it does turn into ice, its lower density keeps it there. The layer of ice that forms then acts as an insulator, preventing the water below from freezing, and allows aquatic life to survive in sub-zero conditions. This decrease in density from liquid to solid water means that the volume of ice is greater than the corresponding volume of water. This explains why bottles of water crack when frozen. It also results in an important weathering phenomenon where water freezes after seeping into cracks in rocks. As it freezes, enough pressure can build up to crack the rocks and contribute to their eventual erosion.

FIGURE 11.8 When ice forms on the surface of water, it acts as an insulator, preventing the water below from freezing.

11.2.3 Melting point and boiling point of group 16 hydrides

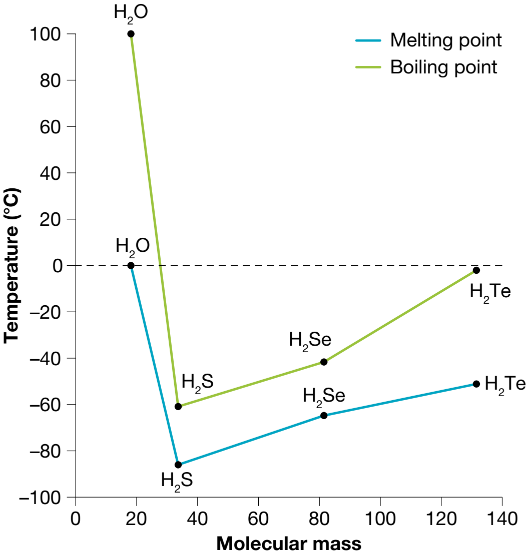
The impact of hydrogen bonding in water can be seen more clearly when we compare melting and boiling temperatures for other **hydrides** in group 16. These temperatures are demonstrated in table 11.2 and figure 11.9.

TABLE 11.2 Melting and boiling temperatures of group 16 hydrides

Hydride	Formula	Molar mass (g mol^{-1})	Melting temperature (°C)	Boiling temperature (°C)
Water	H_2O	18	0	100
Hydrogen sulfide	H_2S	34	−86	−61
Hydrogen selenide	H_2Se	81	−65	−42
Hydrogen telluride	H_2Te	129.6	−51	−2

If either melting or boiling temperatures are compared, water is an anomaly in the group 16 hydrides. If water is ignored, there is an obvious trend of increasing temperatures within the group. The larger the molecule, the higher are its melting and boiling temperatures. This is exactly what would be expected if dispersion forces were the only significant intermolecular force present. The values for water, however, indicate that the intermolecular forces are much stronger than expected for just dispersion forces alone. This is because hydrogen bonding is also present as an additional, much stronger, force between the molecules. Water contains hydrogen–oxygen bonds, which, as mentioned previously, allow for the formation of comparatively strong permanent dipoles called hydrogen bonds.

FIGURE 11.9 Melting and boiling temperatures of group 16 hydrides



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Digital document Experiment 11.1 A water modelling exercise (doc-30859)

11.2 EXERCISE

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1. Why does water have a higher boiling point than other group 16 hydrides?
2. Arrange the following interactions from weakest to strongest: covalent bonding, dispersion forces, hydrogen bonding and dipole–dipole forces.
3. (a) Describe the structure of a water molecule.
   (b) A water molecule described as polar. Explain this statement.
   (c) Draw two water molecules showing correct orientation and label the intramolecular and intermolecular bonds
4. Changing state from solid to liquid is an endothermic process. Explain whether changing state from gas to liquid is an exothermic or endothermic process.
5. Explain if covalent bonds or hydrogen bonds are broken when liquid water evaporates into steam.
6. (a) What quantities are required to find density?
   (b) State a unit for density.
   (c) Is the density of water at 5 °C higher or lower than the density of water at 25 °C?
7. Use the density of water at 25 °C provided in table 11.1 to find the mass (in g) of 65.5 mL of water.
8. Explain why ice floats on water.
9. Use table 11.2 and the graph in figure 11.9 to predict possible melting and boiling points for hydrogen polonide, H_2Po, molar mass (211 g mol^{-1}).
10. The boiling points of water, H_2O, and hydrogen sulfide, H_2S, are 100 °C and –60 °C respectively. Account for the large difference in the two temperatures.

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11.3 Latent heat and specific heat capacity

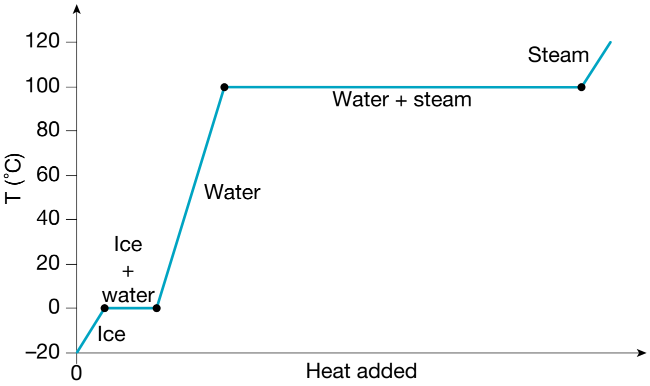
KEY CONCEPT

- Specific heat capacity and latent heat including units and symbols, with reference to hydrogen bonding, to account for the relatively high specific heat capacity of liquid water, and significance for organisms and water supplies of the relatively high latent heat of vaporisation of water

11.3.1 Latent heat

When a substance changes phase (for example, when ice melts and becomes water, or water becomes water vapour), energy is needed. This energy comes from the surrounding atmosphere. The energy needed to change the phase of a substance at its melting or boiling temperature is called the **latent heat** of that substance. When water is changing state, its temperature remains constant so no change occurs in the kinetic energy of the water molecules; the energy comes from or goes to the potential energy stored in the hydrogen bonds between the molecules. That is why it is called latent, or hidden, heat. When ice is melting, the energy supplied is used to break bonds between the molecules and the temperature only begins to rise again once the change of state is complete. The symbol for latent heat is *L*. Common units are kJ mol^{-1} and J kg^{-1}.

FIGURE 11.10 The latent heat of water. Energy is needed to break the bonds between the water molecules. This causes the temperature to remain constant even though the water is being heated.



The **latent heat of fusion (L_f)** of water is the amount of energy needed to change a fixed amount of water from a solid to liquid phase at 0 °C.

To change a fixed amount of water from a solid to liquid phase at 0 °C, the energy required (the latent heat of fusion), breaks the hydrogen bonds between the ice molecules in the crystal lattice. This allows the water molecules to move around more freely in a liquid state. The latent heat of fusion of water is 6.02 kJ mol^{-1}, so 6.02 kJ of energy must be supplied to change each mole of water from solid to liquid.

The **latent heat of vaporisation (L_v)** of water is the amount of energy needed to change a fixed amount of water from a liquid to a gas at 100 °C.

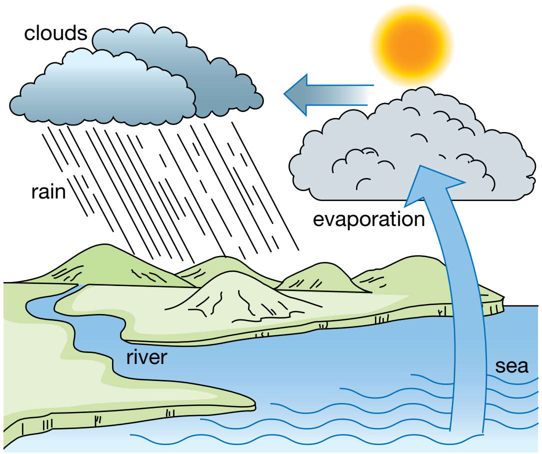
To change a fixed amount of water from a liquid to a gas at 100 °C, the energy required (the latent heat of vaporisation), breaks the remaining forces holding the water molecules together, allowing the molecules to move around freely as a gas. The latent heat of vaporisation is considerably greater than the latent heat of fusion, because the molecules must be completely separated from each other when changing from liquid to gas and not just slightly moved apart as when moving from the crystal lattice in ice to liquid water.

FIGURE 11.11 Sweating uses heat from the body to evaporate the moisture.

The latent heat of vaporisation of water is 40.7 kJ mol^{-1}, which means that 40.7 kJ energy must be supplied to vaporise one mole of water. The relatively high latent heat of vaporisation of water is very useful in keeping organisms cool through perspiration. In humans, when a person perspires, droplets of liquid water on the skin evaporate. This change from liquid water to water vapour requires a large amount of heat energy. This is absorbed from the body, making perspiration an efficient mechanism for removing unwanted heat from the body. It is interesting to note that, if a lot of water vapour is already in the air (as is the case when humidity is high), this process is slow, with the result that people often feel hot and uncomfortable on such days. Condensation, the reverse process of evaporation, would require the release of 40.7 kJ of energy per mole of water.

The higher the latent heat of a substance, the greater the amount of energy it needs when it changes state. Water has much higher latent heat values than other substances of a similar size owing to the stronger intermolecular forces within its solid and liquid states. This means that water evaporates more slowly than many other liquids. Given that we store large amounts of water in dams and lakes, this means that losses due to evaporation from their surfaces are lower than they would otherwise be.

FIGURE 11.12 The relatively high latent heat of vaporisation of water, 40.7 kJ mol^{-1}, means that a large amount of energy from the sun and surroundings is needed for evaporation to occur in the water cycle.



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Video eLesson Phase changes (eles-2487)

Calculations using latent heat

The heat energy required to change the state of a substance is given by the formula:

$$q = nL$$

where:

***q* is the heat energy required or released; kJ**

***n* is the number of mole of substance**

***L* is the latent heat.**

TABLE 11.3 Latent heat values

Substance	Melting point °C	Latent heat of fusion (L_f) kJ mol^{-1}	Boiling point °C	Latent heat of vaporisation (L_v) kJ mol^{-1}
Ethanol, C_2H_5OH	–114	5.02	78.4	38.6
Methanol, CH_3OH	–97.6	3.18	64.7	35.2
Methane, CH_4	–182	0.94	–165	8.18
Ammonia, NH_3	–77.7	5.65	– 33.3	23.4
Water, H_2O	0	6.01	100	40.7

SAMPLE PROBLEM 2

Calculate the heat energy (in kJ) required to fully convert 100 grams of ice to liquid water.

Teacher-led video: SP2 (tlvd-0557)

THINK	**WRITE**
1. State the given information and identify the quantity required.	$m(H_2O) = 100\text{ g}$ $q(H_2O) = ?$
2. To use the latent heat formula, $q = nL_f$, first calculate the number of mole of water, n, where $n = \frac{m}{M}$.	$n(H_2O) = \frac{m}{M}$ $= \frac{100}{(2 \times 1.0 + 16.0)}$ $= 5.56\text{ mol}$
3. Find $q = nL_f$. The latent heat of fusion for water can be found in table 11.3, $L_f = 6.01\text{ kJ mol}^{-1}$. Check significant figures and units are correct.	$q(H_2O) = 5.56 \times 6.01$ $= 33.4\text{ kJ}$

PRACTICE PROBLEM 2

Calculate the heat energy, in kJ, required to fully convert 100 grams of liquid water to steam.

11.3.2 Specific heat capacity

The **specific heat capacity (c)** of a substance is the amount of energy needed to raise the temperature of one gram of the substance by one degree Celsius. It is, therefore, a measure of how easy or how difficult it is to change the temperature of a substance. Water has a high specific heat capacity ($4.18\text{ J g}^{-1}\ ^\circ\text{C}^{-1}$), which means that its temperature is difficult to alter. This is why the beach is popular in hot days. The water can absorb heat from its warmer surroundings so that the air temperature by the seashore is lowered.

FIGURE 11.13 Water has cooling properties.

The energy required to raise the temperature of a particular mass of a substance can be calculated using:

$$q = mc\Delta T$$

where:

q is the energy required in joules

m is the mass in grams

c is the specific heat capacity of the substance

ΔT is the temperature change.

TIP: **This formula can be found in Table 3 of the VCE Chemistry Data Book. Note that heat energy is calculated in joule (not kilojoule) and mass in grams. The symbol Δ (the upper case Greek letter delta) is used to represent a change in a quantity — in this case, a change in temperature.**

Water has a high heat capacity due to the large amount of hydrogen bonding between water molecules. Because temperature is dependent upon the velocities of particles, large amounts of energy are required to overcome the hydrogen bonding, to increase the particles' velocity and hence heat water. This also makes water an excellent insulator of heat. With a heat capacity almost five times higher than soil or rock, water retains heat five times more effectively than land, but may take five times longer than land to heat up. This means water can absorb large amounts of heat energy with only a minimal increase in temperature. On a global scale, short-term extremes in temperature can be evened out by the oceans absorbing heat from the atmosphere in the tropics and releasing it when ocean currents take this water to the polar regions. However, the oceans are struggling to maintain stable temperatures. At the time of writing, 2018 was the warmest year on record for the world's oceans, which provides additional evidence for climate change. Higher surface temperature of the oceans has the added effect of increasing the severity of storms and cyclones.

TABLE 11.4 Specific heat capacity of some common substances

Substance	Specific heat capacity (*c*) ($J\ g^{-1}\ °C^{-1}$)
Water	4.15
Ethanol	2.5
Vegetable oil	2.0
Soil (wet)	1.5
Aluminium	0.90
Concrete	0.88
Glass	0.84
Sand (dry)	0.80
Soil (dry)	0.80
Copper	0.39
Gold	0.13
Lead	0.13

It is possible to measure the energy transferred to water when it is being heated using the formula above.

SAMPLE PROBLEM 3

Calculate how much heat energy (in kJ) is needed to raise 250.0 grams of water from 20.0 °C to 100.0 °C. The specific heat capacity of water is 4.18 J g^{-1} $°C^{-1}$.

Teacher-led video: SP3 (tlvd-0558)

THINK	WRITE
1. State the given information and identify the quantity required.	$m(H_2O) = 250.0$ g $c(H_2O) = 4.18\ J\ g^{-1}\ °C^{-1}$ $\Delta T(H_2O) = 100.0\ °C - 20.0\ °C$ $= 80.0\ °C$ $q(H_2O) = ?$

2. Recall the formula to calculate energy is $q = mc\Delta T$	$q = 250.0 \times 4.18 \times 80.0\,°\text{C}$ $= 8.36 \times 10^4$ J
3. Convert J to kJ by dividing by 1000. Check that significant figures and unit are correct.	$q = \dfrac{8.36 \times 10^4}{1000}$ $= 83.6$ kJ

PRACTICE PROBLEM 3

Calculate the amount of heat energy (in kJ) required to warm 1000 grams of water from the fridge at 4.0 °C to room temperature, 25.0 °C.

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Digital document Experiment 11.2 Measuring specific heat capacity (doc-30858)

Teacher-led video Experiment 11.2 Measuring specific heat capacity (tlvd-0627)

11.3 EXERCISE

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1. What is the name of the energy required to change liquid to gas and vice versa without any change in temperature?
2. Heat energy is continually supplied to keep water boiling at 100 °C. Explain why the temperature remains constant.
3. Water has a specific heat capacity of 4.18 J g^{-1} $°C^{-1}$. Explain what this means.
4. On a hot day at the beach, the sand feels very warm, but the sea is cold. Explain why.
5. Refer to table 11.4 and state whether greater energy is required to raise the temperature of 10 g of copper or 10 g of gold by 1.00 °C.
6. Calculate the amount of energy needed to change the temperature of 300 g of water by 30.2 °C (specific heat capacity of water is 4.18 J g^{-1} $°C^{-1}$).
7. At 25.0 °C, 50.0 g of water absorbs 6274 J of heat. What is its final temperature?
8. A 5.10 kg lump of copper is given 30.0 kJ of energy and it rises in temperature by 15.0 °C. Calculate the specific heat capacity of copper.
9. Calculate the amount of energy needed to change the temperature of 250 g of aluminium from 20 °C to 45 °C (specific heat capacity of aluminium is 0.90 J g^{-1} $°C^{-1}$).
10. Calculate the amount of heat energy (in kJ) that must be released from 152 g of liquid water to convert it to ice.

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11.4 Review

11.4.1 Summary

Intermolecular forces in water

- Water molecules are made up of two hydrogen atoms bonded to an oxygen atom by polar covalent bonds (intramolecular). The presence of two lone pairs repelling the bonding pairs of electrons result in a bent shape and a polar molecule being formed.
- The bonding between water molecules (intermolecular) is hydrogen bonding. A hydrogen bond is formed only when the atoms making up the molecule consist of hydrogen bonded to oxygen, fluorine or nitrogen, which can attract each other very closely.
- The values for the properties of water are higher than for other substances of similar molecular size due to water's strong intermolecular forces.
- Water has a higher density in the liquid state than in the solid state. This is because it expands upon freezing. The more ordered arrangement of water molecules into a crystalline lattice structure is responsible for the expansion of water in the solid state.
- Marine life can survive in winter because the density of ice is less than liquid water. The ice floats on water due to its lower density, providing an insulating layer that prevents the water below from freezing.
- Water has a melting point (of 0 °C) and a boiling point (of 100 °C).
- Water has a higher melting point than any other hydride of group 16 hydrides due to the presence of hydrogen bonding.

Latent heat and specific heat capacity

- The latent heat of a substance is the energy needed to change the state of a substance at its melting or boiling temperature.
- Latent heat can be calculated using the formula $q = nL$ where q is the heat energy required or released in kilojoule, n is the number of mole of water and L is the latent heat.
- Water has a high latent heat of fusion (L_f) — the latent heat of fusion is the amount of energy needed to change a fixed amount of water from the solid to the liquid state.
- Water has a high latent heat of vaporisation (L_v) — the latent heat of vaporisation is the amount of energy needed to change a fixed mass of water from the liquid to the gaseous state.
- While water is changing state when it is being heated, the temperature of the water remains constant. The temperature rises again when the change of state is complete.
- The high latent heat of water means that a large amount of the sun's energy is required to evaporate water from water sources.
- Specific heat capacity is the amount of energy needed to raise the temperature of 1 gram of water by 1 °C. The specific heat capacity of water is 4.18 J g^{-1} $^{o}C^{-1}$.
- To calculate the energy required to change the temperature of a substance, the formula is $q = mc\Delta T$, where q is the energy (in J) required to raise m gram of specific heat capacity c by ΔT degrees Celsius.
- The high specific heat capacity of water means that it can absorb a substantial amount of heat energy without its temperature rising significantly. This makes water useful as a coolant.
- The ability of water to absorb large amounts of heat from the atmosphere with a minimal increase in temperature assists in maintaining a relatively stable environment for ocean organisms.

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11.4.2 Key terms

boiling temperature temperature at which a liquid boils and turns to vapour
covalent bonding sharing of electrons between nuclei that bonds them together in a molecule
crystalline lattice structure of particles that are held together in an ordered, three-dimensional arrangement
density physical property found by dividing the mass of an object by its volume
dipole unequal sharing of electrons between atoms in a molecule
dipole–dipole bonding weak bonding caused by the positive end of one dipole attracting the negative end of another dipole
dispersion force the bond between adjacent molecules formed by instantaneous dipoles
electronegativity the power of an atom to attract a shared pair of electrons
electrostatic force of attraction between positive and negative charges
endothermic chemical process or reaction that absorbs energy
exothermic chemical process or reaction that releases energy through light or heat
hydride binary compound containing hydrogen
hydrogen bond the bond between a hydrogen atom covalently bonded to an atom of F, O or N and another molecule that also contains an atom of H, F, O or N
intermolecular bond bond between molecules
intramolecular bonding bond involving the internal structure of a molecule
latent heat energy needed to change state
latent heat of fusion, (L_f) energy needed to change a fixed amount from the solid to the liquid state
latent heat of vaporization, (L_v) energy needed to change a fixed mass from the liquid to the gaseous state
melting temperature temperature at which state changes from solid to liquid
polar bond bond formed when two atoms that have different electronegativities share electrons unequally
specific heat capacity energy (measured in joules) needed to raise the temperature of 1 g of a pure substance by 1 °C

Resources

Digital document Key terms glossary — Topic 11 (doc-30906)

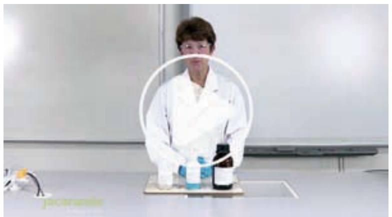
11.4.3 Practical work and experiments

Experiment 11.2

Measuring specific heat capacity

Aim: To measure the specific heat capacity of a number of different liquids

Digital document: doc-30858
Teacher-led video: tlvd-0627



Resources

Digital document Practical investigation logbook (doc-30907)
Teacher-led video Experiment 11.1 A water modelling exercise (doc-30859)

11.4 Exercises

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

11.4 Exercise 1: Multiple choice questions

1. Which one of the following types of bonding would be least likely to be present in a sample of water?
   - A. Ionic
   - B. Covalent
   - C. Dispersion
   - D. Hydrogen
2. In which of the following substances would you NOT expect to find hydrogen bonding?
   - A. HF
   - B. H_2O
   - C. CH_4
   - D. $CH_3CH_2NH_2$
3. What is the hydrogen bonding in water an example of?
   - A. Intermolecular bonding
   - B. Intramolecular bonding
   - C. Covalent bonding
   - D. Ionic bonding.
4. Which of the following processes require the absorption of heat?
   - A. Evaporation
   - B. Condensation
   - C. The transition from a gas to a liquid
   - D. The transition from a liquid to a solid
5. The structure of ice is a crystalline lattice of molecules. The bonding that holds the water molecules in the lattice together is best described as
   - A. hydrogen bonding only.
   - B. dispersion forces and hydrogen bonding.
   - C. ionic bonding.
   - D. covalent bonding.
6. Two solids have different masses but the same volume. The solid with the lower mass
   - A. has a higher density.
   - B. has a lower density.
   - C. has the same density because the volume is the same.
   - D. must be a liquid.
7. The density of water is 0.997 gmL^{-1}, so the mass of 100 mL of water will be
   - A. 0.997 grams.
   - B. 9.97 grams.
   - C. 99.7 grams.
   - D. 997 grams.
8. What is the latent heat of fusion?
   - A. The amount of heat required to raise the temperature of a substance by 1 ^{o}C
   - B. The amount of heat required to increase the temperature of a 1 g of a substance by 1 ^{o}C
   - C. The amount of heat required to change the state of a substance from solid to liquid without any change in temperature
   - D. The amount of heat required to change the phase of a substance from liquid to gas without any change in temperature

9. Why does water have a high heat capacity?
   - A. It can boil at higher temperatures than other liquids.
   - B. It can form hydrogen bonds with any compound.
   - C. It can absorb large amounts of heat before the temperature changes.
   - D. It has a low density as a solid.
10. Why can organisms survive in the ocean even when the temperature is below the freezing point of water?
   - A. Water has a low melting point.
   - B. Ice is denser than water.
   - C. Water has a high heat capacity.
   - D. Ice is less dense than water.

11.4 Exercise 2: Short answer questions

1. Explain why water is different from most liquids with respect to its density in the solid and liquid states.
2. a. What is latent heat?
   b. The latent heat of fusion for water is 6.01 kJ mol^{-1} and the latent heat of vaporisation is 40.7 kJ mol^{-1}. Explain the large difference between these two values.
3. a. What is specific heat capacity?
   b. Explain how water is unusual with respect to this physical property.
   c. Give an example where this property is used.
4. Why do tumble dryers use so much electricity to dry clothes?
5. Cells contain mainly water. If lettuce leaves are placed in the freezer and then defrosted, they become soft and soggy. Explain this observation.
6. Refer to the graph in figure 11.9 and explain why the boiling points of group 16 hydrides decrease and then increase.
7. a. Use table 11.3 to calculate the quantity of heat energy (in kJ) given out when 220 g of steam at 100 °C condenses.
   b. Why is it more dangerous to receive a burn by steam at 100 °C than one by water at 100 °C?
8. a. The temperature of 435 g of water changed from 33.5 °C to 86.4 °C. How much heat energy (in kJ) did this sample absorb?
   b. When 925 J of heat energy is added to a mass of water, the temperature changed by 5.00 °C. Calculate the mass of water used.
9. Calculate the final temperature when 1.05 kg of water at 15.0 °C is supplied with 2.5×10^4 J of energy.
10. 6.10 mole of water at 12.0 °C is heated until its temperature is 37.0 °C. Calculate the amount of heat energy, in kJ, needed to cause this rise in temperature.

11.4 Exercise 3: Exam practice questions

Question 1 (10 marks)

Water has a number of unusual properties.

a. Explain what happens to the density of water as the temperature decreases by referring to the changes in arrangement of its molecules. **5 marks**

b. Why is this significant for organisms that live in the sea? **2 marks**

c. Describe an experiment to compare the effect on density of freezing water and the effect of freezing oil. **3 marks**

Question 2 (3 marks)

When NASA scientists send probes to investigate other planets, one of the first tasks is to search for water. Explain why you think this is the case.

Question 3 (6 marks)

Explain what is occurring at the molecular level when the following observations are made.

a. A damp cloth placed around a water bottle helps to keep it cool. **2 marks**

b. Water droplets form on the sides of a can when it is removed from the fridge. **2 marks**

c. Fresh fish is placed on a bed of crushed ice in the supermarket. **2 marks**

Question 4 (7 marks)

The density of ethanol, C_2H_5OH, is 0.789 g mL^{-1} and specific heat capacity is 2.5 J $g^{-1}\,{}^{o}C^{-1}$.

a. Calculate the mass of the ethanol that fills a 250.0 mL container. **2 marks**

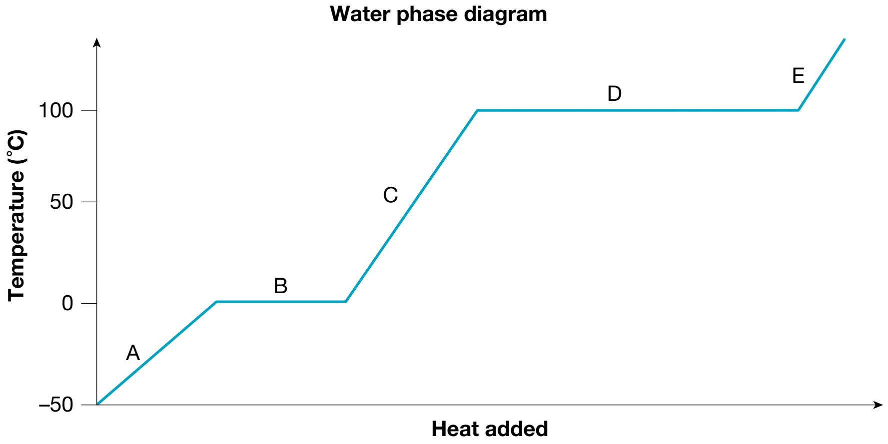
b. Define the term ‘specific heat capacity’. **1 mark**

c. If an equal amount of energy is used to heat 100 g of water and to heat 100 g of ethanol, explain which substance would experience the higher increase in temperature. **2 marks**

d. Calculate the energy (in kJ) required to raise the temperature of 180 g of ethanol by 17 °C. **2 marks**

Question 5 (8 marks)

The following graph indicates what occurs as water is heated over time. Match the letters to the statements, noting that each letter may be used more than once.



a. Molecules have the greatest kinetic energy ___________ **1 mark**

b. Water being heated in a kettle ___________ **1 mark**

c. State change from solid to liquid ___________ **1 mark**

d. Represents the highest latent heat value ___________ **1 mark**

e. Both ice and water present ___________ **1 mark**

f. Latent heat of fusion ___________ **1 mark**

g. Corresponds to boiling occurring ___________ **1 mark**

h. Has the greatest amount of hydrogen bonding ___________ **1 mark**

11.4 Exercise 4: studyON Topic Test online only

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12 Water as a solvent

12.1 Overview

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12.1.1 Introduction

Water is often described as the 'universal solvent' but if this were to mean water can dissolve anything, a number of problems would arise. A major one would be storing water because it would dissolve any container.

Water is, however, an excellent solvent for many ionic and polar substances. We know this because although most of the surface of the Earth is covered with water, very little of it is suitable for human use and only a small fraction of that fresh water is accessible. Pure water does not exist in the environment; all water from nature contains some dissolved and some fine undissolved particles. But why do some substances dissolve and others do not? Previously, when looking at other properties of water, it was shown that hydrogen bonding had a significant impact. How readily a substance dissolves also depends on the polar nature of the water molecules as well as the nature of the substance that dissolves in it. Many ionic compounds seem to disappear when dissolved in water but when oil is added to water it will form a layer on top of it. Some molecular compounds dissolve but others are insoluble. These solvent properties of water are useful in certain situations. The importance of these properties in some biological, domestic and industrial situations is also examined.

Most of the reactions that you will study in this unit occur in water. When two solutions react and produce an insoluble product, this is described as a precipitation reaction. In this topic, you will learn how to write and balance full and ionic equations to represent these reactions.

FIGURE 12.1 Coloured solutions of transition metal salts

12.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- the comparison of solution processes in water for molecular substances and ionic compounds
- precipitation reactions represented by balanced full and ionic equations, including states
- the importance of the solvent properties of water in selected biological, domestic or industrial contexts.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of learning and assessment. Experiments and investigations, supported by a **Practical investigation logbook** and **Teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

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Digital documents Key science skills (doc-30903)
Key terms glossary Topic 12 (doc-30912)
Practical investigation logbook (doc-30913)

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12.2 Solution processes

KEY CONCEPT

- The comparison of solution processes in water for molecular substances and ionic compounds

12.2.1 Solubility

FIGURE 12.2 Dissolving sugar (solute) in a cup of tea (solvent), forming an aqueous solution

Dissolving occurs when one substance is pulled apart by another or moves into the spaces within another. The **solute** is the substance that is dissolved and the **solvent** is the liquid that the solute is dissolved in. Together the solute and the solvent form a **homogeneous** solution. This means that it has uniform composition and properties throughout the whole solution. If the solvent is water, the solution is known as an **aqueous** solution. When sugar is dissolved in a cup of tea, for example, an aqueous solution is formed. Our survival is dependent on reactions that occur in aqueous solutions inside and outside of the cells in our body.

Solubility is the extent to which a solute can dissolve in a solvent. Solubility is determined to a large extent by the polar or non-polar nature of the solute and the solvent. This situation demonstrates the '**like dissolves like**' rule used by chemists; namely, that, if both the solvent and solute particles contain charged regions, it is likely that the solute will dissolve. The process of dissolving depends on the strength of the attractions or bonds between the solute particles, the attractions or bonds between the solvent particles and the new bonds that form between the solute and the solvent. The strength is related to the size of the energy required to break these bonds, compared to the energy released when the new attractions or bonds are formed. In the coming subtopics, the interaction between water molecules and ionic compounds as well as between water molecules and polar molecular compounds will be explored. For some substances, the energy required to break the bonds in either the solute or the solvent is too great and the substance is insoluble.

Water is an effective solvent for many substances due to its polarity and ability to form hydrogen bonds.

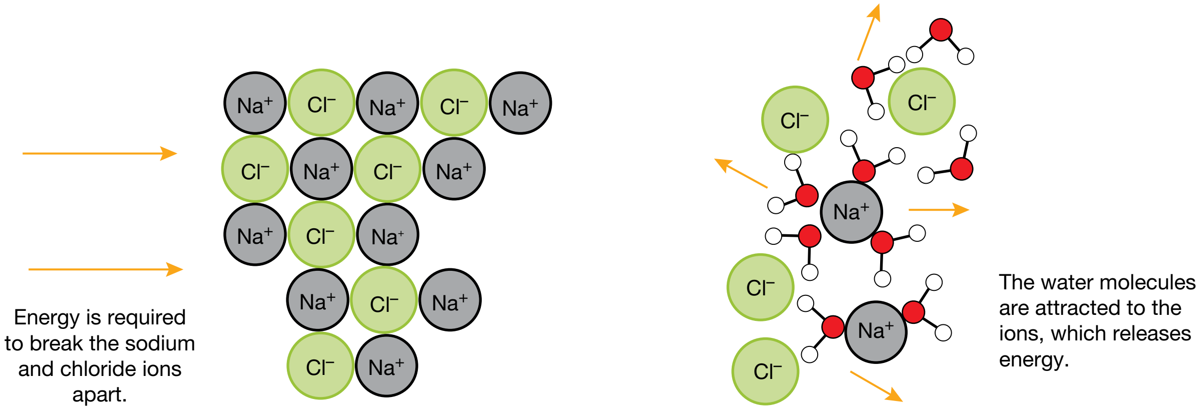
12.2.2 Ionic substances in water

Energy is required to break bonds and energy is released when new bonds form. When **ionic compounds** are mixed with water, the negative ends of the water molecules are attracted to the positive ions (cations) and the positive ends of the water molecules are attracted to the negative ions (anions). If this attraction is strong enough to break the bonds between the cations and the anions, the compound dissolves. The ions move apart and spread throughout the water to form a solution. These ions are described as being **hydrated** and the interaction between a water molecule and an ion is an **ion–dipole attraction**. The ions are surrounded by water molecules and so cannot recombine. The process by which ions separate when an ionic compound dissolves in a solvent is called **dissociation**. An example of dissociation is shown in the following equation:

$$\text{NaCl(s)} \xrightarrow{H_2O} \text{Na}^+\text{(aq)} + \text{Cl}^-\text{(aq)}$$

When ionic substances dissolve in water, they dissociate.

FIGURE 12.3 Sodium chloride dissolving in water



If the strength of the bonds between the ions is very strong, the water molecules are unable to separate the ions and the salt is **insoluble**; for example, silver chloride, AgCl. In this case, not enough energy is released when the ion–dipole interactions are formed to provide the energy needed to break the bonds between the ions.

12.2.3 Molecular substances in water

Molecular compounds are covalently bonded and many do not dissolve in water. As seen in topic 6, the interactions between **covalent** molecules include dispersion forces, dipole–dipole and **hydrogen bonds**.

Water molecules are attracted to each other by comparatively strong hydrogen bonds. Molecules that dissolve in water are highly **polar**, and are described as dipoles. These polar molecules will interact with polar water molecules and dissolve, either because they form ions in water or because they form hydrogen bonds with water.

Molecular compounds that ionise in water

Pure hydrogen chloride, HCl, is a gas consisting of covalently bonded H–Cl molecules, which are highly polar due to the large electronegativity difference between hydrogen and chlorine atoms. Yet, when dissolved in water, HCl reacts to form hydrogen and chloride ions. The hydrogen ions combine with water to form hydronium ions, H_3O^+, and the aqueous solution is called hydrochloric acid. This happens because attraction of the polar water molecules increases the polarity of the H–Cl bond and the slightly negative end of the water molecule attracts the hydrogen atom in HCl, pulling it away from the hydrogen chloride molecule. This hydrogen atom now becomes covalently bonded to the water molecule. Both of these ions are hydrated and form ion–dipole interactions in a similar way to the ions described in the previous section. **Ionisation** is the process by which polar molecular compounds form ions when dissolved in water.

$$\mathrm{HCl(l)} + \mathrm{H_2O(l)} \rightarrow \mathrm{H_3O^+(aq)} + \mathrm{Cl^-(aq)}$$

Highly polar molecular substances ionise in water when they dissolve.

FIGURE 12.4 HCl dissolving in water.

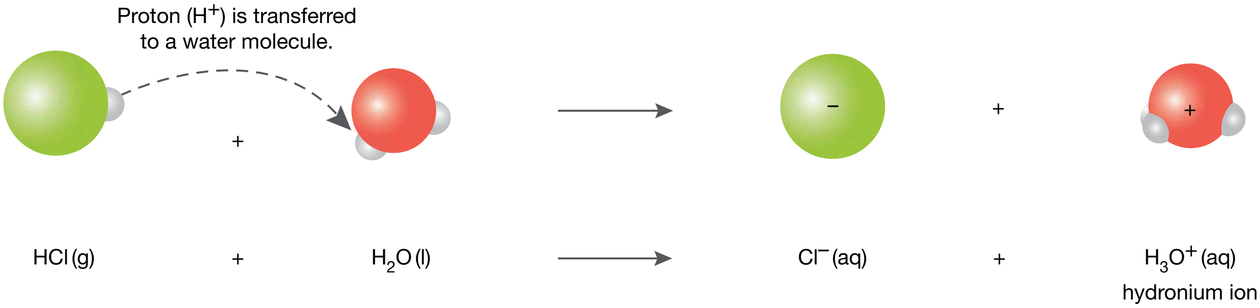
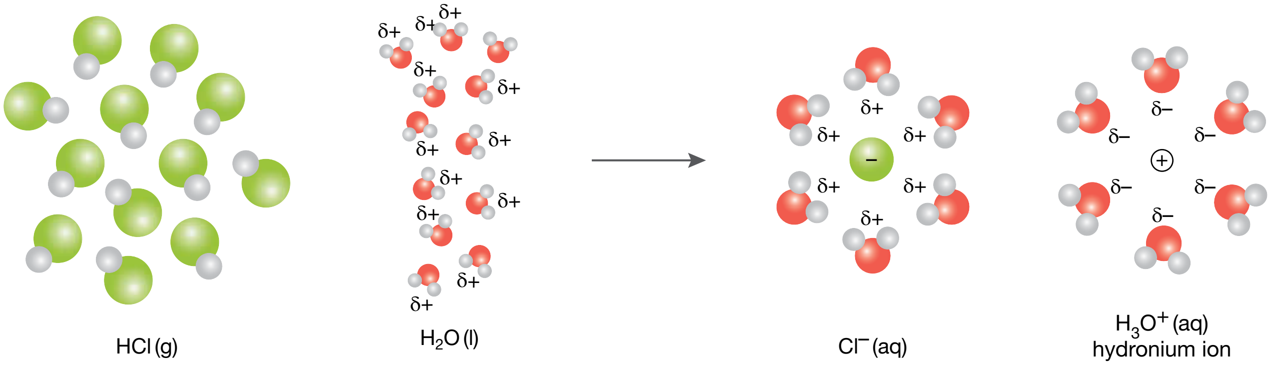


FIGURE 12.5 Hydrated ions with ion–dipole attractions



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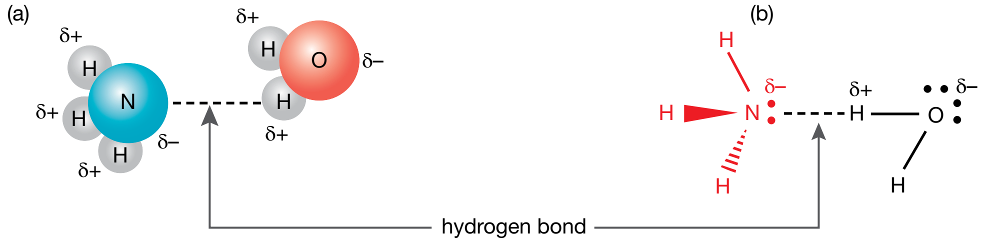
Weblink HCl dissolving in water

Molecular compounds that form hydrogen bonds with water

Other polar molecules, although they do not react with water as outlined in the previous section, may still form hydrogen bonds with water and are soluble. These molecules can be identified because they have a hydrogen atom bonded to one of fluorine, oxygen or nitrogen. Consider ammonia, NH_3: the covalent bonds between the nitrogen atom and hydrogen atoms in the molecule are highly polar due to the electronegativity of nitrogen being considerably greater than that of hydrogen. Water can form a hydrogen bond with nitrogen in ammonia, $NH_3(g) \rightarrow NH_3(aq)$, as shown in figure 12.6.

FIGURE 12.6 Hydrogen bond between ammonia and water molecules shown as (a) a space filling model and (b) Lewis (electron dot) structure

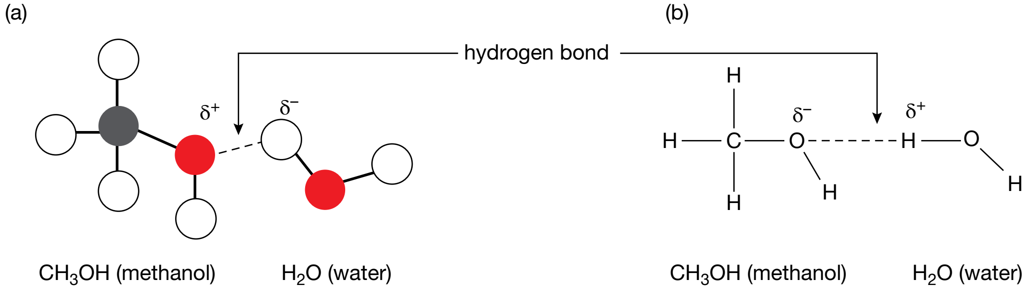


Other molecules that are **soluble** in water include small alcohol molecules. In the simplest alcohol, methanol (CH_3OH), the covalent bond between oxygen and hydrogen in the hydroxy (–OH) group is highly polar. Water can form a hydrogen bond with methanol, as shown in figure 12.7. Hydrogen bonding does not occur between the hydrogen atoms that are bonded to the carbon atoms because the bonds are not very polar — this is due to a relatively small electronegativity difference between carbon and hydrogen.

$$CH_3OH(l) \xrightarrow{H_2O} CH_3OH(aq)$$

It is interesting to note that when water and alcohol molecules interact, they take up less volume due to the rearrangement of the molecules. The volume of the combined liquids is less than the sum of the individual volumes.

FIGURE 12.7 Hydrogen bonding between methanol and water molecules shown as (a) a ball-and-stick model and (b) structural formula



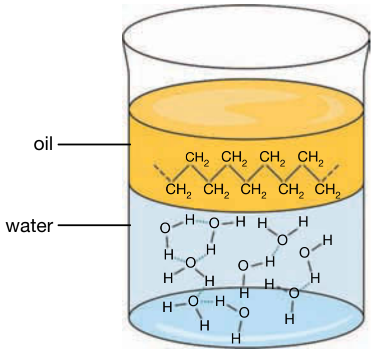
Molecular compounds that contain hydrogen bonded to fluorine, oxygen or nitrogen dissolve by forming hydrogen bonds with water.

Substances that do not dissolve

When two liquids combine to form a solution, they are described as **miscible**. If they do not combine and instead form separate layers, they are described as **immiscible**. For example, when oil is added to vinegar (aqueous solution) in salad dressing, they do not mix and are referred to as immiscible. Non-polar inorganic or organic compounds do not interact with water and so do not dissolve in it. Oil, for example, consists of non-polar molecules, so it does not dissolve in water but will dissolve in non-polar solvents. Non-polar molecules only have weak dispersion forces between them and little energy is required to separate them. As water molecules have stronger hydrogen bonding between them they tend to stick together and not interact with the non-polar molecules.

FIGURE 12.8 Oil is immiscible in water.

FIGURE 12.9 Oil molecules are non-polar and will not dissolve in water.



Polar molecule: In a polar bond, the electronegativity of the atoms will be different.

Non-polar molecule: In a nonpolar bond, the electronegativity of the atoms will be equal.

Polar dissolves polar. → Polar + Polar = Solution

Non-polar dissolves nonpolar. → Non-polar + Non-polar = Solution

Water is polar.
Oil is non-polar. → Polar + Non-polar = Suspension (won't mix evenly)

Organic compounds that have a high proportion of very electronegative atoms such as N or O or, preferably, a number of N–H or O–H bonds are likely to form hydrogen bonds with water and, therefore, be soluble. For example, simple sugars have many hydroxy groups (–OH) and are therefore soluble.

Dissolving gases

Polar gases dissolve easily in water. For example, both ammonia gas, NH_3, and hydrogen chloride gas, HCl, dissolve in water. Both of these gases are composed of polar molecules.

A gas such as oxygen or carbon dioxide is made up of molecules that are **non-polar**. These non-polar gases do not dissolve as easily. Only a very small amount of oxygen gas dissolves in water, but this is enough for cellular respiration in plants and fish. Likewise, the small amount of dissolved carbon dioxide in water enables photosynthesis to take place in aquatic plants.

An increase in pressure on the gas–water mixture causes more of the gas to dissolve in the water. Carbon dioxide is dissolved in water under pressure to make 'carbonated' drinks, for example. When the top is removed from a bottle of carbonated drink, the pressure is released and the gas starts coming out of solution. Carbon dioxide and water form a weakly acidic solution, so soft drink that has been left open for a while loses its carbon dioxide gas and this results in the 'flat' taste.

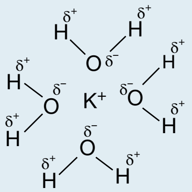
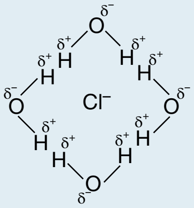
FIGURE 12.10 Gas is dissolved under pressure in carbonated drinks

Temperature also influences the amount of gas that dissolves in water. An increase in temperature generally *decreases* the solubility of gases in liquids. The dissolved oxygen in sea water decreases as the temperature increases and this has significant implications for survival of aquatic organisms.

The solubility of gases decreases with increasing temperature.

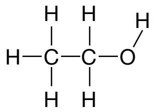
SAMPLE PROBLEM 1

Describe the process by which potassium chloride dissolves in water, including an equation and diagram.

THINK	WRITE
1. Determine the type of bonding present. If it is an ionic compound, bonds will be broken between ions. If it is a polar molecular compound, ions are formed. If it is a molecular compound that contains H bonded to F, O or N, hydrogen bonds are formed.	Potassium chloride is an ionic compound because it consists of a metal and non-metal. Potassium ions and chloride ions are dissociated by the polar water molecules.
2. Interactions occur between the ions and water, and the ions present become hydrated.	The positive potassium ions are attracted to the negative ends of the water molecules by ion–dipole interactions. The negative chloride ions are attracted to the positive ends of the water molecules by ion–dipole interactions.
3. The ions are surrounded by water molecules.	The ions are separated and hydrated.
4. Write the equation and draw a diagram.	$KCl(s) \xrightarrow{H_2O} K^+(aq) + Cl^-(aq)$  

PRACTICE PROBLEM 1

This is the structure of ethanol.



Describe the process by which ethanol dissolves in water.

12.2.4 Electrical conductivity

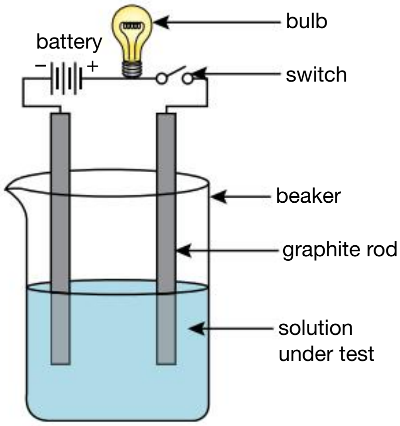
Measuring the electrical conductivity is a useful way to test the quantity of ions formed in a solution. A substance conducts electricity only if it contains charged particles that are able to move. The covalent bonding in water molecules, which results in the formation of neutral molecules, means that it should not conduct electricity — even though these molecules move freely in the liquid state. However, pure water does conduct electricity very slightly. The amount is so tiny that it is still often described as a non-conductor. The reason for this is the ability of water molecules to react slightly with themselves to form a small number of ions. This is called the **self-ionisation** of water and is discussed more fully in topic 13. The equation for this reaction is:

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

Double (reversible) arrows are used to indicate that this reaction is simultaneously reacting in the forward direction and in the reverse direction. The small extent of this reaction means that only a few ions are formed at any instant and the resulting conductivity is virtually zero. This fact may surprise a number of people. We are all accustomed to warnings that water and electricity do not mix and have seen such warnings on appliances such as hair dryers and bathroom heaters. These warnings are necessary because the water that we use every day is never pure, but contains other substances dissolved within it. Often these substances are ionic or react Once such ions are present, the water conducts freely because the charged particles are able to move.

Substances that produce ions when dissolved are classified as **electrolytes**. Mobile ions can be generated in the dissolving process by some ionic compounds and some polar molecular compounds. The degree of solubility of compounds that form ions in solution can be determined by measuring their conductivity. Electrical conductivity can be measured using either a handheld conductivity meter or the apparatus shown in figure 12.11. In this apparatus, the globe lights up if sufficient mobile ions are present to conduct the current.

FIGURE 12.11 Apparatus used to test for conductivity



Conductivity of solutions can be used to compare the quantity of ions present.

12.2 EXERCISES

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1. Explain the difference between the terms ‘solute’ and ‘solvent’. Give an example of each.
2. Describe how the structure of a water molecule makes it a good solvent for ionic and polar molecular compounds.
3. In order to dissolve the salt lithium bromide, LiBr, in water, what types of bonds need to be broken and what interactions are formed?
4. Describe the process by which NaOH dissolves in water.
5. Write dissociation equations for (a) $CuCl_2(s)$, (b) $Na_3PO_4(s)$ and (c) $(NH_4)_2SO_4(s)$.
6. Write ionisation equations for (a) $HNO_3(l)$, (b) $H_2SO_4(l)$ and (c) HBr(g).
7. Why are substances like oil insoluble in water?
8. What is the meaning of the word miscible?
9. Which of the following liquids would be miscible with water: cordial, olive oil, vinegar, wine.
10. Name the interactions that occur when each of the following types of compounds dissolve in water.
    (a) Ionic compound
    (b) Polar molecular compound like HCl
    (c) Molecular compound that contains H bonded to F, O or N

11. The gases NH_3 and HCl dissolve in water. Discuss the similarities and differences in the way the dissolve.
12. What is an electrolyte?

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12.3 Precipitation reactions

KEY CONCEPT

- Precipitation reactions represented by balanced full and ionic equations, including states

12.3.1 Forming a precipitate

Precipitation occurs when two solutions combine and a new compound of low solubility in water is formed, with the unreacted ions remaining in solution. This low-solubility compound forms as solid particles that eventually settle. It is called a **precipitate**. To predict whether a precipitate will form, we must know which substances are soluble in water and which substances are insoluble.

Precipitation reactions are used in many areas of everyday life. Water authorities often add chemicals such as alum (a hydrated double sulfate salt) to their water to deliberately bring about precipitation. As the newly formed precipitate settles, it also traps very fine suspended particles from the water, thus effectively removing them and making the water much clearer. A second common use is the removal of ‘hardness’ from water by adding hydrated sodium carbonate (also known as washing soda). Ions that cause hardness, such as Mg^{2+} and Ca^{2+}, react with the carbonate ions to form the corresponding insoluble carbonate. The ions are removed from the water and cannot react with soap to form a scum.

Precipitation occurs when two solutions combine and a new compound of low solubility in water is formed.

FIGURE 12.12 Precipitate of lead iodide

12.3.2 Rules for solubility

Although water is an excellent solvent for many ionic substances, it cannot dissolve some substances. For substances such as these, water is unable to dissociate the ionic lattice structure and they are therefore insoluble. When different solutions containing dissolved ions are mixed, they will sometimes form an insoluble combination. This produces solid particles that eventually settle out as a precipitate.

To identify the precipitate, refer to a table of solubilities such as shown in table 12.1.

TABLE 12.1 Solubility in water of compounds of common ions

Name of ion	Symbol	Soluble compounds of ion	Insoluble compounds of ion
Group I ions	Li^+, Na^+, K^+, Rb^+, Cs^+, Fr^+	All	None
Ammonium	NH_4^+		
Hydrogen	H^+		
Nitrate	NO_3^-		
Nitrite	NO_2^-		
Chlorides	Cl^-	Most	Ag^+, Pb^{2+}, Hg^{2+} ($PbCl_2$ is moderately soluble in hot water.)
Bromides	Br^-		
Iodides	I^-		Ba^{2+}, Pb^{2+} (Ag_2SO_4 and $CaSO_4$ are slightly soluble.)
Sulfates	SO_4^{2-}		
Carbonates	CO_3^{2-}	Na^+, K^+, NH_4^+	Most
Phosphates	PO_4^{3-}		
Sulfides	S^{2-}	Na^+, K^+	Most (MgS, CaS, BaS, Al_2S_3 and Fe_2S_3 decompose in water.)
Hydroxides	OH^-	Na, K^+, Ba^{2+} (NH_4OH and $(NH_4)_2O$ do not exist as solids.)	Most ($Ca(OH)_2$ is slightly soluble.)
Oxides	O^{2-}		

Note: Most of the ionic compounds not listed here may be assumed to be insoluble in water.

SAMPLE PROBLEM 2

Use table 12.1 to determine whether $PbSO_4$ is soluble.

THINK	WRITE
1. Identify the ions present.	Pb^{2+} and SO_4^{2-}
2. Refer to table 12.1, looking at the symbols column and determine if either ion is listed.	SO_4^{2-} is listed.
3. Identify solubility and check the table for exceptions.	Compounds containing SO_4^{2-} are mostly soluble. $PbSO_4$ is an exception.
4. Answer the question.	$PbSO_4$ is insoluble.

PRACTICE PROBLEM 2

Use table 12.1 to determine whether $MgCl_2$ is soluble.

An easy way of remembering some common soluble and insoluble compounds is to use the SNAPE and CHOPS initials.

Compounds containing these ions are soluble	
Sodium	Na^+
Nitrate	NO_3^-
Ammonium	NH_4^+
Potassium	K^+
Ethanoate	CH_3COO^-

Compounds containing these ions are usually insoluble unless combined with SNAPE ions	
Carbonate	CO_3^{2-}
Hydroxide	OH^-
Oxide	O^{2-}
Phosphate	PO_4^{3-}
Sulfide	S^{2-}

12.3.3 Full equations for precipitation reactions

As covered in section 12.2.2, when a salt is dissolved in water it breaks up, or dissociates, into its constituent ions, which then become hydrated. For example:

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

The symbol (aq) indicates that each ion is separated and has become surrounded by water molecules. Silver nitrate is another salt that is soluble in water.

$$AgNO_3(s) \xrightarrow{H_2O} Ag^+(aq) + NO_3^-(aq)$$

If these two solutions are combined, an insoluble compound, a precipitate, is formed. By examining the solubility table, it becomes apparent that the precipitate observed must be silver chloride. The **full equation** can therefore be written:

$$NaCl(aq) + AgNo_3(aq) \rightarrow NaNO_3(aq) + AgCl(s)$$

That is, the Ag^+ ions combine with the Cl^- ions to form solid AgCl while the Na^+ and NO_3^- ions remain in solution. When equations for precipitation reactions are written, the substance that forms the precipitate is signified by using the solid symbol of state (s). Na^+ and NO_3^- ions are referred to as **spectator ions** because they do not take part in the reaction, as shown in figure 12.13.

FIGURE 12.13 The formation of a precipitate of silver chloride from the mixing of solutions of sodium chloride and silver nitrate

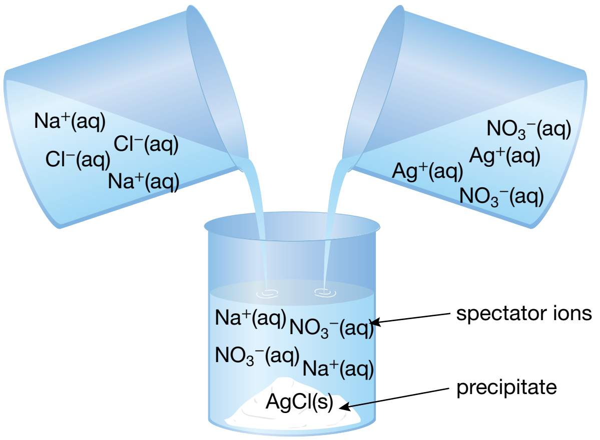

FIGURE 12.14 Silver chloride forms a white precipitate.

Steps for writing balanced precipitation full equations

1. Write out the reactants.
2. Write out the products (swap the anions), ensuring the formulas are correct.
3. Write a skeleton equation (just the reactants and products).
4. Balance the equation.
5. Use the solubility table to find out if one of the products is insoluble. This is the precipitate (sometimes a precipitate may not form and on rare occasions two precipitates are formed).
6. Add states, remembering that the insoluble compound, the precipitate, must be solid (s).

SAMPLE PROBLEM 3

Predict the products and write a balanced equation for the reaction occurring between sodium carbonate solution and copper (II) sulfate solution.

Teacher-led video: SP3 (tlvd-0561)

THINK	WRITE
1. Write the reactants.	$Na_2CO_3 + CuSO_4$
2. Predict the products formed, ensuring that you have written the correct formulas.	$Na_2SO_4 + CuCO_3$
3. Write the skeleton equation for the reaction.	$Na_2CO_3 + CuSO_4 \rightarrow Na_2SO_4 + CuCO_3$
4. Balance the equation.	$Na_2CO_3 + CuSO_4 \rightarrow Na_2SO_4 + CuCO_3$ (No balancing is required.)
5. Refer to the solubility table to determine if either of the compounds is insoluble.	Sodium sulfate is aqueous (since all group I compounds are soluble), but copper (II) carbonate is an insoluble compound.
6. Include symbols of state; the insoluble compound has (s) for the symbol of state.	$Na_2CO_3(aq) + CuSO_4(aq) \rightarrow Na_2SO_4(aq) + CuCO_3(s)$

PRACTICE PROBLEM 3

Predict the products and write balanced equations for the reaction $KOH(aq) + Ca(NO_3)_2(aq)$.

on Resources

Digital document Experiment 12.1 Testing for solubility (doc-30857)

Teacher-led video Experiment 12.1 Testing for solubility (tlvd-0629)

Weblink Precipitation reactions

12.3.4 Ionic equations

In all the preceding equations, the elements and compounds have been written in their molecular or formula unit forms. However, in aqueous solutions, the reactions are best represented by simpler equations called **ionic equations**. Ionic equations are equations that show only the species that are formed or changed in a reaction. Spectator ions are not included in an ionic equation because they remain in solution unchanged in the reaction.

Steps for writing ionic equations

1. Write the balanced full chemical equation as described previously.
2. Expand the equation by dissociating all the soluble compounds into their free ions. Check for any molecular substances such as acids and certain bases that react with water to produce ions. Replace the formulas of these substances by the ions that they form.
3. Cancel all free ions that are unchanged on both sides of the equation (the spectator ions).
4. Write out the ions and formulas remaining. This is the ionic equation.
5. Make sure that equation is balanced in charge as well as in the number of atoms and add symbols of state.
6. Spectator ions are still in solution so write these with commas between (there is no need for coefficients).

SAMPLE PROBLEM 4

Write a balanced ionic equation for the reaction occurring between sodium carbonate solution and copper (II) sulfate solution. State the spectator ions.

Teacher-led video: SP4 (tlvd-0562)

THINK	WRITE
1. Write out the full equation using the steps outlined above.	$Na_2CO_3(aq) + CuSO_4(aq) \rightarrow Na_2SO_4(aq) + CuCO_3(s)$
2. Expand the equation by dissociating or ionising all of the soluble compounds.	$2Na^+(aq) + CO_3^{2-}(aq) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow$ $2Na^+(aq) + SO_4^{2-}(aq) + CuCO_3(s)$
3. Cancel all free ions that are unchanged on both sides of the equation (these are the spectator ions).	$\cancel{2Na^+}(aq) + CO_3^{2-}(aq) + Cu^{2+}(aq) + \cancel{SO_4^{2-}}(aq) \rightarrow$ $\cancel{2Na^+}(aq) + \cancel{SO_4^{2-}}(aq) + CuCO_3(s)$
4. Write out the remaining ions and formula; this is the ionic equation.	$Cu^{2+} + CO_3^{2-} \rightarrow CuCO_3$
5. Check that the equation is balanced and includes states.	$Cu^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CuCO_3(s)$
6. Spectator ions are those that have not reacted.	Na^+, $SO_4{}^{2-}$

PRACTICE PROBLEM 4

Write a balanced ionic equation for the reaction between $MgCl_2$ and NaOH. State the spectator ions.

Spectator ions are omitted in ionic equations

The precipitate formed may be separated from the mixture by filtration, using a funnel, filter paper and a conical flask (figure 12.15). The precipitate remains in the filter paper and is referred to as the **residue**, and the clear solution in the conical flask is the **filtrate** and contains the spectator ions. The filter paper with the residue can be placed in an oven at 110°C to evaporate the water.

FIGURE 12.15 Filtration apparatus (funnel, filter paper and a conical flask) used to separate precipitate from solution

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Weblink Balancing ionic equations

12.3 EXERCISE

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1. (a) What is a precipitation reaction?
   (b) Describe what you would observe by mixing of two solutions which then form a precipitate.
2. State the names and formulas of the ions present in the following solutions.
   (a) K_2SO_4
   (b) $(NH_4)_3PO_4$
   (c) $AgNO_3$
3. Use table 12.1 to determine whether each of the following ionic compounds is soluble.
   (a) $BaSO_4$
   (b) $Pb(NO_3)_2$
   (c) AgI
   (d) sodium sulfide.
4. Explain which of the following reactions would be classified as precipitation reactions.
   (a) $2HCl(aq) + Mg(s) \rightarrow MgCl_2(aq) + H_2(g)$
   (b) $Ca(OH)_2(s) \rightarrow CaO(s) + H_2O(g)$
   (c) $K_2CO_3(aq) + CaCl_2(aq) \rightarrow CaCO_3(s) + 2KCl(aq)$
5. Use table 12.1 to determine
   (a) which sodium compounds are soluble.
   (b) which potassium compounds are insoluble.
   (c) which silver compounds are soluble.
6. Predict which of the following pairs of solutions will form a precipitate when mixed.
   (a) Iron (III) nitrate and sodium hydroxide
   (b) Barium chloride and copper sulfate
   (c) Sodium nitrate and ammonium carbonate
7. Predict products and write balanced equations for the following.
   (a) $2LiOH(aq) + Ca(NO_3)_2(aq)$
   (b) $Na_2S(aq) + (CH_3COO)_2Pb(aq)$
   (c) $2(NH_4)_3PO_4(aq) + 3CaCl_2(aq)$
8. Write ionic equations for the following reactions and identify the spectator ions.
   (a) Barium chloride and magnesium sulfate solutions reacting to form a precipitate of barium sulfate and magnesium chloride solution
   (b) Calcium nitrate reacting with sodium carbonate in solution to form solid calcium carbonate and sodium nitrate solution
9. Suggest a method of obtaining a sample of solid zinc carbonate from solutions of zinc nitrate and potassium carbonate.
10. Zahra wishes to test a solution for the presence of chloride ions. Name a soluble salt solution she could add to produce a precipitate containing chloride ions, and write the ionic equation for the reaction.

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12.4 Solvent properties of water in biological, domestic or industrial concepts

KEY CONCEPT

- The importance of the solvent properties of water in selected biological, domestic or industrial contexts

We know that water is particularly useful as a solvent because of its ability to dissolve an enormous number of substances. Water can transport dissolved chemicals around our body and act as a medium for chemical reactions to occur. It is an important solvent in the home, and for environmental, industrial and agricultural processes.

12.4.1 Solvent properties of water in biological contexts

FIGURE 12.16 Plants and fish rely on gases dissolved in water to survive.

The bodies of most organisms are made up of a large percentage of water. Humans, for example, are about 60% water. Blood plasma and cells contain dissolved substances such as amino acids, proteins, mineral ions, various types of sugars, some vitamins, digestion products and hormones, to name just a few. Waste products are also present, such as urea and carbon dioxide, and these are carried away by water. In plants, water in the cells and sap has a wide range of dissolved materials, including nutrients, amino acids and sugars produced from photosynthesis. The oxygen and carbon dioxide dissolved in waterways are essential for the survival of aquatic animals and plants.

Problems may be caused by the solubility of fertilisers, particularly those containing nitrogen and phosphorus, which are transported through the ground into rivers and lakes and encourage the growth of algae and plants in the water. These eventually die and the decay process uses up oxygen, which results in the death of fish. This process is known as **eutrophication** and is discussed in topic 15.

12.4.2 Solvent properties of water in domestic contexts

In the garden, many pesticides and nutrients are dissolved in water before being applied. Vinegar is a solution of ethanoic acid. Beverages contain a number of solutes responsible for their flavours, and aerated drinks contain dissolved carbon dioxide. In the medicine cabinet are many aqueous solutions or products that are dissolved in water before being taken. Personal care products such as perfumes and mouthwashes can be added to this list and soaps and detergents dissolve in water to help keep household items and ourselves clean.

12.4.3 Solvent properties of water in industrial contexts

Many chemicals are more easily and safely handled and transported as solutions than in their pure states. Water facilitates many reactions by providing a convenient medium to bring reactants together. In electrolysis reactions, solutions in water often provide a less expensive lower temperature alternative to using molten materials to achieve conductivity. From a safety point of view, chemicals can be diluted so that they pose less of a risk — larger volumes of more dilute solutions can be used instead of smaller volumes of more concentrated and potentially dangerous solutions. Water can be used to conveniently capture and dissolve by-products of industrial reactions, allowing them to be turned into useful products or to be processed further before release into the environment. Water is used as a solvent in many manufacturing industries, including the production of foods, medicines, fertiliser, detergents, paints and paper.

A special mention should also be made of soaps and detergents. These products are part of a larger group called **surfactants** and they contain molecules that are polar at one end and non-polar at the other. This enables them to dissolve in both water and oil, making them ideally suited as cleaning agents where oil and grease need to be removed.

FIGURE 12.17 Two soluble aspirin tablets in glass of water

12.4 EXERCISE

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1. Describe how the solvent properties of water are of benefit in the human body.
2. Explain how dissolved substances keep plants and fish alive in a pond.
3. List two examples of water acting as a solvent in a meal.
4. Why is it important that fertilisers are soluble in water and what problem may result from this?
5. How is water useful in the industrial production of materials?
6. What disadvantages might there be in the use of water in chemical industries?

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12.5 Review

12.5.1 Summary

Solution processes

- Water is an excellent solvent for many ionic and polar substances. This is due to its polarity and ability to form hydrogen bonds. Water is, therefore, attracted to ions and polar substances. If these attractions are strong enough to overcome the attractions between the molecules or ions of the other substance, that substance dissolves (the 'like dissolves like' rule).
- A solute is a substance that is dissolved in a solvent to make a solution.
- Solubility is the extent to which a solute can dissolve in a solvent and this depends on the temperature.
- The process of dissolving depends on
  - The strength of the bonds between the solute particles and the energy required to separate them.
  - The strength of the bonds between the solvent particles and the energy required to separate them.
  - The amount of energy released when new interactions or bonds are formed between the solute and the solvent.
- Ionic compounds that dissolve do so because they dissociate in water, forming hydrated ions.

- Dissociation is the process by which particles in the solid state are pulled apart by water and become able to move around. This results in the solid dissolving and a solution being formed.
- Molecular compounds that dissolve either
  - ionise in water (when the molecules are highly polar)
  - form hydrogen bonds (when the molecules contain H bonded to F, O or N) with water.
- Polar gases such as ammonia and hydrogen chloride dissolve readily in water.
- Non-polar gases such as oxygen and nitrogen have low solubility, but enough to sustain aquatic life.
- An increase in pressure generally increases the solubility of a gas in liquids.
- An increase in temperature generally decreases the solubility of gases in liquids.
- Conductivity of solutions can be used to determine the quantity of ions formed.

Precipitation reactions

- Precipitation reactions occur when solutions of soluble substances are mixed and an insoluble substance is formed.
- A solubility table may be used to predict the formation of a precipitate when solutions are mixed.
- Salts that contain any of the ions Na^+, NO_3^-, NH_4^+, K^+, and CH_3COO^- (SNAPE ions) are soluble.
- Salts that contain CO_3^{2-}, OH^-, O^{2-}, PO_4^{3-} and S^{2-} (CHOPS ions) are usually insoluble unless combined with the SNAPE ions.
- An ionic equation is one that shows only the ions that are formed or changed during a chemical reaction.
- Those ions that do not react are known as spectator ions and are not shown in ionic equations.
- Ionic equations are particularly useful for summarising precipitation reactions.

Solvent properties of water in biological, domestic or industrial concepts

- Biological processes involve the transportation of chemicals and waste products, and are dependent on the solvent properties of water.
- Many solutions are common in households, including drinks, detergents and medications.
- Water is an important solvent for the industrial manufacture of various materials; for example, pharmaceuticals, fertilisers and dyes.

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12.5.2 Key terms

aqueous solution with water as the solvent
covalent bond sharing of electrons between nuclei that bonds them together in a molecule
dissociation the process by which ions separate when an ionic compound dissolves in a solvent
double (reversible) arrows arrows used in an equation for a reaction where the forward reaction is occurring at the same time as the reverse reaction
electrolyte solution or liquid that can conduct electricity
eutrophication excess nutrients in rivers and lakes that promote growth and decay of plant life to the detriment of other species
filtrate liquid that has passed through a filter
full equation equation that includes all of the reactants and products that are present
homogeneous uniform composition throughout
hydrated containing water molecules within the structure

hydrogen bond the bond between a hydrogen atom covalently bonded to an atom of F, O or N and another molecule that also contains an atom of H, F, O or N
immiscible when two substances are unable to form a homologous solution
insoluble will not dissolve
ion–dipole attraction attraction between an ion and a polar molecule
ionic equation equation without the spectator ions
ionisation reaction in which a substance reacts with water to produce ions
like dissolves like strongly polar and ionic solutes tend to dissolve in polar solvents, and non-polar solutes tend to dissolve in non-polar solvents
miscible able to be mixed and form a solution
molecular group of atoms bonded together covalently
non–polar molecules that have an even distribution of charge
polar molecules that have an uneven distribution of electrons due to differences in electronegativity of the atoms
precipitate solid with low solubility formed in a liquid
precipitation reaction reaction where two solutions are mixed and a solid is formed
residue substance remaining in filter paper after the process of filtration
self-ionisation reaction in water where water ionises very slightly to form H_3O^+ and OH^-
solubility grams of a substance that will dissolve in an amount of solvent at a particular temperature
soluble will dissolve
solute dissolved substance in a solution
solvent substance in which the solute is dissolved
spectator ion ions that take no part in a reaction but maintain electrical neutrality in a solution
surfactant a molecule that has a polar (hydrophilic) and a non-polar (hydrophobic) end; e.g., a detergent molecule

Resources

Digital document Key terms glossary — Topic 12 (doc-30912)

12.5.3 Practical work and experiments

Experiment 12.1

Testing for solubility

Aim: To investigate the formation of precipitates in aqueous solution and gain skill in writing ionic equations to represent them

Digital document: doc-30857
Teacher-led video: tlvd-0629

Resources

Digital document Practical investigation logbook (doc-30913)

12.5 Exercises

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12.5 Exercise 1: Multiple choice questions

1. Which properties explain the ability of water to act as a solvent?
   (i) Hydrogen bonding, (ii) polarity of water molecules or (iii) high specific heat capacity of water?
   A. (i) and (ii) only
   B. (i) and (iii) only
   C. (ii) and (iii) only
   D. (i), (ii) and (iii)
2. Which of the following terms best describes a substance that dissolves when added to water?
   A. Solvent
   B. Solute
   C. Insoluble
   D. Solution
3. Which of the following is NOT a characteristic of a solution?
   A. Homogeneous
   B. Miscible
   C. Solute and solvent cannot be distinguished by the naked eye
   D. Solute particles can be separated by filtering
4. The following interactions are involved in the solution process:
   (i) solute–solvent, (ii) solute–solute and (iii) solvent–solvent.
   Which of the interparticle attractions must be overcome in order for a solute to dissolve in a solvent?
   A. (i) and (ii)
   B. (i) and (iii)
   C. (ii) and (iii)
   D. (i), (ii) and (iii)
5. In general, to increase the solubility of a gas in water it is necessary to have:
   A. high temperature and high pressure.
   B. low temperature and high pressure.
   C. high temperature and low pressure.
   D. low temperature and low pressure.
6. For which of the following types of ionic compounds are most examples insoluble in water?
   A. Nitrates
   B. Sulfides
   C. Chlorides
   D. Ethanoates
7. Which of the following includes only water-soluble substances?
   A. Ammonium chloride, magnesium nitrate, sodium hydroxide, sodium sulfide
   B. Ammonium nitrate, magnesium hydroxide, ammonium hydroxide, sodium chloride
   C. Magnesium chloride, sodium nitrate, silver chloride, potassium chloride
   D. Potassium nitrate, magnesium nitrate, copper sulfide, copper nitrate
8. Which of the following compounds would appear as a precipitate in solution?
   A. Ammonium sulfate
   B. Barium nitrate
   C. Barium sulfate
   D. Ammonium hydroxide

9. An ionic equation does *not* show
   A. reactants.
   B. charges on ions.
   C. solid products.
   D. spectator ions.
10. Which of the following is a correctly balanced ionic equation?
   A. $Ba^{+}(aq) + SO_4^{-}(aq) \rightarrow BaSO_4(s)$
   B. $HSO_4^{-}(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq) + H_3O^{+}(aq)$
   C. $Na^{+}(aq) + 2Cl^{-}(aq) \rightarrow NaCl_2(s)$
   D. $2H^{+}(aq) + SO_4^{2-}(aq) + 2K^{+}(aq) + 2OH^{-}(aq) \rightarrow 2K^{+}(aq) + SO_4^{2-}(aq) + 2H_2O(l)$

12.5 Exercise 2: Short answer questions

1. a. Explain why only one hydrogen atom of the six present in ethanol (CH_3CH_2OH) can form a hydrogen bond with a water molecule.
   b. Why is ethanol (CH_3CH_2OH) soluble in water and ethane (CH_3CH_3) insoluble?
2. Explain the process by which silver nitrate dissolves in water.
3. Write dissociation equations for the following.
   a. $KOH(s)$
   b. $Na_3PO_4(s)$
4. Is there more dissolved oxygen in warm water or cold water? Why is this significant for aquatic organisms?
5. Complete the table using the following compounds: CH_4, KI, HNO_3 and CH_3OH.

Description	Compound
This compound is insoluble in water	
This compound ionises when dissolved in water	
This compound forms hydrogen bonds when dissolved in water	
This compound dissociates when dissolved in water	

6. Use the rules for solubility to decide whether or not each of the following substances would be soluble in water.
   a. Potassium nitrite
   b. Sodium chloride
   c. Lead sulfate
   d. Calcium carbonate
   e. Ammonium hydroxide
   f. Copper (II) sulfide
7. Chemists often refer to the ‘like dissolves like’ rule when describing the solubility of various substances in different liquids. How does this rule explain the following?
   a. Potassium chloride dissolves in water
   b. Ammonia dissolves in water
   c. Methane is insoluble in water
8. Determine whether the following substances are soluble in water or not.
   a. CaI_2
   b. KOH
   c. AgCl
   d. AgI
   e. $BaSO_4$
   f. NH_4Cl
   g. Na_2S
   h. $MgCl_2$
   i. CuS
   j. $Al(OH)_3$
9. Complete the following equations.
   a. $BaCl_2(aq) + (NH_4)_2CO_3(aq)$
   b. $2AgNO_3(aq) + H_2S(g)$
   c. $K_2CrO_4(aq) + Pb(NO_3)_2(aq)$

10. Predict whether a precipitate forms when solutions of the following are mixed.
    - a. $AgNO_3$ and KCl
    - b. H_2SO_4 and $BaCl_2$
    - c. H_2SO_4 and NaCl
    - d. $NaNO_3$ and KCl
    - e. $(NH_4)_3PO_4$ and $CaCl_2$
    - f. $(NH_4)_2S$ and $Pb(NO_3)_2$
    - g. $(NH_4)_2S$ and $NaNO_3$
    - h. $CaCl_2$ and NaBr
11. Write ionic equations for the reactions that form precipitates in question 10.
12. Write a full equation and an ionic equation for each of the following reactions (if no reaction occurs, write 'no reaction').
    - a. Solutions of potassium sulfate and calcium nitrate are mixed.
    - b. Lead (II) nitrate solution is added to potassium iodide solution.
    - c. Magnesium chloride is added to sodium phosphate solution.
    - d. Solutions of potassium sulfide and zinc chloride are mixed.
    - e. Solutions of sodium sulfate and potassium bromide are mixed.
13. What cations could be present in a solution that did the following.
    - a. Gave a precipitate with sodium sulfate solution but not with sodium chloride.
    - b. Gave a precipitate with sodium sulfate and with sodium chloride.
    - c. Gave a precipitate with sodium sulfate solution but not with sodium hydroxide.
    - d. Gave a precipitate with sodium carbonate but not with sodium hydroxide.
14. Describe simple tests that would allow you to distinguish between solutions of the following.
    - a. Silver nitrate and zinc nitrate
    - b. Barium chloride and magnesium chloride
    - c. Silver nitrate and lead nitrate
    - d. Potassium sulfate and potassium chloride

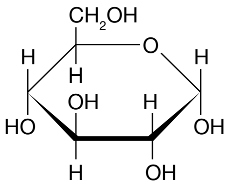
12.5 Exercise 3: Exam practice questions

Question 1 (6 marks)

Explain the following observations.

a. Glucose, $C_6H_{12}O_6$ (shown), is very soluble in water but does not conduct electricity when dissolved. **2 marks**

Glucose



b. Sodium chloride is soluble in water but insoluble in benzene, which is a non-polar liquid. **2 marks**

c. The bonds between water molecules are stronger than the bonds between hydrogen chloride molecules. **2 marks**

Question 2 (6 marks)

Compare the different processes that occur when hydrogen bromide gas, HBr, and ammonia, NH_3, dissolve in water. Include both diagrams and equations.

Question 3 (6 marks)

Solutions of sodium chloride and hydrochloric acid are both good conductors of electricity. This means that they both have mobile ions present.

a. Define the term 'dissociation'. **1 mark**

b. Define the term 'ionisation'. **1 mark**

c. Explain why a solution of sodium chloride can conduct electricity. Use the appropriate term from parts (a) and (b) in your explanation. **2 marks**

d. Explain why hydrochloric acid can conduct electricity. Use the appropriate term from parts (a) and (b) in your explanation. **2 marks**

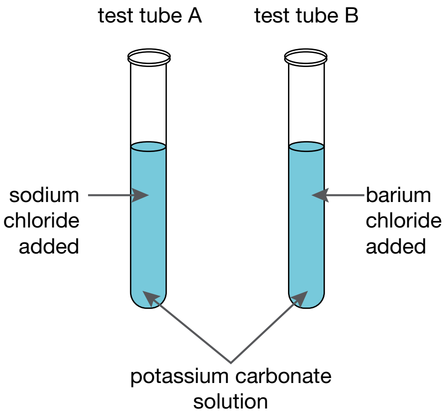
Question 4 (3 marks)

a. From the following list of substances, choose a substance that is insoluble in water: $CuSO_4$, $Ba(OH)_2$, KOH, $BaSO_4$. **1 mark**

b. Write a chemical equation for the reaction between two reactants of your choice that would produce the precipitate you chose in part (a). **2 marks**

Question 5 (8 marks)

Some ammonium carbonate solution is poured into two separate test tubes. Some sodium chloride solution is then added to test tube A, and some barium chloride solution is added to test tube B.



a. In which test tube would a precipitate be observed to form? **1 mark**

b. Write the formula equation for the formation of the precipitate that you have identified from part (a). **2 marks**

c. Write the ionic equation for this precipitation reaction. **1 mark**

d. What are spectator ions? In the example from part (c), list the spectator ions that are present. **2 marks**

e. Suppose that the precipitate is filtered and the remaining liquid (the filtrate) is then heated to evaporate the remaining water. Describe what happens, naming any new substances that might be formed. **2 marks**

12.5 Exercise 4: studyON Topic Test online only

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Test maker

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13 Acid–base reactions in water

13.1 Overview

Numerous **videos** and **interactivities** are available just where you need them, at the point of learning, in your digital formats, learnON and eBookPLUS at www.jacplus.com.au.

13.1.1 Introduction

The notion of acids and their presence in citrus fruits with their sour taste is familiar to most people, as is acid in the stomach to help break down food. Perhaps you also know about acid in car batteries, where it is used as an electrolyte. Less commonly known is information about the chemical opposite of acids, which are called bases. Bases neutralise acids and are also regularly used in everyday life. Baking soda, otherwise known as sodium hydrogen carbonate, has a multitude of uses from cooking to cleaning. Magnesium hydroxide is taken as an antacid and ammonia is widely used as a cleaning agent. Bases are particularly good at dissolving grease. Blocked drains and stained ovens require the heavy-duty power of the strong base sodium hydroxide which, like many acids, is highly corrosive and needs to be used with care.

Most people know more about what acids and bases do rather than what they are. This topic provides definitions of acids and bases, and explains how acidity is measured and how acids and bases react. The distinction between strong and concentrated acids and bases in terms of numbers and types of particles present is also explained. The reactions of acids and bases with water and with each other are closely linked to their interactions with water, which were discussed in the previous topic.

FIGURE 13.1 Baking powder has many uses, including cooking and cleaning.

13.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- the Brønsted–Lowry theory of acids and bases, including polyprotic acids and amphiprotic species, and writing of balanced ionic equations for their reactions with water including states

- the ionic product of water, the pH scale and the use of pH in the measurement and calculations of strengths of acids and bases and dilutions of solutions (calculations involving acidity constants are not required)
- the distinction between strong and weak acids and bases, and between concentrated and dilute acids and bases, including common examples
- the reactions of acids with metals, carbonates and hydroxides, including balanced full and ionic equations, with states indicated
- the causes and effects of a selected issue related to acid–base chemistry.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of learning and assessment. Experiments and investigations, supported by a **Practical investigation logbook** and **Teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

on Resources

Digital documents	Key science skills (doc-30903)
	Key terms glossary — Topic 13 (doc-30927)
	Practical investigation logbook (doc-30928)

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To access key concept summaries and practice exam questions download and print the **studyON: Revision and practice exam question booklet** (doc-30929).

13.2 Acids and bases

BACKGROUNG KNOWLEDGE

- All acids share common properties, including a sour taste, will turn litmus red and neutralise bases.
- All bases share common properties, including a bitter taste, will turn litmus blue and neutralise acids.
- Acids and bases react to produce salts.

13.2.1 Acids

What can send a shiver down your spine, bring tears to your eyes and make your mouth pucker? Simply the taste of a freshly cut lemon or unripe pineapple. The vinegar in salad dressing, though less stringent, has the same effect. Even a cola or a glass of sparkling wine may leave a slightly sour taste in your mouth. All of these foods have a similar effect because of the presence of compounds called **acids**.

Hydrochloric acid is found in our stomachs, where it is used to help break down food. It is also used in industry, where it is sometimes called ‘spirit of salts’, to clean bricks and to clean off the coating of oxide on corroded iron or steel before plating the metal with a protective layer of zinc or tin. Table 13.1 lists common acids and their uses.

FIGURE 13.2 Citrus fruits contain citric acid. This gives lemons and limes their characteristic sour taste.

TABLE 13.1 Common acids and their uses

Acid	Formula	Occurrence or use
Nitric	HNO_3	Used to etch copper and manufacture fertilisers
Hydrochloric	HCl	Produced in the stomach; used to clean bricks and concrete
Sulfuric	H_2SO_4	Used in car batteries, plastics, insecticides, detergents and pharmaceuticals
Ethanoic	CH_3COOH	Found in vinegar (used in pickling and fermentation of foods)
Citric	$C_6H_8O_7$	Found in lemon juice
Carbonic	H_2CO_3	Found in soft drinks
Lactic	$C_3H_6O_3$	Makes milk sour; used in wrinkle-smoothing face cream
Ascorbic	$C_3H_8O_6$	Found in citrus fruits

Properties of acids

Acids have many common properties. They:

- usually taste sour
- are corrosive
- are molecular in structure and dissolve in water to produce an **electrolyte** (substance that conducts electricity)
- affect the colour of certain natural and synthetic dyes (they turn litmus, a plant dye, from blue to red)
- are neutralised by bases.

13.2.2 Bases

Bases are compounds with properties that in some ways complement those of acids. This means that the characteristic properties of an acid can be reduced by adding a base. If enough base is added, the acidic properties completely disappear. When this occurs, we say that the acid has been neutralised by the base. Bases taste bitter, have a soapy feel and are commonly found in cleaning agents. Soap, detergent, oven and bathroom cleaners, toothpaste and shampoo all contain water-soluble bases in varying degrees of strength. Table 13.2 lists some common bases and their uses.

TABLE 13.2 Common bases and their uses

Base	Formula	Use
Ammonia	NH_3	Fertilisers and detergents
Sodium hydroxide (caustic soda)	$NaOH$	Soaps and detergents
Sodium carbonate	Na_2CO_3	Manufacture of glass; washing powder and detergents
Calcium oxide (quicklime)	CaO	Bricklayers' mortar
Lead(II) oxide	PbO	House paint (now phased out)
Calcium hydroxide (slaked lime)	$Ca(OH)_2$	Garden lime, plaster and cement
Ammonium hydroxide	NH_4OH	Cleaning agents
Magnesium hydroxide	$Mg(OH)_2$	Milk of magnesia (for treatment of indigestion)

The properties of bases

Bases have many common properties. They:

- usually taste bitter
- feel slippery (they react with the natural oils in the skin and produce soap, which gives the characteristic 'slippery feel')
- turn litmus from red to blue
- are electrolytes (substances that conduct electricity)
- may be corrosive (for example, NaOH, which breaks down organic substances such as fat, hair and vegetable matter) and are good drain cleaners
- are generally ionic substances
- are oxides or hydroxides of metals (ammonium hydroxide is a base)
- are usually insoluble in water (a base that is soluble in water is called an **alkali**).

FIGURE 13.3 Caustic soda is another name for sodium hydroxide, a strong base.

13.2.3 Safety with acids and bases

Even so-called weak acids can be dangerous. A **concentrated** sample of ethanoic acid (acetic acid or vinegar), which is regarded as weak, can be extremely harmful to eyes and mucous membranes. Strong acids such as nitric acid, sulfuric acid and hydrochloric acid are extremely dangerous in concentrated form and should be handled with care even in their more **dilute** forms. When diluting a concentrated acid, always add acid to water. Adding water directly to concentrated acids may cause them to splatter, and severe burns may result. Strong bases such as sodium hydroxide and potassium hydroxide are also very dangerous, particularly in concentrated form and when hot. Acid and base spills on skin should be rinsed with large quantities of running water.

GARDENING WITH ACIDS AND BASES

Have you ever wondered why hydrangeas can bloom a different shade from year to year? The colour of the hydrangea bloom depends on the acidity of the soil. If the soil becomes too acidic, pink hydrangea blooms turn blue. To correct this, calcium oxide, CaO, generally known as 'lime' or 'quicklime' is added. Calcium oxide is a base that dissolves in water to form calcium hydroxide, $Ca(OH)_2$, or 'slaked lime'. This base changes the colour of the blossoms back to pink.

FIGURE 13.4 Blue hydrangeas grown in acidic oil and pink hydrangeas grown in more basic soil

13.2 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. List three examples of acids in the home.
2. Refer to table 13.1 and explain what type of bonding would be present in acids.
3. Refer to table 13.1 and name the element that is common to all acids.
4. Describe a simple method for determining whether an unknown liquid is acidic or basic.
5. State a use that is common to a number of bases and suggest why they are used in this way.
6. Explain the difference between a base and an alkali.
7. Which of the bases in table 13.2 would be soluble? State the dissolving process that occurs for each of them.
8. Acids and bases are electrolytes; explain what this means.
9. Describe one hazard when using acids and what protection is required.
10. Explain how acids should be diluted.

Fully worked solutions and sample responses are available in your digital formats.

13.3 Brønsted–Lowry theory of acids and bases

KEY CONCEPT

- The Brønsted–Lowry theory of acids and bases, including polyprotic acids and amphiprotic species, and writing of balanced ionic equations for their reactions with water including states

13.3.1 Defining acids and bases according to the Brønsted–Lowry theory

An acid–base theory suggested independently by Brønsted and Lowry in 1923 is still used today. This theory is summarised by the following two statements:

- An acid is a proton donor.
- A base is a proton acceptor.

The Brønsted–Lowry theory defitinion of acids and bases is as follows:

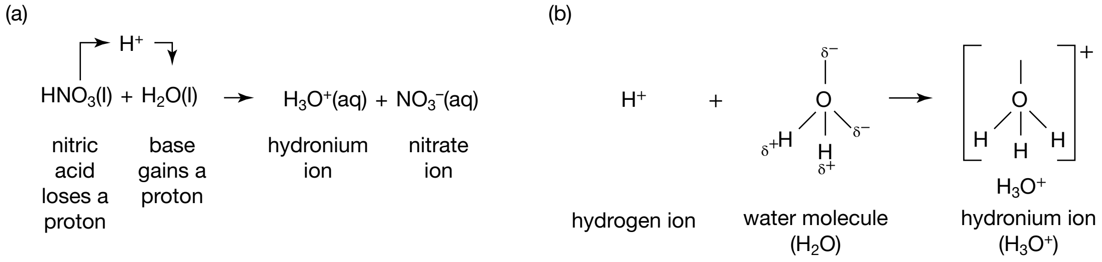
- **An acid is a substance that donates a proton (H^+ ion) to another substance.**
- **A base is a substance that accepts a proton (H^+ ion) from another substance.**

Since a hydrogen atom is simply a proton and an electron, removing the electron leaves a proton.

A proton is simply a hydrogen ion, H^+.

A reaction between an acid and a base involves a proton transfer from an acid to a base. In other words, the substance that loses a hydrogen ion, H^+, is the acid and the substance that accepts the hydrogen ion, H^+, is the base. This is illustrated in figure 13.5(a) where nitric acid, HNO_3, loses a proton to water, which created a **hydronium ion**, H_3O^+, seen in figure 13.5(b). According to the Brønsted–Lowry theory, the water in the equation shown in figure 13.5 is acting as a base because it accepts a proton.

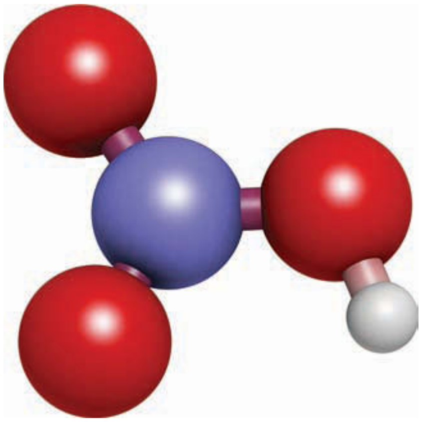
FIGURE 13.5 (a) Nitric acid, HNO_3, loses a proton to water, creating a hydronium ion, and (b) a water molecule accepts a proton to create a hydronium ion, H_3O^+.



The hydronium ion $H_3O^+(aq)$ is often shortened to $H^+(aq)$ but it is important to realise that a water molecule is usually attached to the proton.

Acids are proton donors.
Bases are proton acceptors.

FIGURE 13.6 A nitric acid molecule



on Resources

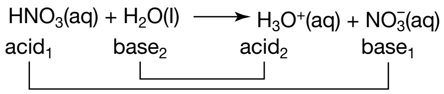
Video eLessons The Brønsted–Lowry theory (eles-2490)

13.3.2 Conjugate acid–base pairs

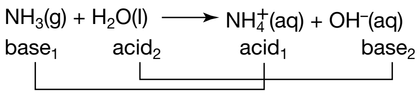
When chemists want to refer to particles that could be atoms, ions or molecules, the term that is used is **species**. The **conjugate base** of a Brønsted–Lowry acid is the species formed after the proton is donated. In the following example, the conjugate base of the acid HNO_3 is NO_3^-.

The **conjugate acid** of a base is the species formed after a base accepts a proton. The conjugate acid of the base H_2O is H_3O^+.

In this case, the conjugate pairs are (HNO_3/NO_3^-) and (H_3O^+/H_2O).



The conjugate pairs in this reaction are NH_4^+/NH_3 and H_2O/OH^-.



Conjugate acid–base pairs differ by a proton, H^+. To find the conjugate base of an acid, subtract one H^+; to find the conjugate acid of a base, add one H^+.

The convention is that the acid is always written first in the conjugate pairs.

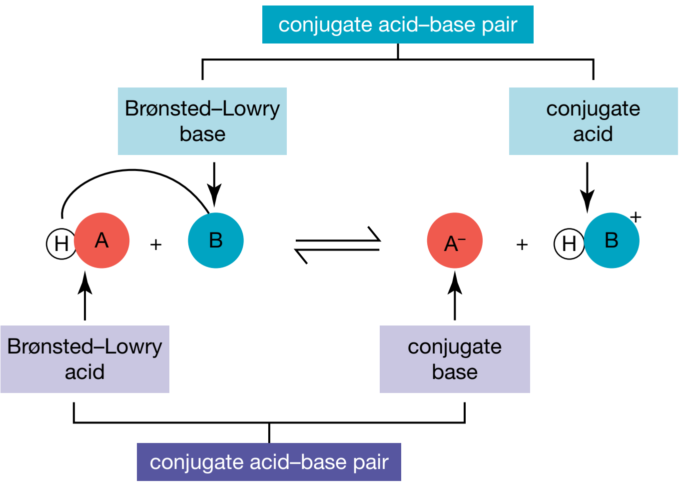
$$\text{conjugate acid} \underset{+H^+}{\overset{-H^+}{\rightleftharpoons}} \text{conjugate base}$$

For example:

conjugate acid		conjugate base
H_3PO_4	$\underset{+H^+}{\overset{-H^+}{\rightleftharpoons}}$	$H_2PO_4^-$

Note that a conjugate acid and a conjugate base can re-form the original acid and base by once again transferring a proton in a 'backwards' reaction according to our definitions.

FIGURE 13.7 Conjugate acids and bases



Conjugate base = acid – H^+
Conjugate acid = base + H^+

SAMPLE PROBLEM 1

Show that the following reaction is a proton transfer reaction and state the conjugate pairs.
$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$

THINK	WRITE
1. Decide whether the each of the reactants has donated or gained a proton.	HCl has become Cl^-; it has lost a proton (H^+). H_2O has become H_3O^+; it has gained a proton, (H^+).
2. The substance that has donated a proton is the acid.	HCl is an acid.
3. The substance that has gained a proton is the base.	H_2O is a base in this reaction.
4. Conjugate base = acid – H^+. Conjugate acid = base + H^+. **TIP:** Remember that the acid should be put first in each pair.	Conjugate base is Cl^-. Conjugate acid is H_3O^+. Conjugate pairs are HCl/Cl^- and H_3O^+/H_2O.

PRACTICE PROBLEM 1

Show that the following reaction is a proton transfer reaction and state the conjugate pairs:

$$NH_3(aq) + HNO_3(aq) \rightarrow NH_4^+(aq) + NO_3^-(aq)$$

13.3.3 Amphiprotic species

You may have noticed in the previous section that water is acting as a base in the first reaction and then acting as an acid in the second reaction. Some substances can act as acids or bases, according to their chemical environment. This means that they can donate or accept protons. Such substances are described as **amphiprotic**.

Water can act as an acid or base depending on whether it is reacting with a stronger acid or base. This is illustrated in the examples in the previous section. Other examples of amphiprotic substances are HSO_4^-, HCO_3^- and HS^-. These ions have both a proton that they can donate and a negative ion to attract an H^+.

Amphiprotic species have both a proton, H^+, that they can donate and a negative ion or pole to attract a proton, H^+.

An example of the hydrogen carbonate ion, HCO_3^-, acting as an amphiprotic species is as follows:

$$\underset{\text{acid}}{HCO_3^-(aq)} + \underset{\text{base}}{H_2O(l)} \rightarrow \underset{\text{acid}}{H_3O^+(aq)} + \underset{\text{base}}{CO_3^{2-}(aq)}$$

$$\underset{\text{base}}{HCO_3^-(aq)} + \underset{\text{acid}}{NH_4^+(aq)} \rightarrow \underset{\text{acid}}{H_2CO_3(aq)} + \underset{\text{base}}{NH(aq)}$$

13.3.4 Monoprotic, diprotic and triprotic acids

Acids can be classified as **monoprotic**, **diprotic** or **triprotic** depending on the actual number of protons that can be donated. A **polyprotic** acid can donate more than one proton.

Hydrochloric acid, HCl, is an example of a monoprotic acid because it can donate only one proton. Ethanoic acid, CH_3COOH, despite the fact that it has four hydrogen atoms, is also monoprotic because it can only donate the proton attached to the oxygen atom in the polar covalent bond.

FIGURE 13.8 Ethanoic acid, CH_3COOH, is a monoprotic acid.

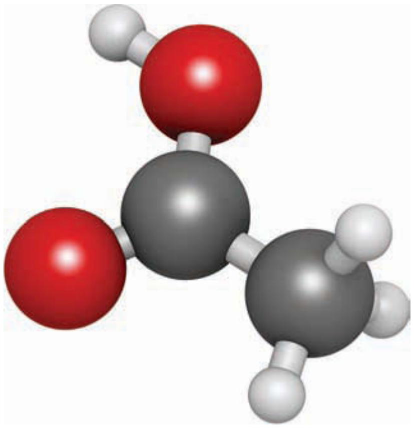
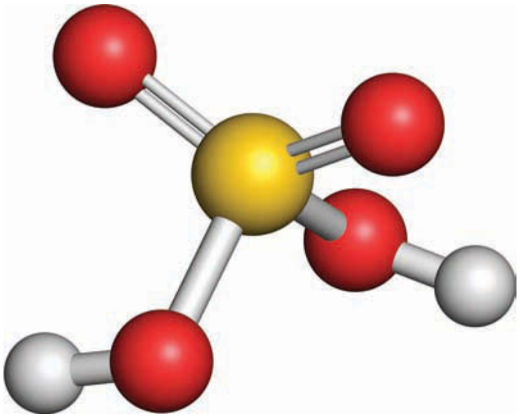


FIGURE 13.9 Sulfuric acid, H_2SO_4, can donate two protons so is a diprotic acid.



A polyprotic acid can donate more than one proton per molecule of acid. An acid such as sulfuric acid, H_2SO_4, is called a diprotic acid because it can donate two protons. This occurs in two stages, as shown in the following equations:

$$\text{Stage 1: } H_2SO_4(aq) + H_2O(l) \rightarrow H_3O^+(aq) + HSO_4^-(aq)$$

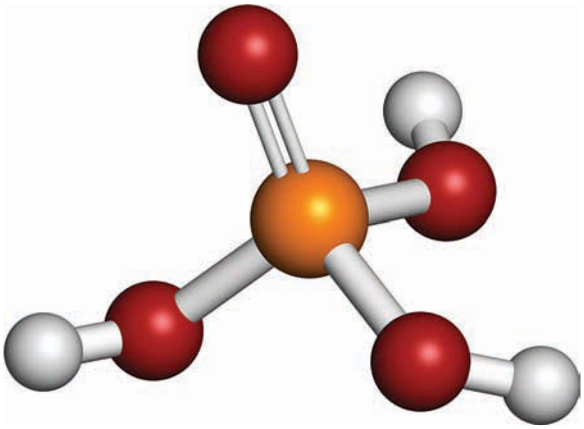
$$\text{Stage 2: } HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_4^{2-}(aq)$$

The first stage is more extensive than the second. The single arrow shows that virtually all of the H_2SO_4 molecules ionise. In stage two, the **double arrow** shows that while some HSO_4^- ions are reacting with water to form SO_4^{2-} some SO_4^{2-} ions are recombining with water to form HSO_4^- ions.

Phosphoric acid, H_3PO_4, is an example of a triprotic acid because it has three protons that it can donate.

Polyprotic acids can donate more than one proton.

FIGURE 13.10 Phosphoric acid, H_3PO_4, can donate three protons.



13.3 EXERCISE

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1. In the following reaction, which substance is acting as an acid and which is acting as a base?

$$H_2SO_4\,(l) + H_2O\,(l) \rightarrow H_3O^+\,(aq) + HSO_4^-\,(aq)$$

2. Identify the reactant that is acting as an acid in each of the following.
   - (a) $OH^-(aq) + HBr(aq) \rightarrow H_2O(l) + Br^-(aq)$
   - (b) $HSO_4^-(aq) + OH^-(aq) \rightarrow SO_4^{2-}(aq) + H_2O(l)$
   - (c) $HCO_3^-(aq) + HSO_4^-(aq) \rightarrow H_2O(l) + CO_2(g) + SO_4^{2-}(aq)$
   - (d) $H_3O^+(aq) + OH^-(aq) \rightarrow 2H_2O(l)$
3. Which of the following reactions can be classified as proton-transfer reactions?
   - (a) $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
   - (b) $H_3PO_4(aq) + H_2O(l) \rightarrow H_3O^+(aq) + H_2PO_4^-(aq)$
   - (c) $O^{2-}(aq) + H_2O(l) \rightarrow 2OH^-(aq)$
   - (d) $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
   - (e) $CO_3{}^{2-}(aq) + H_2O(l) \rightarrow HCO_3^-(aq) + OH^-(aq)$
4. Write the formulas of the conjugate bases of the following acids.
   - (a) H_2SO_4
   - (b) H_2S
   - (c) HS^-
   - (d) H_2O
   - (e) $NH_4{}^+$
5. Write the formulas of the conjugate acids of the following bases.
   - (a) OH^-
   - (b) HCO_3^-
   - (c) S^{2-}
   - (d) H_2O
   - (e) CN^-
6. Consider this equation: $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$. Which substances are acting as acids in this reaction?

7. Classify each of the following as a Brønsted–Lowry acid, base or amphiprotic substance.
   (a) Cl^-
   (b) H_2O
   (c) HSO_4^-
   (d) HNO_3
   (e) OH^-
8. (a) Write equations to illustrate H_3PO_4 acting as a triprotic acid in water.
   (b) Identify the conjugate acid–base pairs for the equations you wrote in part (a).
9. Write equations to illustrate the amphiprotic nature of the following species.
   (a) HS^-
   (b) HSO_4^-
10. Classify each of the following as monoprotic, diprotic, triprotic or amphiprotic.
   (a) H_3PO_4
   (b) CH_3COOH
   (c) HCO_3^{2-}
   (d) H_2CO_3
   (e) HNO_3
   (f) H_2O

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13.4 Concentrating on acids and bases

KEY CONCEPT

- The distinction between strong and weak acids and bases, and between concentrated and dilute acids and bases, including common examples

13.4.1 Calculating concentration using molarity

When working with solutions, the **concentration** can have implications on how they react. Concentration refers to the quantity of solute in a particular volume of solution. Concentration in chemistry is usually measured in units of moles of solute per litre (mol L^{-1} or M) of solution. This is known as the **molarity** of a solution. You may already have noticed this on bottles of chemicals that you have used. The molarity of solutions is one of the most commonly used measures in chemistry. For example, a 3.0 M HCl solution contains 3.0 moles of HCl dissolved in 1 litre of solution, and a 0.5 M NaOH solution has 0.5 mole of sodium hydroxide dissolved in 1 litre of solution.

Concentration is number of moles of solute per litre (mol L^{-1}); also known as molarity (M).

$$c = \frac{n}{V}$$

Alternatively, the moles of solute in a given volume of solution of a known concentration:

$$n = c \times V$$

where:
n is the number of moles of solute
c is the concentration or molarity (mol L^{-1} or M)
V is the volume (L).

SAMPLE PROBLEM 2

Calculate the number of mole present in 250 mL of 3.00 M HCl.

THINK	WRITE
1. Identify the given information and compare units given to units required. The volume of the solution is given in mL, while the concentration is calculated in L. Convert from mL to L.	$c(\text{HCl} = 3.00\,\text{M}$ $V(\text{HCl}) = \dfrac{250\,\text{mL}}{1000}$ $= 0.250\,\text{L}$
Identify the unknown quantity.	$n(\text{HCl} = ?$
2. Use the formula to calculate the number of moles: $n = c \times V$ **TIP:** Remember to check significant figures and units.	$n(\text{HCl}) = 3.00 \times 0.250$ $= 0.750\,\text{M}$

PRACTICE PROBLEM 2

Calculate the number of moles present in 100 mL of 1.35 M NaOH.

If you are required to make up a solution of a particular concentration, you will need to start with a mass of the substance because it is not possible to measure out moles directly. In this case, you need to use the formula introduced in topic 5 to calculate n:

$$n = \frac{m}{M}$$

where m is the mass in grams and M is the molar mass in g mol^{-1}. These calculations will be explored in topic 16.

TIP: Remember that uppercase M represents concentration of a solution in mol L^{-1}. Uppercase *M* in italics is molar mass in g mol^{-1}.

Concentrated and dilute solutions

A concentrated acid has a large amount of acid dissolved in a given volume, and a dilute acid has a small amount of acid dissolved in a given volume. It is possible to have a concentrated strong acid or strong base as well as a concentrated weak acid or weak base. The concentration of a solution can be changed by adding water; however the strength cannot be changed.

Concentrated solutions have a large number of solute particles in a given volume.
Dilute solutions have a small number of solute particles in a given volume.

13.4.2 Strength of acids and bases

The strength of an acid or base is related to the ease with which it donates or accepts a proton.

Strong acids and bases

A **strong acid** donates protons readily.
A **strong base** accepts protons readily.

Common strong laboratory acids include hydrochloric acid, HNO_3, and sulfuric acid, H_2SO_4. There are only a few strong acids and they fully **ionise** in water, producing many ions. Their solutions, therefore, are strong electrolytes. All protons are donated to water, forming hydronium ions. Hydrochloric acid fully reacts with water to produce a hydronium ion and a chloride ion.

$$\underset{\text{strong acid}}{\mathrm{HCl\,(g)}} + \mathrm{H_2O\,(l)} \xrightarrow{\text{virtually complete}} \mathrm{H_3O^+(aq)} + \mathrm{Cl^-(aq)}$$

Some strong bases include lithium hydroxide, LiOH, sodium hydroxide, NaOH, and potassium hydroxide, KOH. Strong bases are generally hydroxides of metals found in group 1 of the periodic table. Strong bases fully **dissociate** in water producing many ions. Their solutions, therefore, are also strong electrolytes. Sodium hydroxide ionises in water to form sodium ions and hydroxide ions.

$$\mathrm{NaOH(s)} \xrightarrow{\mathrm{H_2O}} \mathrm{Na^+(aq)} + \mathrm{OH^-(aq)}$$

FIGURE 13.11 Hydrochloric acid must be handled with care.

FIGURE 13.12 Strong electrolytes show high conductivity because they produce many ions in solution; water is classified as a non-conductor because it very few ions.

Strong acids completely ionise in water.
Strong bases completely dissociate in water.

Resources

Weblink Acid strength

Weak acids and bases

Examples of **weak acids** are ethanoic acid, lactic acid and citric acid. Weak acids only partially ionise in water, resulting in a solution that is a poor conductor of electricity because not many ions are formed. Ethanoic acid partially reacts with water to form hydronium ions and ethanoate ions. Double arrows are used to show that these are partial reactions.

$$\mathrm{CH_3COOH(l)} + \mathrm{H_2O(l)} \rightleftharpoons \mathrm{H_3O^+(aq)} + \mathrm{CH_3COO^-(aq)}$$

However, a **weak base** such as ammonia does not readily ionise and accept protons. It forms very few ions in solution. Ammonia is a poor conductor of electricity.

Weak acids partially ionise in water.
Weak bases partially dissociate in water.

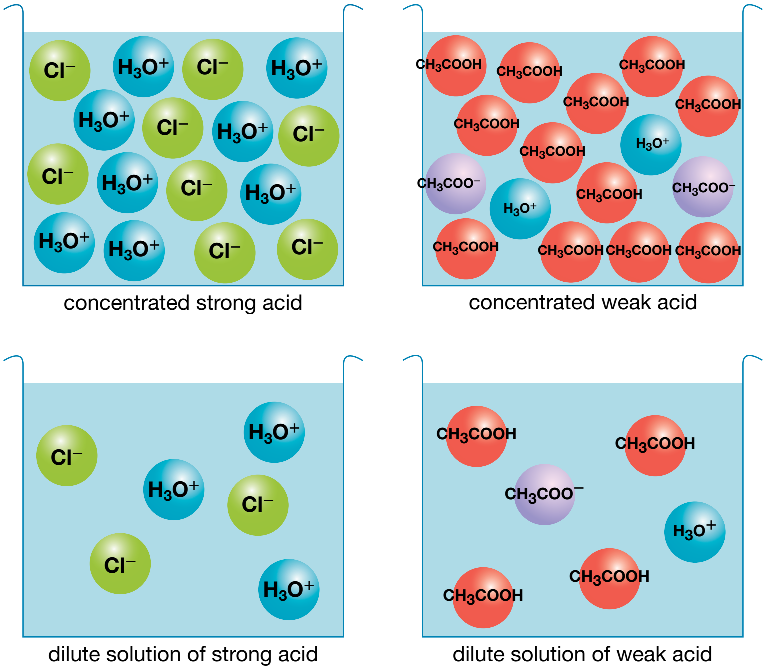
TABLE 13.3 Relative strengths of common acids and their conjugate bases

	Name of acid	Formula	Conjugate base	Name of base	
very strong	Hydrochloric	HCl	Cl^-	Chloride ion	very weak
	Nitric	HNO_3	NO_3^-	Nitrate ion	
	Sulfuric	H_2SO_4	HSO_4^-	Hydrogen sulfite ion	
strong	Hydronium ion	H_3O^+	H_2O	Water	weak
	Phosphoric	H_3PO_4	$H_2PO_4^-$	Dihydrogen phosphate ion	
	Hydrofluoric	HF	F^-	Fluoride ion	
	Ethanoic (acetic)	CH_3COOH	CH_3COO^-	Ethanoate (acetate) ion	
decreasing strength of acids	Carbonic	H_2CO_3	HCO_3^-	Hydrogen carbonate ion	increasing strength of bases
	Hydrogen sulfide	H_2S	HS^-	Hydrogen sulfide ion	
	Ammonium ion	NH_4^+	NH_3	Ammonia	
	Hydrogen carbonate ion	HCO_3^-	CO_3^{2-}	Carbonate ion	
weak	Hydrogen sulfide ion	HS^-	S^{2-}	Sulfide ion	strong
	Water	H_2O	OH^-	Hydroxide ion	
very weak	Hydroxide ion	OH^-	O^{2-}	Oxide ion	very strong
	Hydrogen	H_2	H^-	Hydride ion	

13.4.3 Concentration versus strength

In summary, the concentration of an acid or base refers to the quantity of solute in a given volume of solution, whereas the strength of a solution refers to the extent of ionisation or dissociation of the acid or base. A strong acid fully ionises in solution and a strong base fully dissociates in solution.

FIGURE 13.13 Concentrated and dilute weak and strong acids



concentrated strong acid

concentrated weak acid

dilute solution of strong acid

dilute solution of weak acid

13.4 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Calculate the number of mole present in the following.
   (a) 0.100 L of 1.30 M NaOH.
   (b) 120 mL of 2.00 M HCl
   (c) 50.0 mL of 0.950 M NHO_3
2. If a 0.0850 litre solution contains 0.0750 moles of H_2SO_4, what is the molarity?
3. Calculate the volume, in L, of each of the following solutions that contain 2.50 moles of the substance.
   (a) 3.50M NaOH
   (b) 0.500 M KOH
   (c) 0.240 M CH_3COOH
4. Sea water contains about 35.0 g of the neutral salt, NaCl, per litre. What is the molarity of sodium chloride in sea water?
5. How many mole of H_3O^+ ions are present in 10.5 mL of 0.180 M HCl?
6. The acids found in fruit juice are weak acids. Explain what this means.
7. Using HBr and HS^- as examples, write equations, with appropriate arrows, to illustrate the difference between a strong acid, HBr, and a weak acid, HS^-. Explain the difference in strength in terms of the conductivity of the resulting solution.
8. Dimethylamine, $HN(CH_3)_2$, is a weak base that reacts in water like ammonia, NH_3. Write an equation to illustrate the reaction of dimethylamine with water.
9. Acids, such as nitric and hydrochloric, are used by artists to make engravings on metal plates made from copper, zinc or iron.
   (a) Use table 13.3 to find the conjugate pairs of the acids used by the artists for their etchings.
   (b) What safety precautions would you recommend an artist working with strong acids should take?

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13.5 pH and acidity

KEY CONCEPT

- The ionic product of water, the pH scale and the use of pH in the measurement and calculations of strengths of acids and bases and dilutions of solutions (calculations involving acidity constants are not required).

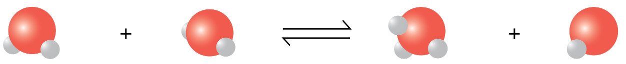
13.5.1 Ionic product of water

Because water can act as both an acid and a base, it can also react with itself according to the equation shown in figure 13.14.

Careful measurements have shown that this does indeed occur and that the concentration of these ions in neutral water is 10^{-7} M (at 25 °C). Such values indicate that this reaction occurs only to an extremely small extent.

FIGURE 13.14 Self-ionisation of water

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$



Furthermore, measurements have shown that, even when dissolved acids and bases are present, the mathematical product of these two ions is always 10^{-14}. If the concentration of H_3O^+ ions increases due to the presence of a dissolved acid, the concentration of OH^- ions decreases such that the product is still 10^{-14}. The reverse applies if a source of OH^- ions (a dissolved base) is present.

For any acidic, basic or neutral solution, use square brackets to show concentration in mol L^{-1}.

K_w is termed the **self-ionisation constant** for water, and $[H_3O^+][OH^-]$ is called the ionic product.

It follows that:

- if $[H_3O^+] > [OH^-]$, the solution is acidic
- if $[OH^-] > [H_3O^+]$, the solution is basic
- if $[H_3O^+] = [OH^-]$, the solution is neutral.

However, in all cases, the product $[H_3O^+][OH^-]$ equals 10^{-14}.

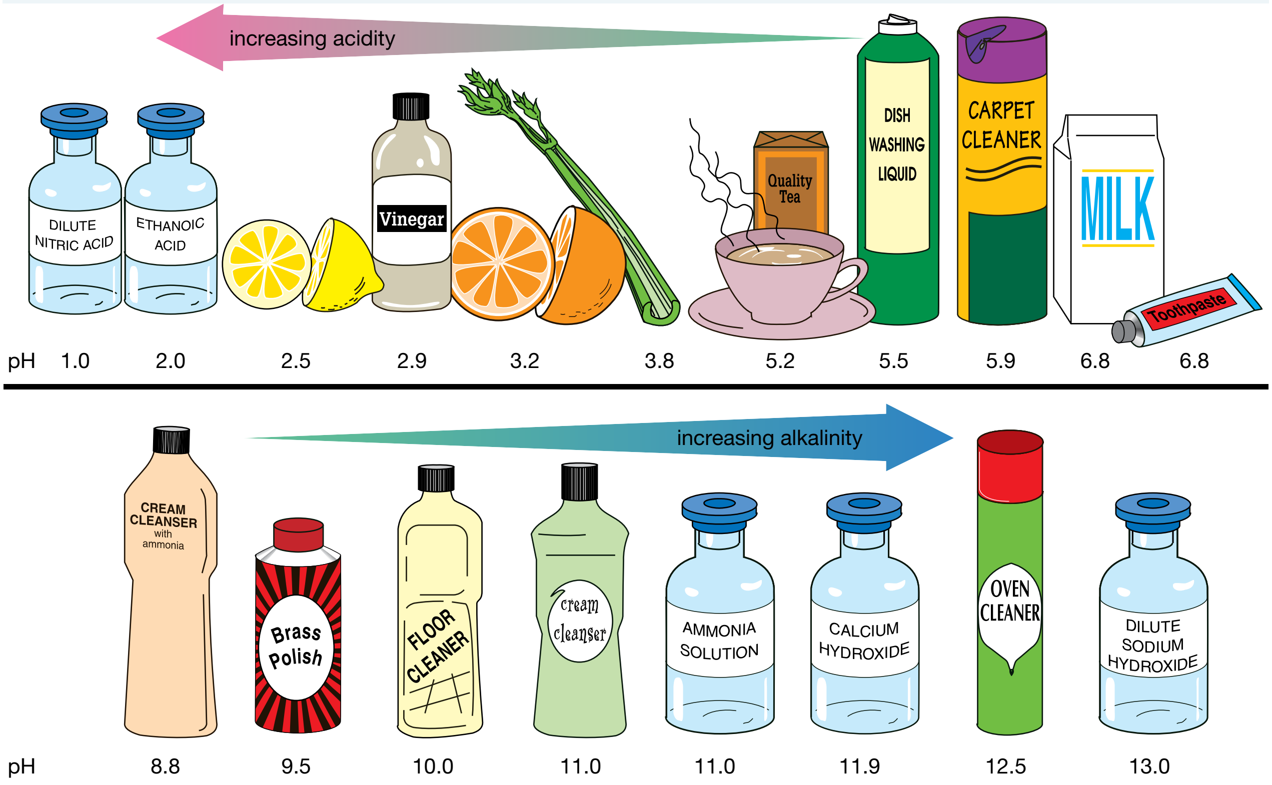
This information is useful when calculating the pH of basic solutions.

For all aqueous solutions $K_w = [H_3O^+][OH^-] = 10^{-14}$ M^2 at 25 °C.

13.5.2 The pH scale

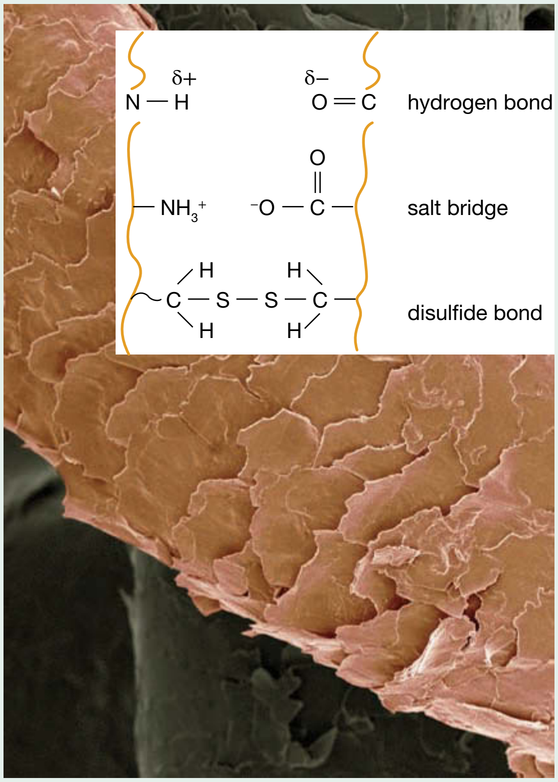
The **pH scale** is used to measure the acidity of a solution. It is usually applied over a range from 1 to 14. Using this scale, a neutral solution has a pH of 7. Values lower than 7 indicate an acidic solution. The stronger the acid, the lower the pH value. For example, a solution of pH 3 is stronger than a solution of pH 4. Any pH value greater than 7 indicates a basic solution. In this case, the higher the value, the stronger the base. It is interesting to note that pH does not have a unit because it is just a number on a scale and that the temperature is specified because the pH can change slightly if the temperature changes. The pH scale of some common acids and bases is shown in figure 13.15.

FIGURE 13.15 The pH of common acids and bases



THE CHEMISTRY OF HAIR

Hair is made from amino acids that polymerise into long chains held together by hydrogen bonding, salt bridges between the acid and base groups, and disulfide bonds. Hair is normally acidic (pH 4 to 5), and for this reason alkaline shampoos, which clean the hair, are followed by an 'acidic rinse' or conditioner. However, both strongly acidic solutions and slightly alkaline solutions can damage the hair. Strongly acidic solutions (pH 1 to 2) break the H–bonding and salt bridges of the hair. Slightly alkaline (pH 8.5) solutions break some of the disulfide bonds, causing the outer surface of hair strands to become ruffled and appear dull, as light is unevenly reflected from their surfaces.



on Resources

Digital document Experiment 13.1 Indicators and pH (doc-30861)

13.5.3 Calculating pH of strong acids

If we know the concentration of the hydronium ion in moles per litre of solution, we can calculate the pH using the following formula.

$$\text{pH} = -\log_{10}\left[\text{H}_3\text{O}^+\right]$$

This means that pH is the negative logarithm of the molar hydronium concentration. This formula enables pH to be calculated for any known H_3O^+ or OH^- concentration by using a calculator with a log function. A **logarithm** is a quantity representing the power to which a fixed number (the base) must be raised to produce a given number. In this case the base is 10, so the logarithm is the power of 10. So, this formula can be rearranged to find the concentration of acid when given the pH:

$$\left[\text{H}_3\text{O}^+\right] = 10^{-\text{pH}}$$

If the $[H_3O^+]$ is easily converted into a power of 10, a calculator is not needed. This formula can only be used for strong acids because strong acids are completely ionised. So for a strong acid, HA:

$$\left[\text{H}_3\text{O}^+\right] = [\text{HA}]$$

This is not true for weak acids because if the acid (HA) is weak, it is only partially ionised and the $[H_3O^+] << [HA]$ and cannot be calculated using this formula.

To calculate the pH of strong acids, HA:

$$[\mathrm{HA}] = \left[\mathrm{H_3O^+}\right] = 10^{-\mathrm{pH}}$$

$$\mathrm{pH} = -\log_{10}\left[\mathrm{H_3O^+}\right].$$

SAMPLE PROBLEM 3

Calculate the pH of a 0.0001 M solution of HCl at 25 °C.

Teacher-led video: SP3 (tlvd-0565)

THINK	WRITE
1. Identify the given information and the unknown information.	$[\mathrm{HCl}] = 0.0001\ \mathrm{M}$ $\mathrm{pH} = ?$
2. HCl is a strong acid and so is fully ionised. This means the concentration of $H_3O^+ = [\mathrm{HCl}]$	$[\mathrm{H_3O^+}] = [\mathrm{HCl}]$ $= 0.0001\ \mathrm{M}$
3. Convert this concentration to a power of 10.	$\left[\mathrm{H_3O^+}\right] = 10^{-4}\ \mathrm{M}$
4. Find the pH using $\mathrm{pH} = -\log_{10}[\mathrm{H_3O^+}]$.	$\mathrm{pH} = -\log_{10}\left(10^{-4}\right)$ $\mathrm{pH} = -(\text{power of } 10)$ $= -(-4)$ $= 4$

PRACTICE PROBLEM 3

A solution is found to have a H_3O^+ concentration of 0.01 M at 25 °C. Calculate its pH.

If the $[\mathrm{H_3O^+}]$ is not a number that can easily be converted into a power of 10, the procedure is simple using a calculator.

SAMPLE PROBLEM 4

Calculate the pH of 0.054 M solution of HCl at 25 °C.

Teacher-led video: SP4 (tlvd-0566)

THINK	WRITE
1. Identify the given information and the unknown information.	$[\mathrm{HCl}] = 0.054\ \mathrm{M}$ $\mathrm{pH} = ?$
2. HCl is a strong acid and so is fully ionised. This means the concentration of $H_3O^+ = [\mathrm{HCl}]$.	$[\mathrm{H_3O^+}] = [\mathrm{HCl}]$ $[\mathrm{H_3O^+}] = 0.054\ \mathrm{M}$

3. Find the pH using $pH = -\log_{10}[H_3O^+]$. This time a calculator is needed:
    - Press the negative key
    - Press the log key
    - Key in the number
    - Press enter

   TIP: Remember to give your answer to the correct number of significant figures.

$pH = -\log_{10}(0.054)$

$pH = 1.3$

PRACTICE PROBLEM 4

Calculate the pH of 0.037 M solution of HNO_3 at 25 °C.

If you need to find the concentration of $[H_3O^+]$, the pH formula must be rearranged.

SAMPLE PROBLEM 5

Calculate the concentration of a solution of HCl with a pH of 2.1 at 25 °C.

THINK	WRITE
1. Identify the given information and the unknown information.	$pH = 2.1$ $[HCl] = ?$
2. HCl is a strong acid and so is fully ionised. This means $[HCl] = [H_3O^+]$. Find $[H_3O^+]$ using $[H_3O^+] = 10^{-pH}$.	$\left[H_3O^+\right] = 10^{-2.1}$ M
3. Calculate using a calculator: • Press 10 • Press ˆ • Press −2.1 • Press enter	$\left[H_3O^+\right] = 7.9 \times 10^{-3}$ M
4. Since HCl is a strong acid $[HCl] = [H_3O^+]$	$[HCl] = [H_3O^+] = 7.9 \times 10^{-3}$ M

PRACTICE PROBLEM 5

Calculate the concentration of a solution of HNO_3 with a pH of 1.8 at 25 °C.

13.5.4 Relative strengths of acids and bases and pH

The strength of an acid or base is measured by its tendency to donate (in the case of acids) or accept (in the case of bases) protons. The strength of an acid has to be taken into consideration when calculating the pH, because it determines the acid's ability to ionise.

It has been found experimentally that not all acids of 0.1 M concentration have the same pH, and do not conduct electricity to the same extent. Remember that the conductivity of a solution is directly related to the number of ions in solution. Consider two acids, HCl and CH_3COOH, with the same concentration of 0.1 M at 25 °C. Experimental results confirm that hydrochloric acid has a pH of 1; however, ethanoic acid has a pH of 3. This is explained by the differing abilities of the two acids to ionise. Almost every HCl molecule donates protons to water. This is consistent with the fact that HCl is a strong acid and ionises virtually completely to produce a good electrolyte. CH_3COOH, on the other hand, is a weak acid. Only one in every hundred CH_3COOH molecules loses a proton, and therefore a solution of ethanoic acid is a poor electrolyte and hence a poor conductor of electricity. This variability is illustrated in table 13.4.

TABLE 13.4 pH values for solutions of a range of acids

Acid	Formula	pH (0.01 M solution)	pH (0.001 M solution)
Benzoic acid	C_6H5COOH	3.1	3.6
Ethanoic acid	CH_3COOH	3.4	3.9
Hypochlorous acid	HOCl	4.8	5.3
Lactic acid	HC_3H5O_3	2.9	3.4
Methanoic acid	HCOOH	2.9	3.4
Nitrous acid	HNO_2	2.6	3.1

The pH of a basic solution is also affected by the ability of the base to dissolve or dissociate in water. For example, a 0.1 M solution of NaOH, a strong base, has a higher pH than a 0.1 M solution of NH_3, a weak base. The sodium hydroxide solution is, therefore ,a stronger base than the ammonia solution.

This dependence of pH on acid strength can be used to compare the relative strengths of a range of acids. Provided that the concentrations of the acids are all the same, the lower the pH the stronger the acid is (refer to table 13.4). Note that nitrous acid is the strongest acid of those listed (lowest pH) while hypochlorous acid (highest pH) is the weakest. It does not matter which concentration is used to make this comparison, as long as it is the same for all the acids that you are comparing.

13.5.5 Calculating pH of strong bases

The definition of pH is based on the concentration of the hydronium ion H_3O^+, so to find the pH of a basic solution it is necessary to use the expression given for the self-ionisation of water, remembering that in aqueous solutions $K_w = [H_3O^+][OH^-] = 10^{-14}\ M^2$.

Therefore, by rearranging this formula, the concentration of hydronium ion can be found and the pH can then be found using a calculator as shown in sample problem 5.

To calculate the pH of strong bases:

$$\begin{aligned}[H_3O^+] &= \frac{10^{-14}}{[OH^-]} \\ pH &= -\log_{10}[H_3O^+].\end{aligned}$$

SAMPLE PROBLEM 6

Calculate the pH of a solution of 1.5 M NaOH solution at 25 °C.

Teacher-led video: SP6 (tlvd-0568)

THINK	WRITE
1. Identify the given information and the unknown information.	[NaOH] = 1.5 M pH = ?

2. $[OH^-] = [NaOH]$ because NaOH is a strong base.	$[OH^-] = [NaOH]$ $[OH^-] = 1.5\ M$
3. To find $[H_3O^+]$ use the formula from the self-ionisation of water: $\left[H_3O^+\right] = \frac{10^{-14}}{[OH^-]}$	$[H_3O^+] = \frac{10^{-14}\ M}{1.5}$
4. To find $[H_3O^+]$ use a calculator: • Press 10 • Press ^ • Press −14 • Press enter • Press ÷ 1.5	$\left[H_3O^+\right] = \frac{10^{-14}\ M}{1.5}$ $\left[H_3O^+\right] = 6.7 \times 10^{-15}\ M$
5. Now to find the pH, apply the formula: $pH = -\log_{10}[H_3O^+]$ And use a calculator: • Press −log (2^{nd} ans) • Press enter	$pH = -\log_{10}[6.7 \times 10^{-15}]$
6. Check that the pH is greater than 7 (because NaOH is a base).	$pH = 14.2$

PRACTICE PROBLEM 6

Calculate the pH of a solution of 2.4×10^{-2} M NaOH solution at 25 °C.

If you are required to find the concentration of a base given the pH, the method is shown in sample problem 7.

SAMPLE PROBLEM 7

Calculate the concentration of a solution of NaOH with a pH of 12.5 at 25 °C.

Teacher-led video: SP7 (tlvd-0569)

THINK	WRITE
1. Identify the given information and the unknown information.	$pH = 12.5$ $[NaOH] = ?$
2. Recall the formula $[H_3O^+] = 10^{-pH}$	$\left[H_3O^+\right] = 10^{-12.5}\ M$
3. To find $[OH^-]$ use the formula from the self-ionisation of water: $\left[H_3O^+\right] = \frac{10^{-14}}{[OH^-]}$	$[OH^-] = \frac{10^{-14}}{10^{-12.5}}$ $= 10^{-1.5}\ M$
4. Use a calculator: • Press 10 • Press ^ • Press −1.5 • Press enter	$[OH^-] = 0.032\ M$
5. $[NaOH] = [OH^-]$ because NaOH is a strong base.	$[NaOH] = 0.032\ M$

PRACTICE PROBLEM 7

Calculate the concentration of a solution of KOH with a pH of 11.8 at 25 °C.

CONTROLLING THE pH OF SWIMMING POOLS

The growth of algae and bacteria is controlled in swimming pools through the use of chlorine compounds. Sodium hypochlorite, NaOCl, and calcium hypochlorite, $Ca(OCl)_2$, are used, undergoing hydrolysis to produce the weak acid hypochlorous acid, HOCl, according to the equation:

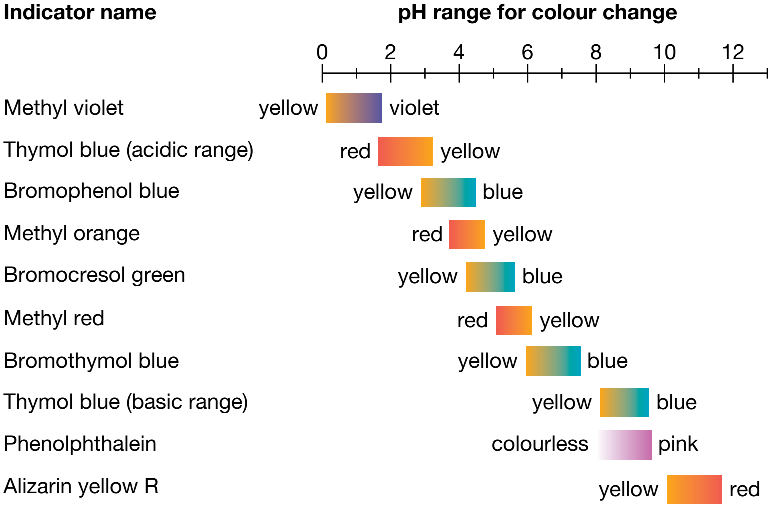
$$OCl^- (aq) + H_2O (l) \rightarrow HOCl (aq) + OH^- (aq)$$

However, if the pH is too high, the concentration and effectiveness of HOCl are reduced. To lower the pH, an acid solution (such as hydrochloric acid or solid sodium hydrogen sulfate) can be used to neutralise the excess hydroxide ions. If the pH is too low, however, the acid content of the water can cause eye and skin irritation and damage to pool fittings. The pH can be raised by neutralising with sodium carbonate or sodium hydrogen carbonate.

13.5.6 Measuring pH

Measuring pH can be performed using indicators or a pH meter.

FIGURE 13.16 Colour changes for different indicators across a range of pH



Indicators

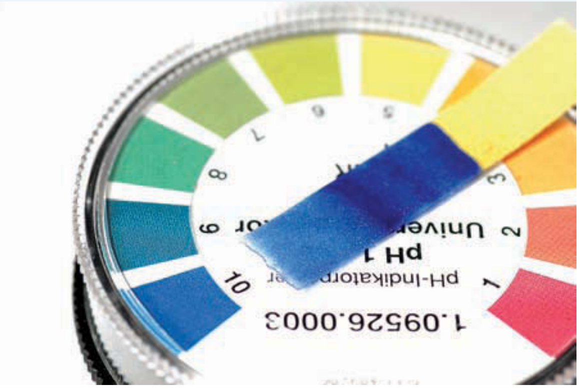
Indicators are dyes that are themselves weak acids or bases. They have the important characteristic that they show a different colour in an acidic solution from a basic solution, because their conjugate base has a different colour from their conjugate acid. They can be used to determine whether a substance is acidic or basic and also how acidic or basic it is. Each indicator changes colour at a certain pH so it is important to choose an indicator that is appropriate for a particular reaction.

Litmus, phenolphthalein, methyl orange and universal indicator are the most commonly used indicators. Universal indicator, a mixture of several indicators, changes colour gradually from red to green to violet as the solution changes from acidic to neutral to basic.

The pH of a solution can be measured by adding universal indicator and matching the colour shade obtained against a special chart (see figure 13.17).

Indicators are useful tools, although they have certain limitations. The indicator colours found in charts are quoted for 25 °C; at other temperatures an indicator may change colour at a different pH. If a solution being tested has a colour of its own, the colour of the indicator may be distorted. As the naked eye is used to make the comparison between the indicator colour and the pH chart, this can be a problem, especially for people who are colourblind. More accurate readings can be obtained through the use of specially designed pH meters.

FIGURE 13.17 Universal indicator paper



The pH meter

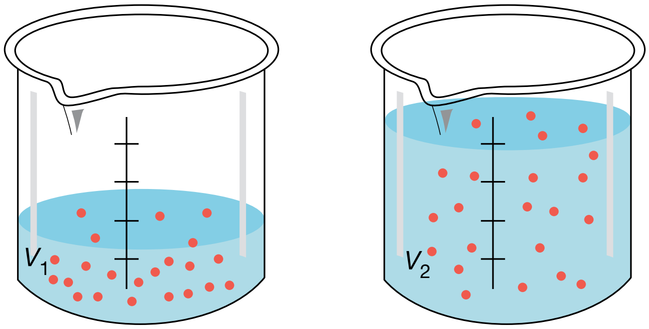
Most laboratories in industry have a pH meter that is used to make rapid, accurate measurements of pH. It can be connected to a computer to monitor pH changes continuously.

The voltage of the electrode changes with the [H_3O^+] in the solution into which it is dipped. Values of pH obtained by a pH meter are accurate to within 0.01 units of the true pH and are not affected by the colour and cloudiness of the unknown solution. Hospitals use pH meters to find small but significant changes of pH in blood and other body fluids. Soils, ponds, rivers, sewage and industrial effluents are also easily monitored with a pH meter.

13.5.7 Dilution of solutions

Chemicals are often bought and stored in concentrated form known as **stock solutions**. They are then diluted to the concentration required. Dilution is the process of adding more water to a solution. In this process, the volume increases and the concentration decreases but the number of mole has not changed because no solute has been added or removed. When using this process, it is essential to be able to accurately calculate the concentration of the diluted solution or the volume of water required to be added to produce a desired concentration.

FIGURE 13.18 Before and after a dilution; the volume changes but the number of moles remains the same.



Consider a beaker with volume V_1 and concentration c_1. The number of moles present is $n = c_1V_1$. If water is added to this beaker, there is a new volume V_2 and a different concentration c_2; however, the moles present are the same, $n = c_2V_2$ (see figure 13.18).

When a solution is diluted, the new volume (V) or concentration (c) can be determined by:

$$c_1V_1 = c_2V_2$$

The units of volume can be either mL or L but both V_1 and V_2 must be the same unit.

SAMPLE PROBLEM 8

If 50.0 mL of a 2.50 M solution is diluted to 85.0 mL, what is the final concentration?

Teacher-led video: SP8 (tlvd-0570)

THINK	WRITE
1. Identify the given information and the unknown information.	$c_1 = 2.50\ \text{M}$ $V_1 = 50.0\ \text{mL}$ $V_2 = 85.0\ \text{mL}$ $c_2 = ?$
2. Rearrange the dilution formula, $c_1V_1 = c_2V_2$, to find c_2. **TIP:** Remember to check significant figures and units and that the concentration is lower after dilution.	$c_2 = \frac{c_1V_1}{V_2}$ $c_2 = \frac{2.50 \times 50.0}{85.0}$ $= 1.47\ \text{M}$

PRACTICE PROBLEM 8

If 125.0 mL of a 1.80 M solution is diluted to 220 mL, what is the final concentration?

To calculate the water added to achieve a new concentration, remember to subtract the initial volume.
Water added $= V_2 - V_1$

SAMPLE PROBLEM 9

How much water would be added to 500 mL of a 2.20 M HCl solution to make a 1.50 M solution?

Teacher-led video: SP9 (tlvd-0571)

THINK	WRITE
1. Identify the given information and the unknown information.	$c_1 = 2.20\ \text{M}$ $V_1 = 500.0\ \text{mL}$ $c_2 = 1.50\ \text{M}$ $V_2 = ?$ Water added = ?
2. Rearrange the dilution formula, $c_1V_1 = c_2V_2$, to find V_2.	$V_2 = \frac{c_1V_1}{c_2}$ $V_2 = \frac{2.20 \times 500}{1.50}$ $= 733\ \text{mL}$
3. Determine amount of water added. Remember to check for significant figures and units.	$V_2 - V_1 = 733 - 500$ $= 233\ \text{mL}$

PRACTICE PROBLEM 9

How much water would need to be added to 750 mL of a 3.20 M HCl solution to make a 2.00 M solution?

If pH values are provided instead of concentrations, it is necessary to convert the pH into a concentration of acid or base (as shown in section 13.5.3) before proceeding with the dilution formula.

For solutions dilutions that involve acidic or basic solutions, this affects the concentration of H_3O^+ ions that are present and, hence, the pH. Because pH is on a logarithmic scale, the factor by which H_3O^+ changes is not the same as that by which pH changes. A change in pH of 1 means a change in concentration of a factor of 10.

This means that a solution of pH 3 is 10 times more dilute than a solution of pH of 2.

SAMPLE PROBLEM 10

60.0 mL of a solution of hydrochloric acid of concentration 0.0550 M is diluted by adding 40.0 mL of water. Calculate the pH of the resulting solution.

Teacher-led video: SP10 (tlvd-0572)

THINK	**WRITE**
1. Identify the given information and the unknown information.	$c_1 = 0.0550\,\text{M}$ $V_1 = 60.0\,\text{mL}$ $V_2 = 100.0\,\text{mL}$ $\text{pH} = ?$
2. To determine the pH, the concentration must first be determined by applying the dilution formula: $c_IV_1 = c_2V_2$	$c_1V = c_2V$ $c_2 = \dfrac{c_1V_1}{V_2}$ $= \dfrac{0.0550 \times 60.0}{100}$ $= 0.0330\,\text{M}$ $c_2 = [\text{HCl}] = 0.0330\,\text{M}$
3. $[H_3O^+] = [\text{HCl}]$ because HCl is a strong acid. Use the formula $\text{pH} = -\log_{10}[H_3O^+]$ to find pH of the diluted solution.	$\text{pH} = -\log_{10}\left[H_3O^+\right]$ $= -\log_{10}(0.033)$ $= 1.48$

PRACTICE PROBLEM 10

20.0 mL of a solution of hydrochloric acid of concentration 0.250 M is diluted by adding 40.0 mL of water. Calculate the pH of the resulting solution.

on Resources

Digital documents	Experiment 13.2 Finding the pH of common household substances (doc-30855)
	Experiment 13.3 The reactions of acids (doc-30860)
Teacher-led video	Experiment 13.3 The reactions of acids (tlvd-0632)

13.5 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. (a) Refer to the ionic product of water and determine [OH^-] if the concentration of [H_3O^+] = 10^{-5} M.
   (b) Is this solution acidic or basic?
2. What is the pH scale and how can pH be calculated?
3. What is the difference between a dilute and a concentrated solution of an acid or base?
4. Determine the pH of the following given that [H_3O^+] is:
   (a) 10^{-5} M.
   (b) 0.1 M.
5. Determine the pH of the strong acid $HClO_4$, which has a concentration of 0.0015 M.
6. Determine the pH of the following given that [OH^-] is:
   (a) 10^{-4} M.
   (b) 0.1 M.
7. 90.0 mL of water is added to 10.0 mL of 0.650 M HNO_3. What is the new pH?
8. Determine the pH of a solution where:
   (a) [H_3O^+] = 9.2×10^{-4} M.
   (b) [OH^-] = 1.8×10^{-5} M.
9. Using the formula for pH, explain why a neutral solution has a pH of 7 (at 25 °C).
10. Nitric acid (HNO_3) is a corrosive acid. When purchased from a chemical company, its molarity is 16 M. How much of the stock solution of nitric acid needs to be used to make 2.5 L of a 0.20 M solution?

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13.6 Reactions of acids

KEY CONCEPT

- The reactions of acids with metals, carbonates and hydroxides, including balanced full and ionic equations, with states indicated.

13.6.1 Salts

Acid reactions occur all around us and in us, and are used in agriculture and in industry. Many industrial products rely on the use of acids as raw materials. The physical properties of acids and bases have been discussed so it is appropriate to look at some chemical properties of acids, in particular. In chemical reactions a new product is formed and, as with many other aspects of chemistry, there are patterns in the products that make it easier to learn what happens. A salt solution is often a product of a reaction of an acid with another substance. If the solid salt is required, the solution can be heated to evaporate the water. The specific reactions that will be studied are reactions of acids with metals, carbonates and hydroxides.

Salt is commonly thought of as a substance that is added to food and has been used for flavouring and as a food preservative for thousands of years. In chemistry, however, there are other salts besides table salt, sodium chloride. A salt is an ionic compound formed when an acid is neutralised by a base. Ionic compounds consist of a positive ion (cation), usually a metal, which comes from the base and a negative ion (anion), usually a non-metal, which comes from the acid. These compounds were studied in topic 3. Examples of salts include KCl, Na_2CO_3 and $Zn(NO_3)_2$.

TABLE 13.5 Some common salts and their applications

Name	Formula	Application
Ammonium sulfate	$(NH_4)_2SO_4$	Fertiliser
Barium sulfate	$BaSO_4$	X-ray of digestive tract
Calcium sulfate dihydrate	$CaSO_4 \cdot 2H_2O$	Plasterboard
Copper sulfate pentahydrate	$CuSO_4 \cdot 5H_2O$	Dyeing; fungicides
Sodium chloride	NaCl	Table salt
Potassium chloride	KCl	Sodium-free salt substitute

13.6.2 Reactions of acids with metals

The reaction of an acid on a metal depends on the type of metal and the strength and concentration of the acid. Many metals, but not all, react with acids and the reactions can be very vigorous (Na, K, Li) to very slow (Ni) or not at all (Cu, Hg, Au, Ag). Other metals that react with acid include Ca, Mg, Mn, Zn, and Fe. The general reaction for an acid and a metal is as follows.

$$\text{Acid} + \text{metal} \rightarrow \text{salt} + \text{hydrogen}$$

The salt produced depends on the acid used: if the acid is hydrochloric acid, the salt is a chloride salt; if the acid is nitric acid, the salt is a nitrate salt; and if the acid used is sulfuric acid, the salt is a sulfate. For example:

$$\begin{array}{ccccccccc} \text{Hydrochloric acid} & + & \text{zinc} & \rightarrow & \text{zinc chloride} & + & \text{hydrogen} \\ 2HCl(aq) & + & Zn(s) & \rightarrow & ZnCl_2(aq) & + & H_2(g) \end{array}$$

Even though this reaction involves an acid, it is not officially classified as an acid reaction because no proton transfer occurs. It is a redox reaction, which involves electron transfer. Redox reactions are discussed in the next topic.

The production of hydrogen can be tested in this experiment using the **'pop' test**. The gas produced can be collected above the reaction text tube (because hydrogen is lighter than air) and then a lit match or ignited splint is placed at the opening. A 'pop' sound is heard if enough gas is collected (figure 13.20), showing the presence of hydrogen gas.

$$2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(l)$$

FIGURE 13.19 Measuring the volume of gas produced during a reaction between acid and zinc

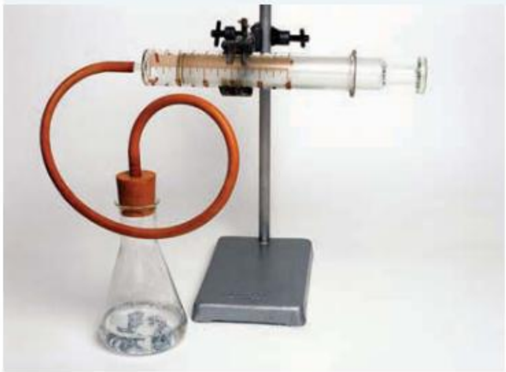
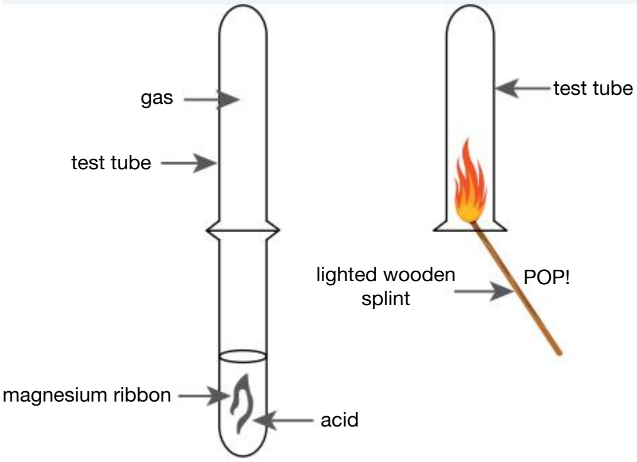


FIGURE 13.20 Pop test for hydrogen



Ionic equations

As with precipitation reactions, it is possible to write ionic equations for reactions of acids. Ionic equations include only the species that have taken part in the reaction and omit spectator ions.

For the reaction of zinc and hydrochloric acid, the acid is ionised and so exists as separate ions in solution, and the product, zinc chloride, is an ionic compound that will be dissociated in solution.

Inspection of this expanded equation shows that the chloride ions are the only species that have not taken part in the reaction. This means that they can be omitted from the ionic equation.

$$2H^+(aq) + \cancel{2Cl^-(aq)} + Zn(s) \rightarrow Zn^{2+}(aq) + \cancel{2Cl^-(aq)} + H_2(g)$$

The ionic equation for this reaction is $2H^+(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + H_2(g)$.

SAMPLE PROBLEM 11

For the reaction between dilute nitric acid and zinc metal, write

a. the full equation representing the reaction.

b. the ionic equation representing the reaction.

THINK	WRITE
a. 1. Write out the correct formulas of the reactants. The general equation for this reaction is acid + metal → salt + hydrogen.	Reactants are HNO_3 and Zn Acid: HNO_3 Metal: Zn
2. The salt formed from nitric acid is a nitrate salt. Write out the correct formulas for the products.	$Zn(NO_3)_2 + H_2$
3. Write out a skeleton equation.	$HNO_3 + Zn \rightarrow Zn(NO_3)_2 + H_2$
4. Balance the equation.	$2HNO_3 + Zn \rightarrow Zn(NO_3)_2 + H_2$
5. Add states to complete the full equation.	$2HNO_3(aq) + Zn(s) \rightarrow Zn(NO_3)_2(aq) + H_2(g)$
b. 1. Write out the equation showing ions present.	$2H^+(aq) + 2NO_3^-(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + 2NO_3^-(aq) + H_2(g)$
2. Write the ionic equation by omitting the species that have not reacted — in this case, the NO_3^- ions.	$2H^+(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + H_2(g)$

PRACTICE PROBLEM 11

For the reaction between magnesium metal and dilute sulfuric acid, write:

a. the full equation representing the reaction

b. the ionic equation representing the reaction.

13.6.3 Reactions of acids with carbonates

When having a break from writing equations, you may decide to have a piece of cake. Consider while you are eating your cake that it is likely that the cake is light and fluffy because a reaction occurred while it was cooking between sodium hydrogen carbonate (baking soda) and an acid to produce carbon dioxide. The carbon dioxide made the cake rise.

Like reactions of acids with metals, the reaction between acids and metal carbonates produce salt and a gas but this time the gas is carbon dioxide and not hydrogen. The salt is the same and water is also produced.

If the metal carbonate is in solution, the equations for the reaction are as follows.

$$\text{Acid} + \text{metal carbonate} \rightarrow \text{salt} + \text{carbon dioxide} + \text{water}$$

For example:

$$2HCl\,(aq) + Na_2CO_3\,(aq) \rightarrow 2NaCl\,(aq) + CO_2\,(g) + H_2O\,(l)$$

This equation can be written as follows.

$$2H^+\,(aq) + \cancel{2Cl^-\,(aq)} + \cancel{2Na^+\,(aq)} + CO_3^{2-}\,(aq) \rightarrow \cancel{2Na^+\,(aq)} + \cancel{2Cl^-\,(aq)} + CO_2\,(g) + H_2O\,(l)$$

The sodium and chloride ions have not reacted and so are spectator ions. The ionic equation is

$$2H^+\,(aq) + CO_3^{2-}\,(aq) \rightarrow CO_2\,(g) + H_2O\,(l)$$

If the metal carbonate is in *solid* form the equations would be as follows.

$$2HCl\,(aq) + Na_2CO_3\,(s) \rightarrow 2NaCl\,(aq) + CO_2\,(g) + H_2O\,(l)$$

This equation can be written as the ions are not dissociated.

$$2H^+\,(aq) + \cancel{2Cl^-\,(aq)} + Na_2CO_3\,(s) \rightarrow 2Na^+\,(aq) + \cancel{2Cl^-\,(aq)} + CO_2\,(g) + H_2O\,(l)$$

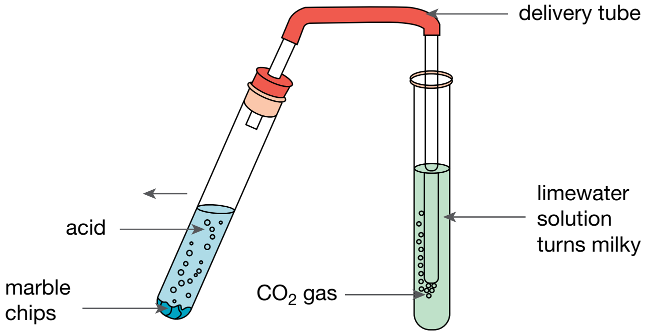
Only the chloride ions have not reacted and so are spectator ions. The ionic equation is

$$2H^+\,(aq) + Na_2CO_3\,(s) \rightarrow Na^+\,(aq) + CO_2\,(g) + H_2O\,(l)$$

To test for carbon dioxide, direct the gas produced (carbon dioxide is heavier than air) into a tube of limewater ($Ca(OH)_2$). If enough gas is produced, a precipitate of calcium carbonate forms.

$$Ca\,(OH)_2\,(aq) + CO_2\,(g) \rightarrow CaCO_3\,(s) + H_2O\,(l)$$

FIGURE 13.21 Test for carbon dioxide



SAMPLE PROBLEM 12

For the reaction between dilute sulfuric acid and solid potassium carbonate, write

a. the full equation representing the reaction.
b. the ionic equation representing the reaction.

THINK	WRITE
a. 1. Write out the correct formulas of the reactants. The general equation is acid + metal carbonate → salt + carbon dioxide + water.	Reactants: H_2SO_4 and K_2CO_3 Acid: H_2SO_4 Metal carbonate:K_2CO_3
2. The salt formed from sulfuric acid is a sulfate salt. Write out the correct formulas for the products.	$K_2SO_4 + CO_2 + H_2O$
3. Write out a skeleton equation.	$H_2SO_4 + K_2CO_3 \rightarrow K_2SO_4 + CO_2 + H_2O$
4. Balance the equation.	$H_2SO_4 + K_2CO_3 \rightarrow K_2SO_4 + CO_2 + H_2O$
5. Add states to complete full equation.	$H_2SO_4\,(aq) + K_2CO_3\,(s) \rightarrow K_2SO_4\,(aq) + CO_2\,(g) + H_2O\,(l)$
b. 1. Write out the expanded equation to show ions present.	$2H^+\,(aq) + SO_4^{2-}(aq) + K_2CO_3\,(s) \rightarrow$ $2K^+\,(aq) + SO_4^{2-} + CO_2\,(g) + H_2O\,(l)$
2. Write the ionic equation by omitting the spectator ions – in this case, the SO_4^{2} ions.	$2H^+\,(aq) + K_2CO_3\,(s) \rightarrow 2K^+\,(aq) + CO_2\,(g) + H_2O\,(l)$

PRACTICE PROBLEM 12

For the reaction between dilute ethanoic acid and solid sodium carbonate, write

a. the full equation representing the reaction.
b. the ionic equation representing the reaction.

Metal hydrogen carbonates also react with acids to produce a salt, carbon dioxide and water; for example, the reaction of hydrochloric acid and sodium hydrogen carbonate is as follows.

$$HCl(aq) + NaHCO_3(s) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

13.6.4 Reactions of acids with hydroxides

Hydrochloric acid is present in our stomachs to assist enzymes to catalyse the digestion of the proteins in the food that we eat. If too much acid is produced, we can feel pain. Relief of this discomfort can be achieved by using antacids, which neutralise some of the excess acid. Antacids contain bases like aluminium hydroxide and magnesium hydroxide.

FIGURE 13.22 Indicator changing colour showing neutralisation has occurred

When an acid and metal hydroxide react, they usually produce a salt and water. This is called **neutralisation**. The general equation is:

$$\text{Acid} + \text{metal hydroxide} \rightarrow \text{salt} + \text{water}$$

For example:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

This equation can be written as follows.

$$H^+(aq) + \cancel{Cl^-(aq)} + \cancel{Na^+(aq)} + OH^-(aq) \rightarrow \cancel{Na^+(aq)} + \cancel{Cl^-(aq)} + H_2O(l)$$

If the spectator ions are omitted, the ionic equation is:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Metal oxides also react with acids to produce a salt and water.

SAMPLE PROBLEM 13

For the reaction between dilute nitric acid and potassium hydroxide solution, write

a. the full equation representing the reaction.

b. the ionic equation representing the reaction.

THINK	WRITE
a. 1. Write the correct formulas of the reactants. The general equation is acid + metal hydroxide → salt + water.	Reactants: HNO_3 and KOH Acid: HNO_3 Metal hydroxide: KOH
2. The salt formed from nitric acid is a nitrate salt. Write out the correct formulas for the products.	$KNO_3 + H_2O$
3. Write out a skeleton equation.	$HNO_3 + KOH \rightarrow KNO_3 + H_2O$
4. Balance the equation.	$HNO_3 + KOH \rightarrow KNO_3 + H_2O$
5. Add states to complete full equation.	$HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$
b. 1. Write out the equation showing all of the ions present.	$H^+(aq) + NO_3^-(aq) + K^+(aq) + OH^-(aq) \rightarrow K^+(aq) + NO_3^-(aq) + OH^-(aq) + H_2O(l)$
2. Write the ionic equation by omitting the spectator ions – in this case, the K^+ and NO_3^- ions.	$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

PRACTICE PROBLEM 13

For the reaction between dilute hydrochloric acid and magnesium hydroxide solution, write

a. the full equation representing the reaction.

b. the ionic equation representing the reaction.

An unusual acid–base reaction does not involve water. Pure hydrogen chloride, HCl, is a gas; it can be dissolved in water to produce a concentrated solution. Both concentrated HCl and ammonium hydroxide, NH_4OH, produce fumes. These fumes react to produce a fine white solid, ammonium chloride, NH_4Cl.

Summary of general reactions for acids:

- **acid + metal → salt and hydrogen**
- **acid + metal carbonate → salt + water + carbon dioxide**
- **acid + metal hydroxide → salt + water.**

FIGURE 13.23 Hydrogen chloride gas and ammonia gas reacting to form a salt.

on Resources

Digital documents Experiment 13.4 Making copper sulfate crystals (doc-30845)
Experiment 13.5 The reactivity of metals and their salts (doc-30865)

Teacher-led video Experiment 13.4 Making copper sulfate crystals (tlvd-0633)

13.6 EXERCISE

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1. Complete the following table.

Acid	Formula of acid	Type of salt formed from the acid	Anion present
Hydrochloric			
		Nitrate	
			SO_4^{2-}
Ethanoic			

2. Write the general equations for the reactions of acids with the following.
   (a) Metals
   (b) Metal carbonates
   (c) Metal hydroxides
3. Complete and balance the following equations.
   (a) $Zn(s) + HBr(aq)$
   (b) $Ca(s) + H_2SO_4(aq)$
   (c) $K(s) + HNO_3(aq)$
4. Complete and balance the following equations.
   (a) $MgCO_3(s) + CH_3COOH(aq)$
   (b) $Na_2CO_3(s) + HCl(aq)$
   (c) $MnCO_3(s) + HNO_3(aq)$

5. Complete and balance the following equations.
   (a) $NaOH + H_2SO_4$
   (b) $Ca(OH)_2 + HCl$
   (c) $Mg(OH)_2 + HNO_3$
6. Write ionic equations for part (a) of questions 3, 4 and 5. State the spectator ions.
7. For each of the following reactions, write full equations and ionic equations.
   (a) The acid and solid metal that could be used to produce calcium sulfate
   (b) The acid and metal carbonate that could be combined to produce zinc nitrate
   (c) The acid and metal carbonate solution that could be combined to produce sodium chloride
8. (a) Describe how to collect and test for hydrogen gas.
   (b) Write a balanced equation for this reaction.
   (c) If no reaction was observed when this test was applied, does this imply that no hydrogen was produced?
9. (a) Describe how to collect and test for carbon dioxide gas.
   (b) Write a balanced equation for this reaction.
10. Write both full and ionic equations for the reactions that result when the following substances are mixed. State the spectator ions. Assume all reactions go to completion.
   (a) Nickel metal and dilute sulfuric acid
   (b) Aluminium metal and dilute sulfuric acid
   (c) Solid calcium carbonate and dilute nitric acid

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13.7 Acids and the environment

KEY CONCEPT

- The causes and effects of a selected issue related to acid–base chemistry.

Acidic oxides react with water to form acids, and are usually oxides of non-metals. Some acidic oxides are released into the atmosphere as a result of human activity and have an impact on the environment.

13.7.1 Acid rain

Normal rain is slightly acidic due to dissolution of carbon dioxide in water according to the following equation:

$$CO_2(g) + H_2O(g) \rightarrow H_2CO_3(aq)$$

The product is the weak acid carbonic acid, which reacts slightly with water to produce a few hydronium ions, as per the following equation:

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

Many industrial processes involve the burning of coal, oil or some other fossil fuel. Many of these fuels contain sulfur in varying amounts. When sulfur is burned in air, it forms sulfur dioxide, SO_2. This gas is often released into the air in vast quantities.

It must be noted that sulfur dioxide is released in many natural processes as well. In particular, active volcanoes release a large amount of sulfur dioxide into the air. The atmosphere can cope with large quantities of sulfur dioxide if it is given time to disperse the gas and break it down. When a large number of industries are all producing the gas over a small area, however, it cannot disperse in the air fast enough and becomes too concentrated to be safe.

Sulfur dioxide (SO_2) can react with water to produce sulfurous acid (a weak acid). It can also react with oxygen in the air to produce sulfur trioxide, SO_3, which then reacts with water to produce the strong acid sulfuric acid. The equations for these three processes are:

$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$$
$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$
$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

Other gases contribute to the **acid rain** problem and many of these are produced by industry and by every car that uses the internal combustion engine.

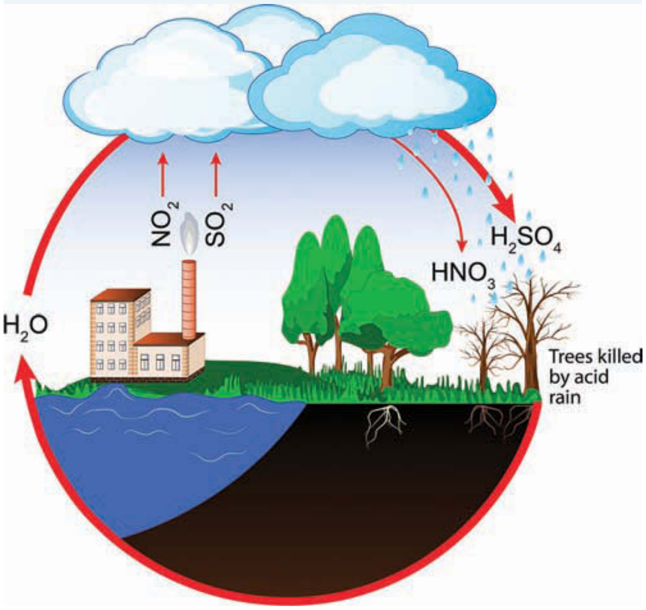
In a car engine, the temperature produced by the spark plug is sufficient to cause the nitrogen and oxygen in the air to combine and form nitrogen monoxide, NO. Nitrogen monoxide combines spontaneously with oxygen in the air to form nitrogen dioxide, NO_2. This nitrogen dioxide reacts with rainwater to form nitric acid, another contributor to the problem known as acid rain.

The equation for this reaction is:

$$3NO_2(g) + H_2O(g) \rightarrow 2HNO_3(aq) + NO(g)$$

It must be noted, however, that carbonic acid and the oxides of sulfur are largely responsible for acid rain.

FIGURE 13.24 The production of acid rain



Impacts and control

Acid rain has been responsible for the defoliation of significant amounts of forest in both Europe and North America. In Australia, it is one of the principal causes of the 'lunar' landscape that occurred around Queenstown in Tasmania where large amounts of sulfur dioxide were produced as a result of copper ore smelting during the last century.

Where acid rain has run off into lakes, the pH of such lakes has in some instances dropped to the point where they have been unable to sustain life.

A range of responses has now evolved to reduce the input of gases into the atmosphere that cause acid rain. These include:

- catalytic converters in car exhausts to change nitrogen oxides into nitrogen and oxygen
- the use of natural gas as a fuel (natural gas contains lower levels of sulfur impurity) in place of other fuels such as oil and coal
- switching to coal with a lower sulfur content
- taking measures to assist the dispersal of SO_2, such as by using taller chimney stacks
- the use of scrubbing, which is where exhaust gases are treated to remove undesirable emissions.

An increasing number of processes are now removing SO_2 from their emissions and concentrating it so that it can be used as a feedstock to produce sulfuric acid. This has an extra advantage because sulfuric acid is a valuable industrial acid and can be sold to offset the costs of the original process.

13.7.2 Ocean acidification

As we have seen previously, carbon dioxide dissolves in water to produce carbonic acid. The oceans are slightly alkaline, but research has shown that, since the start of the Industrial Revolution, their pH has dropped from 8.2 to 8.1. While many localised variations are in these figures, such a change is significant. This is directly attributable to the increased level of carbon dioxide in the atmosphere. Although the effect of this is the subject of ongoing research, problems that might arise as a result of this acidification include:

- coral bleaching
- interference with marine organisms, particularly those at the lower end of the food chain
- reproductive disorders in certain marine organisms
- interference with shell building in shellfish and other similar organisms.

The shells of many marine species have calcium carbonate as an important constituent. Others, including corals, form skeletons using calcium carbonate. Although calcium carbonate is virtually insoluble, Ca^{2+} and CO_3^{2-} ions are present in sea water at concentrations that are low but nevertheless important. However, carbonate ions can be removed by a reaction between themselves, water and carbon dioxide according to:

$$CO_2(g) + H_2O(l) + CO_3^{2-}(aq) \rightarrow 2HCO_3^-(aq)$$

If the carbonate ion concentration is lowered too much, it is possible that the calcium carbonate in the shells of marine organisms may dissolve in an attempt to replace the removed carbonate ions.

$$CaCO_3(s) \rightarrow Ca^{2+}(aq) + CO_3^{2-}(aq)$$

13.7 EXERCISE

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1. What is acid rain?
2. Write a word and symbol equation showing the formation of carbonic acid.
3. Explain whether carbonic acid a strong or weak acid. Use an equation in your answer.
4. Show using equations how the presence of SO_2 in the atmosphere results in the production of sulfuric acid.
5. List some sources of SO_2 in the atmosphere.
6. Describe three possible methods of controlling the input of SO_2 in the atmosphere.
7. Explain how acid–base reactions are involved in the acidification of the ocean.
8. Calculate the concentration of hydroxide ions and hydrogen ions in the ocean where the pH is 8.1.
9. State problems that might occur as oceans become less basic.
10. Outline the effect the decrease in pH has on shell fish, including an equation to illustrate your answer.

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13.8 Review

13.8.1 Summary

Acids and bases

- Common acids include HCl, H_2SO_4, HNO_3, H_3PO_4, CH_3COOH and H_2CO_3. These have similar properties. They taste sour, are corrosive, molecular in structure and turn litmus red.

- Common bases include metal hydroxides (for example, LiOH, NaOH, KOH) and molecular ammonia, NH_3. These have similar properties. They taste bitter, feel slippery, are corrosive, turn litmus blue and apart from ammonia are ionically bonded.
- Alkalis are bases that dissolve in water (for example, group 1 hydroxides).
- Acids and bases need to be handled with care because some can be corrosive.
- When diluting acids, add acid to water.

Brønsted–Lowry theory of acids and bases

- The Brønsted–Lowry theory defines acids and bases as follows:
  - An acid is a proton donor.
  - A base is a proton acceptor.
- Acid–base reactions thus involve proton transfer, in which one hydrogen ion, H^+, is transferred from the acid to the base.
- A conjugate acid–base pair is formed when an acid reacts with a base.
  - These pairs differ by a proton, H^+ (for example, HCl/Cl^-, H_3O^+/H_2O).
  - To find a conjugate pair, subtract one H^+ from the acid and that will be the conjugate acid.
- Acids can be classified according to their ability to donate protons:
  - Monoprotic acids can donate one proton (for example, HCl).
  - Polyprotic acids can donate more than one proton: diprotic acids can donate two protons (for example, H_2SO_4), and triprotic acids can donate three protons (for example, H_3PO_4).
  - Amphiprotic substances can act as acids or bases depending on their chemical environment (for example, H_2O, HS^-, HSO_4^-).

Concentrating on acids and bases

- Concentration refers to the quantity of solute in a particular volume of solution.
- Concentration is usually measured in moles per litre (mol L^{-1} or M).
- To calculate the number of moles in a solution, use the formula $n = c \times V$, where c is the concentration in mol L^{-1} and V is the volume measured in litres.
- Concentrated solutions have a large quantity of solute dissolved in a given volume of solution. Dilute solutions have a small quantity of solute dissolved in a given volume of solution.
- Acids and bases may behave in water in the following ways:
  - Ionisation occurs when an acid donates a proton to water. The water acts as a base and forms the hydronium ion, H_3O^+.
  - When an ionic base dissolves in water, it dissociates or separates into its constituent ions.
- The strength of acids (or bases) relates to their ability to readily donate (or accept) protons.
- Concentration of acids (or bases) relates to the amount of the substance dissolved in a given volume of water. A large amount results in a concentrated solution while a small amount gives a dilute solution.
- There are only a few strong acids including nitric, sulfuric and hydrochloric acids. Most acids are weak acids; for example, ethanoic acid is a weak acid.
- Strong bases include sodium hydroxide and potassium hydroxide. Ammonia is a weak base.

pH and acidity

- In any aqueous solution at 25 °C, $K_W = [H_3O^+][OH^-] = 10^{-14}$ M^2. This is known as the ionic product of water.
- The pH scale is a convenient way of measuring the level of acidity or basicity in a solution.
  - In neutral solutions, $pH = 7$.
  - In acidic solutions, $pH < 7$.
  - In basic solutions, $pH > 7$.
- For strong acids $[H_3O^+]$ = the concentration of the acid, because the acid is fully ionised.
- pH depends on H_3O^+ concentration and is defined by the formula $pH = -\log_{10}[H_3O^+]$.
- To find the $[H_3O^+]$ rearrange the formula: $[H_3O^+] = 10^{-pH}$.
- To calculate the pH of a basic solution, first find the $[H_3O^+]$ using the ionic product of water.

- pH measurements can be used to compare the relative strengths of acids (and bases), providing the concentrations of the acids (or bases) are the same.
- An indicator is a dye that is a weak acid or base. It has a different colour in acidic and basic solutions.
- Universal indicator is a mixture of indicators and is a useful and commonly used indicator.
- Solutions may be required to be diluted for different purposes.
- To calculate volumes or concentrations after dilution, the formula $c_1V_1 = c_2V_2$ is used. The same unit for volume must be used for V_2 and V_1.
- Water added in a dilution can be calculated using $V_2 - V_1$.

Reactions of acids

- Chemical reactions of acids include:
  - acid + metal (not Cu, Hg, Ag or Ag) → salt + hydrogen gas
  - acid + metal carbonate → salt + carbon dioxide gas + water
  - acid + metal hydrogen carbonate → salt + carbon dioxide gas + water
  - acid + metal oxide (basic oxide) → salt + water
  - acid + base (metal hydroxide) → salt + water.
- When writing ionic equations, omit spectator ions
- To test for hydrogen, collect gas and test with lit match. A ‘pop’ shows that H_2 is present.
- To test for carbon dioxide, collect gas and direct into limewater. Cloudiness shows CO_2 is present.
- When an acid and base react, the process is called neutralisation and a salt and water are usually produced. The ionic equation of neutralisation is:

$$H^+ (aq) + OH^- (aq) \rightarrow H_2O \text{ (l) (where } H^+ \text{ is written in place of } H_3O^+\text{).}$$

Acids and the environment

- Two significant environmental issues associated with acid–base chemistry are acid rain and ocean acidification.
- Acidic oxides in the atmosphere such as CO_2, SO_2, SO_3 and NO_2 react with water and contribute to acid rain. They have detrimental effects on the environment.
- Carbon dioxide dissolved in the ocean decreases the pH and affects the growth and development of aquatic life.

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13.8.2 Key terms

acid	substance that can donate a proton to a base
acid rain	rain made acidic by mixing with atmospheric pollution resulting a pH less than 5
acidic oxide	oxide of a non-metal
alkali	soluble base
amphiprotic	substance that can accept or donate protons
base	substance that can accept a proton from an acid
concentrated	has a large quantity of solute in a known volume of solution
concentration	the amount of solute that is dissolved in a known volume of solution
conjugate acid	the product that a base forms when it has accepted a proton from an acid
conjugate base	the product that an acid forms when it has donated a proton to a base
dilute	has a small quantity of solute in a known volume of solution
diprotic	can donate two protons

dissociation the process by which ions separate when an ionic compound dissolves
double arrow used to show a chemical equation is reversible
electrolyte solution or liquid that can conduct electricity
hydronium ion H_3O^+, combination of a proton with a water molecule
indicator substances that undergo an observable change in an acidic or basic solution
ionise reaction in which a substance reacts with water to form ions
logarithm quantity representing the power to which a fixed number (the base) must be raised to produce a given number
molarity concentration measured in units of moles of solute per litre (mol l^{-1} or M) of solution
monoprotic can donate one proton
neutralisation chemical reaction in which an acid and a base react to form a salt
pH scale used to measure acidity; pH can be calculated using $\text{pH} = -\log_{10}[H_3O^+]$
polyprotic able to donate more than one proton to a base
pop test test for hydrogen gas where a small amount of the gas is collected and a lit match is placed nearby
salt ionic compound formed when an acid is neutralised by a base
self-ionisation constant K_w; in all aqueous solutions this value is constant and results from the self-ionisation of water: $H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$

$$K_w = [H_3O^+][OH^-] = 10^{-14}\ M^2$$

species general term that could refer to atoms, molecules, molecular fragments or ions
stock solution concentrated solution that will be diluted as required for later use
strong acid acid that fully ionises in water
strong base base that fully dissociates in water
triprotic acid that can donate three protons
weak acid acid that partially ionises in water
weak base base that partially dissociates in water

on Resources

Digital document Key terms glossary — Topic 13 (doc-30927)

13.7.3 Practical work and investigations

Experiment 13.3

The reactions of acids

Aim:
To investigate the reactions of a typical acid (dilute hydrochloric acid) with metals, metal hydroxides, and metal carbonates

Digital document: doc-30860
Teacher-led video: tlvd-0632

Experiment 13.4

Making copper sulfate crystals

Aim: To observe the reaction between copper oxide and sulfuric acid.

Digital document: doc-30845
Teacher-led video: tlvd-0633

on Resources

Digital documents Practical investigation logbook (doc-30928)
Experiment 13.1 Indicators and pH (doc-30861)
Experiment 13.2 Finding the pH of common household substances (doc-30855)
Experiment 13.5 The reactivity of metals and their salts (doc-30865)

13.8 Exercises

To answer questions online and to receive immediate feedback and sample responses for every question, go to your learnON title at www.jacplus.com.au.

13.8 Exercise 1: Multiple choice questions

1. Which of the following is the best definition of an acid?
   - A. An acid is an electron acceptor.
   - B. An acid is a substance that tastes sour.
   - C. An acid is a substance that donates H^+.
   - D. An acid is a substance that accepts H^+.
2. Hydrogen sulfide, H_2S, is a gas at room temperature and dissolves in water. When this solution is tested, it conducts electricity and turns litmus solution red. Why is the solution acidic?
   - A. The H_2S has gained a proton.
   - B. The H_2O has gained a proton.
   - C. The H_2S and H_2O have gained protons.
   - D. The H_2S and H_2O have lost protons.
3. What is the conjugate base of the species $H_2PO_3^-(aq)$?
   - A. $H_3PO_3(aq)$
   - B. $HPO_4^{2-}(aq)$
   - C. $H_2PO_4(aq)$
   - D. $HPO_3^{2-}(aq)$
4. Consider the following equations:
   i. $Zn(s) + 2H_3O^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g) + 2H_2O(l)$
   ii. $Cl^-(aq) + H_3O^+(aq) \rightarrow HCl(aq) + H_2O(l)$
   iii. $H_3O^+(aq) + OH^-(aq) \rightarrow 2H_2O(l)$.
   Which can represent acid–base reactions?
   - A. (i) and (ii)
   - B. (i) and (iii)
   - C. (ii) and (iii)
   - D. (i), (ii) and (iii)
5. Which of the following materials is the most acidic?
   - A. Grapefruit juice of pH 3.0
   - B. Detergent of pH 10.5
   - C. Gastric juice of pH 1.0
   - D. Sea water of pH 8.5
6. When hydrogen chloride gas is dissolved in water, the resulting solution can conduct an electric current. What is the best explanation for this?
   - A. Ions in the hydrogen chloride gas are released when dissolved in water.
   - B. The water reacts with the hydrogen chloride to form ions.
   - C. The water molecules carry the electric current in one direction while the hydrogen chloride molecules carry it in the other direction.
   - D. The water contains impurities that react with the hydrogen chloride to form ions.

7. Which one of the following species acts as an amphiprotic species in aqueous solution?

A. SO_4^{2-}

B. HCO_3^-

C. PO_4^{3-}

D. Cl^-

8. What is the strength of an acid determined by?

A. The number of hydrogen ions present

B. Its concentration

C. Its degree of ionisation in aqueous solution

D. Its ability to change the colour of litmus

9. Consider the reaction between methanoic acid and water:

$HCOOH + H_2O \rightleftharpoons HCOO^- + H_3O^+$

Which of the following pairs are Brønsted–Lowry acids?

A. H_2O, HCOOH

B. HCOOH, H_3O^+

C. H_2O, H_3O^+

D. HCOOH, $HCOO^-$

10. What is a substance that is capable of acting as both an acid and as a base known as?

A. Amphiprotic

B. Conjugated

C. Miscible

D. Diprotic

11. At 25 °C, the equation representing the self-ionisation of water is

A. $H_2O(l) + H_2O(l) \rightleftharpoons H_2O_2(aq) + O_2(g)$

B. $H_2O(l) + H_2O(l) \rightleftharpoons 4H^+(aq) + 2O^{2-}(aq)$

C. $H_2O(l) + H_2O(l) \rightleftharpoons 2H_2(g) + O_2(g)$

D. $H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$

12. What is the pH of 0.0001 molar HCl solution?

A. 1

B. 2

C. 3

D. 4

13. Which of the following is correct for an acidic solution at 25 °C?

A. $[H_3O^+] > [OH^-]$, pH > 7

B. $[H_3O^+] > [OH^-]$, pH < 7

C. $[H_3O^+] < [OH^-]$, pH > 7

D. $[H_3O^+] < [OH^-]$, pH > 7

14. Which of the following reactions takes place when a metal carbonate reacts with an acid?

A. Acid + metal carbonate → salt + water + carbon dioxide

B. Acid + metal carbonate → salt + water + oxygen

C. Acid + metal carbonate → salt + water + hydrogen

D. Acid + metal carbonate → metal oxide + carbon dioxide

15. When dissolved in rain, which of the following gases does not cause it to be acidic?

A. O_2

B. CO_2

C. NO_2

D. SO_3

13.8 Exercise 2: Short answer questions

1. **a.** What are the characteristic properties of acids?
  b. Illustrate two of these properties with an appropriate equation.

2. **a.** Define a base and give two examples.
  b. Define an alkali and give two examples.

3. Classify the following list of properties and uses as acidic, basic, both acidic and basic, or neutral.
  a. A sour taste
  b. The bubbles in champagne
  c. Corrosive
  d. Dissolves oxide coating on metals
  e. Dissolves marble statues
  f. Feels slippery
  g. A bitter taste
  h. Breaks down vegetable matter
  i. Can be used to clean drains
  j. Produced in stomachs
  k. Used for upset stomachs

4. **a.** In terms of the Brønsted–Lowry theory, define an acid, a base and a proton-transfer reaction. You may use an equation in your explanations.
  b. Explain, using equations, why a solution of hydrochloric acid conducts electricity but pure hydrogen chloride does not.

5. Antacid tablets are used to neutralise the hydrochloric acid in the stomach. Use the two bases mentioned in the discussion of antacids in section 13.6.3 to write neutralisation equations for the reactions that may happen in our bodies.

6. **a.** What are conjugate acid–base pairs?
  b. State the formulas of the conjugate acid of the following bases:
    i. Cl^- **ii.** CO_3^{2-} **iii.** HSO_4^- **iv.** H_2O **v.** OH^-.
  c. State the formulas of the conjugate base of the following acids:
    i. HF **ii.** HCO_3^- **iii.** HSO_4^- **iv.** H_2O **v.** H_2S.

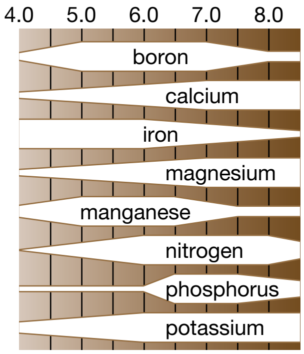
7. Write two equations to illustrate the amphiprotic ion HCO_3^-.

8. Sulfuric acid is a strong, diprotic acid. Explain, using equations, what this term means.

9. Beatrice took the lid off a bottle of glacial ethanoic acid, CH_3COOH, and immediately noticed the very potent smell of vinegar. 'This must be a very strong acid', she exclaimed. 'No', said her sister Freyja. 'It is merely a concentrated acid'. Using appropriate definitions, explain who is right.

10. In a pH test on a swimming pool, the water is found to have a pH of 4.5.
  a. What would you add to the water to raise the pH?
  b. The pH should be about 7.5 for best conditions. Is this acidic or alkaline?

11. Soil contains a number of different elements. Plants need to take up these elements in order to grow and flower. If the pH of soil is too high or too low, plants cannot take up these elements. The figure shows how the pH of the soil influences the amount of an element that can be taken up by a plant. The narrower the bar, the harder it is for the plant to take up the element.



  a. Which elements can easily be absorbed at a pH below 4.5?
  b. Which elements can be absorbed at a pH of 8?
  c. Which elements cannot be taken up easily if the pH is 6?
  d. Which elements cannot be taken up easily if the soil pH is 8?
  e. Fuchsias require the soil to be acidic. Which elements would be most important to these plants?
  f. Azaleas require a pH of around 5.5. Which elements are not important to these plants?

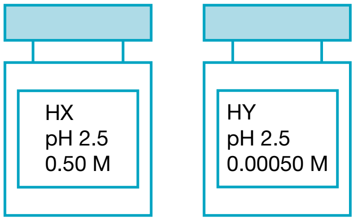
g. A gardener comes to you with a problem. He says that he cannot grow lettuce successfully any longer. It is not an insect problem or a water problem. Design an experiment you could do to find out the problem with his soil. (Note: Lettuces need a lot of iron.)

12. Calculate the pH of solutions that have the following H_3O^+ concentrations.
    a. 0.0010 M
    b. 0.00050 M
13. Explain why, when an acid is diluted by a factor of 2, its pH does not change by the same factor.
14. Calculate the pH of solutions that have the following OH^- ion concentrations.
    a. 0.0010 M
    b. 0.000500 M
15. Write both full and ionic equations for the reactions that result when the following substances are mixed. Assume all reactions go to completion.
    a. Solid sodium sulfide and phosphoric acid, H_3PO_4
    b. Solid copper(II) oxide and hydrochloric acid
    c. Potassium hydroxide solution and sulfuric acid

13.8 Exercise 3: Exam practice questions

Question 1 (1 mark)

Two acids, HX and HY, are shown in the figure with their concentrations and pH values.



Which of the following statements is/are correct?

The increase in carbon dioxide in the atmosphere is causing the pH of rain water to decrease.

i. HX is a stronger acid than HY.
ii. Both acids ionise to the same extent in water.
iii. The reaction showing the ionisation of HY in water should be written with a double arrow.

A. (i) only
B. (ii) and (iii) only
C. (ii) and (iii) are correct
D. No statements are correct

Question 2 (1 mark)

Ethanoic (acetic) acid is a weak monoprotic acid.

a. What is meant by the term ‘monoprotic’? **1 mark**
b. Write the equation for the ionisation of ethanoic acid in water. **1 mark**
c. Write an equation for the neutralisation of ethanoic acid by sodium hydroxide. **1 mark**
d. What is meant by the term ‘weak’? Use your equation from part (b) to illustrate your answer. **2 marks**
e. Describe a test that could be used to show that ethanoic acid is a weak acid. **1 mark**
f. Draw the structure of an ethanoic acid molecule. Given that it is monoprotic, circle the hydrogen atom that is donated when it reacts with a base. **2 marks**

Question 3 (4 marks)

The hydrogen carbonate ion is amphiprotic.

a. What is meant by the term ‘amphiprotic?’ **1 mark**
b. Write two different equations for the reaction of the hydrogen carbonate ion with water that demonstrate its amphiprotic nature. **2 marks**
c. Solutions of the hydrogen carbonate ion in water are slightly basic. Comment on this with respect to the two equations that you have written for part (b). **1 mark**

Question 4 (5 marks)

Ammonium sulfatenitrate is used as a fertiliser.

a. State two compounds that could be mixed to produce ammonium sulfate. **1 mark**

b. Write the full equation for this reaction. **1 mark**

c. Write the ionic equation for this reaction. **1 mark**

d. How would solid ammonium sulfate be obtained? **1 mark**

e. When the solid fertiliser dissolves in water it dissociates. Write an equation showing this process. **1 mark**

Question 5 (6 marks)

The increase in carbon dioxide in the atmosphere is causing the pH of rain water to decrease.

a. Write the equation for this reaction. **1 mark**

b. Explain using a balanced equation how this acid can produce the ion that causes the decrease in pH. **2 marks**

c. Acid rain causes significant damage to limestone, $CaCO_3$, used in buildings. Write an equation showing the reaction of sulfuric acid with limewater. **1 mark**

d. Name three acidic oxides that contribute to acid rain and provide their formulas. **2 marks**

13.8 Exercise 4: studyON Topic Test online only

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Create unique tests and exams from our extensive range of questions, including practice exam questions. Access the Assignments section in learnON to begin creating and assigning assessments to students.

14 Redox reactions in water

14.1 Overview

Numerous **videos** and **interactivities** are available just where you need them, at the point of learning, in your digital formats, learnON and eBookPLUS at www.jacplus.com.au.

14.1.1 Introduction

FIGURE 14.1 Fireworks are very rapid redox reactions

Oxidation–reduction, or redox, reactions continually occur all around us and are one of the largest groups of chemical reactions. Redox chemistry is a fundamental part of life — photosynthesis and respiration are redox reactions. Our bodies work by redox reactions — the food we eat is oxidised to enable us to obtain the energy we need to live. If we break a bone, electrical currents are generated around the damaged area as part of the healing process. The conduction of impulses in nerves and the supply of battery power in our mobile phones, calculators, computers and cars rely on spontaneous redox reactions. Many metals are extracted by reduction reactions and municipal water supplies and swimming pools are treated with oxidising agents that act as bactericides and algaecides. Bushfires are uncontrolled redox reactions on a large scale, while fireworks and explosions are very fast redox reactions.

In this topic, you will learn about these types of reactions.

14.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- oxidising and reducing agents, conjugate redox pairs and redox reactions, including writing of balanced half and overall redox equations with states indicated
- the reactivity series of metals and metal displacement reactions, including balanced redox equations with states indicated
- the causes and effects of a selected issue related to redox chemistry.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of learning and assessment. Experiments and investigations, supported by a **Practical investigation logbook** and **Teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

on Resources

Digital documents Key science skills (doc-30903)
Key terms glossary — Topic 14 (doc-30944)
Practical investigation logbook (doc-30945)

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To access key concept summaries and practice exam questions download and print the **studyON: Revision and practice exam question booklet** (doc-30946).

14.2 Redox reactions and agents

KEY CONCEPT

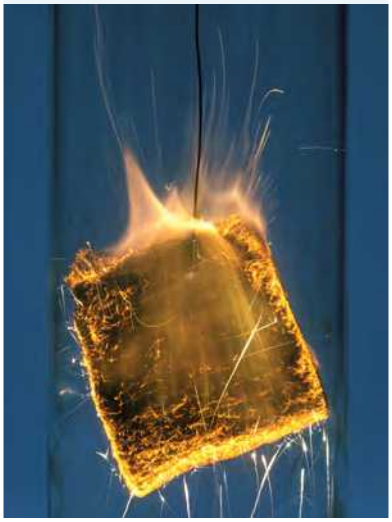
- Oxidising and reducing agents, conjugate redox pairs and redox reactions, including writing of balanced half and overall redox equations with states indicated

14.2.1 Introduction and history of redox reactions

The word **redox** is a combination of shortened forms of the words '**reduction**' and '**oxidation**'. Redox reactions (oxidation–reduction reactions) were originally defined in terms of loss or gain of oxygen but are now characterised as electron transfer reactions. Oxidation is the loss of electrons and reduction is the gain of electrons. The substance that is being reduced is called the **oxidising agent** and the substance being oxidised is called the **reducing agent**. Although this sounds confusing, with practice, you will become familiar with these terms. Equations for redox reactions can be split into separate half-equations, one showing oxidation and the other showing reduction. These can be balanced using a series of steps that are demonstrated in this topic. The reactivity of metals is explored as an example of redox reactions and an important real-life application, corrosion, is investigated.

Early chemists knew that a gas called oxygen was essential for the survival of living things. This prompted them to study the reactions of oxygen with other substances in an effort to find out more about its behaviour. They found that many substances combined with oxygen, so the term oxidation was used to describe these reactions. The combustion reactions of hydrocarbons such as the burning of propane:

FIGURE 14.2 Steel wool (iron) burning in air to form iron(III) oxide



$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

and the burning of iron in air:

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

were described as oxidation reactions because a reactant, or some part of it, combined with oxygen.

Conversely, reactions that involved the decomposition of a compound, with the loss of oxygen, were called reduction reactions because the compound seemed to be 'reduced' to something simpler. For example, copper(II) oxide may be reduced to copper by hydrogen:

$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$$

Similarly, iron(III) oxide is reduced to iron by carbon monoxide in a blast furnace:

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$$

As one reactant is reduced, the other reactant is oxidised. Oxidation and reduction must occur at the same time.

Oxidation is the gain of oxygen; reduction is the loss of oxygen.

14.2.2 Oxidising and reducing agents

In any oxidation–reduction reaction:

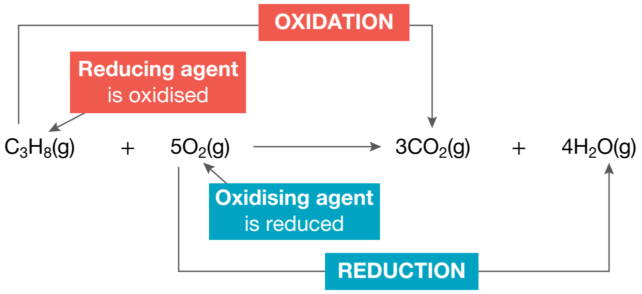
- The oxidising agent *causes* the oxidation of *another substance* and is *itself* reduced.
- The reducing agent *causes* the reduction of *another substance* and is *itself* oxidised.

Oxidising agents are also known as **oxidants** and reducing agents are known as **reductants**.

The oxidising agent is the species that is reduced.
The reducing agent is the species that is oxidised.

These terms are demonstrated in the reaction $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$ in figure 14.3.

FIGURE 14.3 Oxidation–reduction terminology in the equation $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$



Modern definition of oxidation and reduction

The modern definition of oxidation and reduction processes describes redox reactions in terms of electron transfer. For example, the combustion of magnesium in oxygen to form magnesium oxide involves the oxidation of magnesium.

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

A similar reaction occurs when magnesium chloride is formed by the combustion of magnesium in chlorine, but in this case no oxygen is involved.

$$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$$

Analysis of these two reactions in terms of a transfer of electrons reveals that they are essentially the same process. Both **products** are ionic substances and so contain Mg^{2+} ions. In the process of oxidation, each atom of magnesium has lost two electrons.

$$Mg \rightarrow Mg^{2+} + 2e^-$$

In general, a substance that is oxidised is one that loses electrons and is, therefore, an electron donor.

When magnesium reacts in this way with either oxygen or chlorine, these atoms accept these electrons and become O^{2-} and Cl^- anions in the ionic lattice of the products.

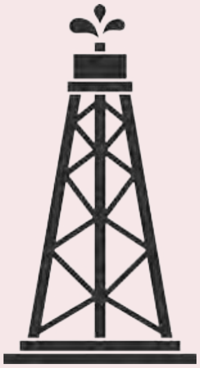
$$O_2 + 4e^- \rightarrow 2O^{2-}$$
$$Cl_2 + 2e^- \rightarrow 2Cl^-$$

In general, a substance that is reduced is one that gains electrons and is, therefore, an electron acceptor.

Redox reactions, therefore, are those in which electrons are transferred from one reactant to another.

Redox reactions are electron transfer reactions.
Oxidation is defined as a loss of electrons to another substance.
Reduction is defined as a gain of electrons from another substance.

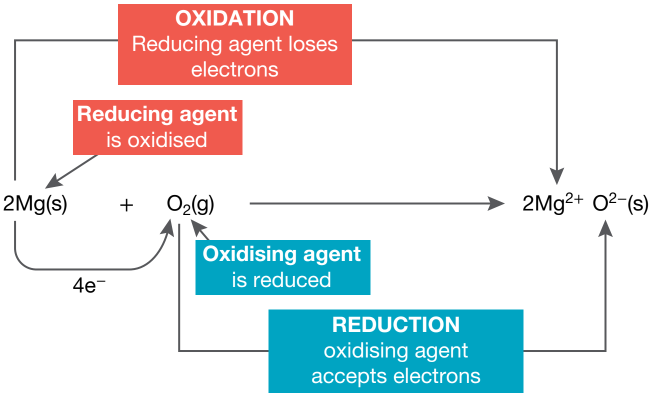
This can be remembered using the acronym 'OIL RIG':
Oxidation
Is
Loss of electrons
Reduction
Is
Gain of electrons



Oxidation must be accompanied by reduction; that is, oxidation cannot occur unless reduction occurs simultaneously. When writing equations for redox reactions, we do not show electrons because all electrons given off during oxidation are taken in during reduction. However, if we are considering oxidation or reduction reactions separately, it is appropriate (and necessary) to write reactions that do show the electrons, as has been done in the examples earlier in this section. Such equations are called partial equations or **half-equations**. It should be noted if you multiply the two half-equations by a factor so the same number of electrons are on each side, and then add them together, you do get the overall ionic equation.

In summary, the reactant that is losing electrons (undergoing oxidation) is called the reducing agent or reductant, because it transfers electrons to another substance and causes it to be reduced. The reactant that is gaining electrons (undergoing reduction) is called the oxidising agent or oxidant, because it accepts the electrons from the reducing agent and causes it to be oxidised. This is shown in figure 14.4.

FIGURE 14.4 Oxidation and reduction in terms of electron transfer



on Resources

Video eLesson Redox-electron transfer (eles-2495)

SAMPLE PROBLEM 1

Explain why the following reaction is described as a redox reaction and identify the species oxidised and reduced:

$$\mathbf{Fe(s) + 2Ag^{+}(aq) \rightarrow 2Ag(s) + Fe^{2+}(aq)}$$

THINK	WRITE
1. Redox reactions are electron transfer reactions. To decide if electrons have been transferred, compare reactants and products and identify the changes.	$Fe(s)$ has become $Fe^{2+}(aq)$. $Ag^{+}(aq)$ has become $Ag(s)$.
2. Identify which reactant has lost electrons and which reactant has gained electrons. The species that has lost electrons is oxidised. The species that has gained electrons is reduced. **TIP:** When naming ions, is it important to refer to them as ions and not just state the metal. In this case, the silver **ion** is reduced.	Fe has lost two electrons to become Fe^{2+}. Ag^{+} has gained an electron to become Ag. Fe is oxidised Ag^{+} is reduced; that is, the silver ion is reduced.
3. Electron transfer has occurred.	This is a redox reaction because electron transfer has occurred from Fe to Ag^{+}.

PRACTICE PROBLEM 1

Explain why the following reaction is described as a redox reaction and identify the species oxidised and reduced:

$$\mathbf{Cl_2(aq) + 2Br^{-}(aq) \rightarrow Br_2(aq) + 2Cl^{-}(aq)}$$

14.2.3 Redox equations

Redox reactions can often be described using half-equations. For more complex redox reactions, the half-equations need to be balanced and then added together to correctly capture the reaction.

Half-equations

Half-equations are a useful way of understanding the processes involved in a redox reaction. The following example describes how half-equations are written.

When an iron nail is placed in a blue copper sulfate solution, the nail becomes coated with metallic copper and the blue colour of the solution fades (figure 14.5). A redox reaction has taken place because electrons have been transferred from the iron nail to the copper ions in the solution, allowing solid copper to form. The full equation for the reaction is as follows:

$$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$$

Since copper sulfate and iron sulfate are in solution, we can write an equation for this reaction showing all of the ions present.

$$Fe\,(s) + Cu^{2+}\,(aq) + SO_4^{2-}\,(aq) \rightarrow Fe^{2+}\,(aq) + SO_4^{2-}\,(aq) + Cu\,(s)$$

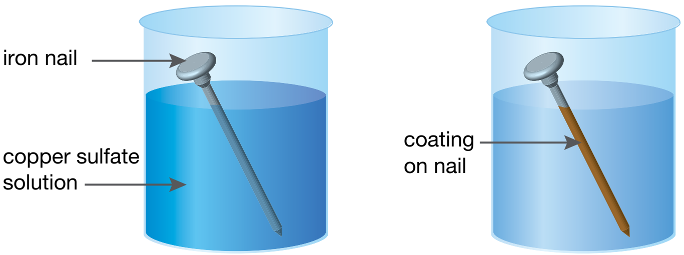
Then eliminating the sulfate spectator ions, the ionic equation becomes:

$$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$$

Although the oxidation and reduction reactions occur simultaneously, it is possible to consider the two reactions separately. To do this we separate the conjugate pair of oxidising agent and reducing agent. A **conjugate redox pair** is made up of two species that differ by a certain number of electrons. Each has its own half-equation.

So for the oxidation conjugate pair, we can write:

FIGURE 14.5 An iron nail placed in copper sulfate solution



$$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$$

and for the reduction conjugate pair, we can write:

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

These half-equations are balanced with respect to both atoms and charge. Combining these two half-equations will yield the ionic equation already provided for the reaction as a whole. The conjugate redox pairs are:

$$Fe^{2+}(aq)/Fe(s)$$
$$Cu^{2+}(aq)/Cu(s)$$

TIP: To remember which sides of the half-equations the electrons are on:

$$\overset{+e^-}{RED}/\overset{+e^-}{OX}$$

This shows the electrons are on the left for reduction and the right for oxidation.

SAMPLE PROBLEM 2

For the following reaction, write the:

a. **ionic equation**
b. **balanced half-equation**
c. **conjugate pairs.**

$$\mathbf{2NaBr(aq) + Cl_2(g) \rightarrow 2NaCl(aq) + Br_2(l)}$$

Teacher-led video: SP2 (tlvd-0577)

THINK	WRITE
a. 1. Show all ions and other species present in the full equation.	$2Na^+(aq) + 2Br^-(aq) + Cl_2(g) \rightarrow 2\,Na^+ + 2Cl^-(aq) + Br_2(l)$
2. Write the ionic equation, omitting spectator ions (in this case, Na^+).	$2Br^-(aq) + Cl_2(g) \rightarrow 2Cl^-(aq) + Br_2(l)$
b. 1. Write the half-equation for the first reactant, *balancing the charges* by adding electrons to the appropriate side.	$2Br^-(aq) \rightarrow Br_2(l) + 2e^-$
2. Write the half-equation for the second reactant, *balancing the charges* by adding electrons to the appropriate side.	$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$
c. Identify the conjugate pairs from each equation.	$Br_2(l)/Br^-(aq)$ $Cl_2(g)/Cl^-(aq)$

PRACTICE PROBLEM 2

For the following equation, write the:

- **a. ionic equation**
- **b. balanced half-equation**
- **c. conjugate pairs.**

$$2AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$

Resources

Digital document Experiment 14.1 Simple redox equations (doc-30862)

Writing more complex redox reactions

At first glance, the statement that MnO_4^- is reduced to Mn^{2+} seems incorrect because, according to the charges on the ions, there appears to be a loss, rather than gain, of electrons:

$$MnO_4^-(aq) \rightarrow Mn^{2+}(aq)$$

It is only when the entire half-equation for the change is written that its true nature as reduction becomes obvious, with five electrons being accepted by each permanganate ion:

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$$

The steps involved in obtaining such equations in an acidified solution are provided here.

These equations *must* be split into two half-equations and then these half-equations need to be balanced before being added together.

To balance half-equations: balance all elements except hydrogen and oxygen. Next, balance the oxygen atoms by adding water molecules, balance the hydrogen atoms by adding hydrogen ions, and balance the difference in charge by adding electrons. Multiply each half-equation by factors that will lead to the same number of electrons in each half-equation.

In summary, to balance each separate half-equation, use the following KOHES procedure:

1. Balance the **K**ey element; that is, the one undergoing oxidation or reduction.
2. Balance the **O**xygen atoms by adding water molecules.
3. Balance the **H**ydrogen atoms by adding hydrogen ions.
4. **E**qualise the charges on both sides of the equation by adding electrons to the more positive side.
5. Add symbols of **S**tate.

on Resources

Video eLesson Balancing redox equations (eles-2489)

SAMPLE PROBLEM 3

Write the half-equation for the reduction of permanganate ions, $MnO_4^-(aq)$, to manganese ions, $Mn^{2+}(aq)$.

THINK	WRITE
1. To balance the key element, state the conjugate pair and write it as a skeleton equation, then balance the equation, except for oxygen and hydrogen, which will be balanced next.	$MnO_4^-(aq)/Mn^{2+}(aq)$ $MnO_4^- \rightarrow Mn^{2+}$ Manganese atoms are already balanced.
2. Balance oxygen atoms, where needed, by adding water.	$MnO_4^- \rightarrow Mn^{2+} + 4H_2O$
3. Balance hydrogen atoms, where needed, by adding H^+.	$MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O$
4. Equalise the charge of each half-equation by adding electrons. The total charge on the left is $-1 + 8 = +7$. On the right side the total charge is $+2$. So, five electrons must be added to the left side. **TIP:** Take care when balancing electrons because this is commonly where mistakes are made.	$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O$
5. Add states; these reactions generally occur in aqueous solutions so the states are usually aqueous but be aware of common gases (g) and water, which is (l).	$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

PRACTICE PROBLEM 3

Write the half-equation for the oxidation of sulfur dioxide gas, SO_2, to sulfate ions, SO_4^{2-}, in solution.

The method described here is used to determine one of the two half-equations required to write an overall ionic equation for a redox reaction. The same method can be used to find the other half-equation; in this case, the reduction reaction. If the number of electrons is different in each half-equation, it is necessary to multiply one or both by a factor before adding them together. Electrons should never be in the final overall ionic equation, however, and nor should any spectator ions.

If water molecules or hydrogen ions are on both sides, cancel the same number of each from both sides to get the lowest number. For example, for the reaction of dichromate ions, $Cr_2O_7^{2-}$, and nitrite ions, NO_2^-, to form chromium, Cr^{3+}, and nitrate ions, NO_3^-, the following half-equations are obtained.

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) \quad [1]$$

$$NO_2^-(aq) + H_2O(l) \rightarrow NO_3^-(aq) + 2H^+(aq) + 2e^- \quad [2]$$

There are 6e⁻ in the first equation and 2e⁻ in the second. To ensure that the electrons cancel out when these equations are added, the second equation must be multiplied by three:

$$3NO_2^-(aq) + 3H_2O(l) \rightarrow 3NO_3^-(aq) + 6H^+(aq) + 6e^- \quad [3]$$

Now equations [1] and [3] can be added together:

$$Cr_2O_7^{2-}(aq) + \overset{8}{\cancel{14}}H^+(aq) + \cancel{6e^-} + 3NO_2^-(aq) + \cancel{3H_2O(l)} \rightarrow$$
$$2Cr^{3+}(aq) + \overset{4}{\cancel{7}}H_2O(l) + 3NO_3^-(aq) + \cancel{6H^+(aq)} + \cancel{6e^-}$$

By cancelling out common terms — that is, $2e^-$, $3H_2O(l)$ and $6H^+(aq)$ — we are left with the final ionic equation:

$$Cr_2O_7^{2-}(aq) + 8H^+(aq) + 3NO_2^-(aq) \rightarrow 2Cr^{3+}(aq) + 4H_2O(l) + 3NO_3^-(aq)$$

on Resources

Digital document Experiment 14.2 Complex redox reactions (doc-30866)

Teacher-led video Experiment 14.2 Complex redox reactions (tlvd-0636)

14.2 EXERCISE

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1. State two situations where redox reactions might occur.
2. Identify each of the following half-equations as either oxidation or reduction.
   (a) $K(s) \rightarrow K^+(aq)$
   (b) $Cu^{2+}(aq) \rightarrow Cu(s)$
   (c) $Br^-(aq) \rightarrow Br_2(l)$
   (d) $H_2(g) \rightarrow H^+(aq)$
   (e) $Sn^{4+}(aq) \rightarrow Sn^{2+}(aq)$

3. Complete the following for the redox reaction $Pb(NO_3)_2(aq) + Zn(s) \rightarrow Zn(NO_3)_2(aq) + Pb(s)$.
   (a) Write the ionic equation.
   (b) Write the two half-equations.
   (c) Name the substance that has been oxidised.
   (d) Name the substance that has been reduced.
   (e) State the oxidising agent.
   (f) State the reducing agent.
4. The equation for the reaction between magnesium and hydrochloric acid is

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

   This is actually a redox reaction and not an acid–base reaction. Explain why it does not fit the definition of acid–base and how it can be correctly described as a redox reaction. Identify the substances oxidised and reduced, and include the half-equations to support your explanation.
5. The balanced ionic equation for the displacement of silver from an aqueous silver nitrate solution by metallic lead is as follows.

$$2Ag^+(aq) + Pb(s) \rightarrow 2Ag(s) + Pb^{2+}(aq)$$

   (a) Write balanced oxidation and reduction half-equations for the reaction.
   (b) Which reactant accepts electrons?
   (c) Which reactant is oxidised? What is it oxidised to?
   (d) Which reactant is the reducing agent?
   (e) From which reactant are electrons taken?
   (f) State the conjugate pairs.
6. The reaction of aluminium with hydrogen ions in a dilute solution of hydrochloric acid can be represented by the following half-equations:

$$Al(s) \rightarrow Al^{3+}(aq) + 3e^-$$
$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

   (a) Which reactant accepts electrons?
   (b) Which reactant is reduced? What is it reduced to?
   (c) Which reactant is the reducing agent?
   (d) From which reactant are electrons taken?
   (e) Which reactant is oxidised? What is it oxidised to?
   (f) Which reactant is the oxidising agent?
   (g) State the conjugate pairs.
   (h) Write the balanced ionic equation.
7. Write balanced half-equations for the following reaction and state the conjugate pairs.
   (a) $Ca(s) + Cl_2(g) \rightarrow CaCl_2(s)$
   (b) $Mg(s) + S(s) \rightarrow MgS(s)$
   (c) $Al(s) + 3Ag^+(aq) \rightarrow Al^{3+}(aq) + 3Ag(s)$
8. Balance the following half-equations using the half-equation method (use KOHES, from section 14.2.3, for each).
   (a) $NO_3^-(aq) \rightarrow NO_2(g)$
   (b) $ClO^-(aq) \rightarrow Cl^-(aq)$
   (c) $BrO_3^-(aq) \rightarrow BrO_2(aq)$
9. For the reaction of dichromate ions, $Cr_2O_7^{2-}$, and nitrite ions, NO_2^-, to form chromium, Cr^{3+}, and nitrate ions, NO_3^-, explain how each of the half-equations were obtained.
10. Balance the following equations by first separating them into half-equations.
   (a) $I_2(s) + H_2S(g) \rightarrow I^-(aq) + S(s)$
   (b) $SO_3^{2-}(aq) + MnO_4^-(aq) \rightarrow SO_4^{2-}(aq) + Mn^{2+}(aq)$
   (c) $Cu(s) + NO_3^-(aq) \rightarrow NO(g) + Cu^{2+}(aq)$

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14.3 Oxidation numbers

BACKGROUND KNOWLEDGE

- The use of oxidation numbers to identify if a redox reaction has occurred

14.3.1 Using oxidation numbers in redox reaction investigations

An **oxidation number**, sometimes called oxidation state, is a theoretical number that corresponds to the charge that an atom would have if the compound was composed of ions.

When determining oxidation numbers, they are written *above* the atom and are not to be confused with the charge on an ion. A set of rules can be used to determine the oxidation number of an atom or ion. These are summarised with examples in table 14.1.

TABLE 14.1 Oxidation number rules and examples

Oxidation number rule	Oxidation number example
1. The oxidation number of an element is 0.	$\overset{0}{Zn}, \overset{0}{S}, \overset{0}{O_2}, \overset{0}{N_2}$
2. In monatomic ions, the oxidation number is the same as the charge on the ion.	$\overset{+2}{Mg^{2+}}, \overset{-2}{O^{2-}}, \overset{+1}{Na^{+}}$
3. In all compounds of hydrogen, the oxidation number of hydrogen is +1. Except in metal hydrides, where the oxidation number of hydrogen is –1.	$\overset{+1}{H_2}\overset{-2}{O}, C\overset{+1}{H_4}$ $\overset{+1}{Na}\ \overset{-1}{H}$
4. The oxidation number of oxygen in a compound is usually –2. Except in: • peroxide compounds, where the oxidation number is –1 • compounds with oxygen bonded to fluorine, where the oxidation number is +2.	$\overset{+1}{H_2}\overset{-2}{O}$ Peroxide compounds: $\overset{+1}{H_2}\overset{-1}{O_2}$ Oxygen difluoride: $\overset{+2}{O}\overset{-1}{F_2}$
5. Fluorine always has oxidation number of –1 because it is the most electronegative element.	$\overset{-1}{F}$
6. In a neutral molecule, the sum of all the oxidation numbers must equal 0.	In H_2S, the oxidation number of hydrogen is +1. When adding oxidation numbers, you need to multiply the oxidation number by the number of atoms present, so the total is +2 for the hydrogen atoms. The overall total must add to zero, so the oxidation number of sulfur is –2.

Oxidation number rule	Oxidation number example
7. In a polyatomic ion, the sum of the oxidation numbers must equal the charge on the ion.	The oxidation number of the chromate ion, $Cr_2O_7^{2-}$, is –2. If the oxidation number of oxygen is –2, the total is –14 so the oxidation numbers of the two chromium atoms must add up to 12 so that the overall total is –2. Each chromium atom must have an oxidation number of +6; that is, $\overset{+6}{Cr}_2\overset{-2}{O}{}_7^{2-}$.
8. In ionic compounds, it is often easier to separate the compound into its ions first and calculate the oxidation numbers of each ion separately.	$NaCl \rightarrow Na^+ + Cl^-$ $\overset{+1}{Na}, \overset{-1}{Cl}$
9. In covalent compounds that do not involve oxygen or hydrogen, the more electronegative element has the negative oxidation number. This is equal to the charge that it would have if it were a negative ion.	In ICl_3, the chlorine is the more electronegative atom. It is, therefore, assigned an oxidation number of –1, because this is the charge on a chloride ion. (*Note:* this is just the way the oxidation number is worked out. This molecule is a covalent, neutral molecule; it *does not* contain chloride ions.) Using rule 6, we can now calculate that the oxidation number of the iodine in ICl_3 is +3.

SAMPLE PROBLEM 4

Determine the oxidation numbers of each element in the following examples.

a. Br_2 **b.** Cu^{2+} **c.** Na_2O **d.** HSO_4^-

THINK	WRITE
For each of the examples determine if the species is an element, monatomic ion, molecule or polyatomic ion.	
a. Br_2 is an element, so its oxidation number is 0. (Rule 1)	$\overset{0}{Br_2}$
b. Cu^{2+} is a monatomic ion, so its oxidation number will be the charge on the ion. (Rule 2)	$\overset{+2}{Cu}{}^{2+}$
c. Na_2O is a compound, so the oxidation numbers must add up to zero. (Rule 6)	Na_2O $\overset{+1}{Na}_2\overset{-2}{O}$ because $(2 \times (+1)) + (-2) = 0$
d. HSO_4^- is a polyatomic ion, so the oxidation numbers must add up to the charge, which in this case is –1. The oxidation numbers of hydrogen and oxygen are known, so it is necessary to find the oxidation number of sulfur, S. (Rule 7) Then the oxidation numbers can be stated.	$\overset{+1}{H}S\overset{-2}{O}{}_4^-$ Let oxidation number of S be x $+1 + x + (4 \times -2) = -1$ $+1 + x - 8 = -1$ $x - 7 = -1$ $x = +6$ The oxidation number of S is +6. $\overset{+1}{H}\overset{+6}{S}\overset{-2}{O}{}_4^-$

PRACTICE PROBLEM 4

Find the oxidation numbers of the following elements.

a. **P_2O_5**
b. **NO_3^-**
c. **Cl^-**
d. **K**

Metallic elements always have positive oxidation numbers. Non-metallic elements can have positive or negative oxidation numbers, and the highest oxidation an element can have relates to its group in the periodic table.

The use of oxidation numbers greatly simplifies the identification of redox reactions and the subsequent identification of oxidising agents and reducing agents. If oxidation numbers change during a reaction, it indicates that the reaction is a redox reaction.

If an oxidation number increases, oxidation has occurred.

If an oxidation number decreases, reduction has occurred.

Consider the reaction $Ca(s) + H_2O(l) \rightarrow Ca(OH)_2(s) + H_2(g)$.
We can assign oxidation numbers to all of the atoms present.

$$\overset{0}{Ca}(s) + \overset{+1}{H_2}\overset{-2}{O}(l) \rightarrow \overset{+2}{Ca}(\overset{-2}{O}\overset{+1}{H})_2(s) + \overset{0}{H_2}(g)$$

Firstly, the oxidation numbers have changed, so this is a redox reaction. The calcium atom has *increased* its oxidation number from 0 to +2; therefore, it has been oxidised and it is the reducing agent. The hydrogen atoms have *decreased* their oxidation numbers from +1 to 0; therefore, the water has been reduced and it is the oxidising agent. Note that because the hydrogen is present as part of the water molecule, we say that the water is reduced.

SAMPLE PROBLEM 5

Use oxidation numbers to explain why the following reaction is a redox reaction and state the reducing agent and the oxidising agent.

$$\mathbf{MnO_2(s) + 2H^+(aq) + NO_2^-(aq) \rightarrow NO_3^-(aq) + Mn^{2+}(aq) + H_2O(l)}$$

Teacher-led video: SP5 (tlvd-0580)

THINK	**WRITE**
1. Assign oxidation numbers to all atoms. This is a redox reaction because the oxidation numbers of reactants have changed.	$\overset{+4}{Mn}\overset{-2}{O_2}(s) + \overset{+1}{2H^{+1}}(aq) + \overset{+3}{N}\overset{-2}{O_2^-}(aq) \rightarrow \overset{+5}{N}\overset{-2}{O_3^-}(aq) + \overset{+2}{Mn^{2+}}(aq) + \overset{+1}{H_2}\overset{-2}{O}(l)$
2. Identify which atom has increased its oxidation number; this belongs to the species being oxidised.	$\overset{+3}{N}\overset{-2}{O_2^-}$ has been oxidised to $\overset{+5}{N}\overset{-2}{O_3^-}$.

3. Identify which atom has decreased its oxidation number; this belongs to the species being reduced. $\overset{+4}{\mathrm{Mn}}\overset{-2}{\mathrm{O}}_2$ has been reduced to $\overset{+2}{\mathrm{Mn}}{}^{2+}$.
4. The species being oxidised is the reducing agent. NO_2^- is the reducing agent.

   The species being reduced is the oxidising agent. MnO_2 is the oxidising agent.

PRACTICE PROBLEM 5

Use oxidation numbers to explain why the following reaction is a redox reaction and state the reducing agent and the oxidising agent.

$$\mathbf{Pb(s) + Br_2(l) \rightarrow Pb^{2+}(aq) + 2Br^-(aq).}$$

REDOX REACTIONS IN FIREWORKS

Fireworks are spectacular examples of exothermic redox reactions.

To create the redox reactions that occur in fireworks, clearly an oxidising agent and a reducing agent are required. Traditionally, the combination used was gunpowder, which is composed of an oxidising agent (potassium nitrate, also known as saltpetre), sulfur and carbon (in the form of charcoal). The mass ratio of these three substances in gunpowder, which has remained unchanged for over 500 years, is 75:15:10.

Today the oxidising agents used include nitrates, NO_3^-, chlorates, ClO_3^-, and perchlorates, ClO_4^-. The chlorine in the chlorates help strengthen some colours and, when burnt, potassium perchlorate, $KClO_4$, or ammonium perchlorate, NH_4ClO_4, can reach temperatures of up to 2000°C. Nitrate oxidising agents include potassium nitrate, KNO_3 (saltpetre); strontium nitrate, $Sr(NO_3)_2$, which produces a red colour; and barium nitrate $Ba(NO_3)_2$, which produces a green colour. You may remember seeing these colours when you performed flame tests on metal compounds.

The reducing agents are metals (which have a high heat output when oxidised), carbon (in the form of charcoal or organic compounds) and/or sulfur. The high temperatures of reaction of the oxidising agents allow a simultaneous release of a large amount of oxygen that is rapidly accepted by the reducing agents, oxidising (burning) the fuel, which releases more heat and gas (SO_2 and CO_2) and further propagating the reaction to produce explosive energy and gas — what we see as spectacular explosions of colour.

While elements in the oxidising and reducing agents contribute to the colours seen (for example, burning aluminium creates a white light), the variety and intensity of colours in fireworks is also due to the addition of colour ingredients and intensifiers. Red is created by strontium compounds, orange by calcium, yellow by sodium, green by barium, blue by copper, purple by copper-strontium and silver by magnesium-aluminium compounds.

14.3 EXERCISE

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1. Determine the oxidation numbers for each atom in the following examples.
   (a) P
   (b) F_2
   (c) Zn^{2+}
   (d) HCO_3^-
   (e) $NaNO_3$
   (f) BeO
   (g) Ag_2S
2. Arrange these substances in order of the increasing oxidation numbers of nitrogen: NO_3^-, N_2, NH_4^+, NO_2.
3. Sulfur is in group 16. What is the maximum oxidation number that sulfur can have?
4. Use oxidation numbers to identify which of the following equations are redox reactions.
   (a) $Br_2(l) + SO_2(g) \rightarrow 2Br^-(aq) + SO_4^{2-}(aq)$
   (b) $KOH(aq) + HNO_3(aq) \rightarrow KNO_3(aq) + H_2O(aq)$
   (c) $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$
   (d) $5Zn(s) + 2NO_3^-(aq) + 12H^+(aq) \rightarrow 5Zn^{2+}(aq) + N_2(g) + 6H_2O(l)$
5. In the unbalanced chemical reactions given, complete the following.
   (i) Determine the oxidation number of each element.
   (ii) State which species is oxidised and which is reduced.
   (iii) State the oxidising agent and the reducing agent.
   (a) $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$
   (b) $PbO_2(s) + I^-(aq) \rightarrow Pb^{2+}(aq) + I_2(g)$
   (c) $I^-(aq) + NO_2^-(aq) \rightarrow I_2(s) + NO(g)$
   (d) $H_2O_2(aq) + MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + O_2(g)$
6. For the reaction $H_2O_2(l) + Sn^{2+}(aq) + 2H^+ \rightarrow 2H_2O(l) + Sn^{4+}(aq)$, use oxidation numbers to answer the following questions.
   (a) What is the oxidation number of oxygen in H_2O_2?
   (b) What substance is oxidised?
   (c) What substance is reduced?
   (d) Name the oxidising agent?
   (e) Name the reducing agent?
7. Determine the oxidation numbers of the elements in the following changes. Place a tick in the relevant box to indicate if the changes are oxidation, reduction or neither.

Change	Oxidation	Reduction	Neither
$O_2 \rightarrow O_3$			
$MnO_2 \rightarrow MnO_4^-$			
$P_2O_5 \rightarrow P_4H_{10}$			
$NH_3 \rightarrow NO_2$			
$Fe^{2+} \rightarrow Fe^{3+}$			
$S_2O_3^{2-} \rightarrow S_2O_4^{2-}$			

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14.4 Reactivity series of metals

KEY CONCEPT

- The reactivity series of metals and metal displacement reactions, including balanced redox equations with states indicated

You may have noticed that many of the redox reaction examples involved a metal reacting with a different metal ion solution. Metals have different levels of reactivity, not only with other metal solutions but also with oxygen, water or acids.

14.4.1 Reactivity of metals

FIGURE 14.6 The flask on the right shows the initial set up. The flask on the left shows the silver forming on the copper strip and the solution has become blue due to the copper ions forming.

When a copper strip is placed in silver nitrate solution, a reaction occurs, causing silver metal to form on the copper (see figure 14.6). This is a **displacement reaction**: copper removes silver from the solution and, as a result, a deep blue colour of the solution forms as the copper metal forms copper ions. If you place a silver strip, however, in a solution of copper nitrate, no reaction occurs. Therefore, copper is more reactive than silver.

Different metals have varying abilities to react with other substances. Potassium is so reactive that it is only found naturally as a compound rather than a pure metal. When extracted, it must be stored in oil to prevent exposure to moisture and oxygen. It reacts very vigorously with water to produce hydrogen gas and potassium hydroxide. The heat given off in this exothermic reaction is sufficient to ignite the hydrogen gas and allows it to burn in air with a bright lilac flame (see figure 14.7).

$$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$$

Gold, on the other hand, is very unreactive. It occurs in nature as a pure element and can be found in its pure form both by mining or panning in rivers. Ornamental gold is valuable because it remains untarnished for centuries. Potassium and gold are at opposite ends of the **reactivity series of metals**, shown in table 14.2. This series, sometimes called the activity series, lists the metals in order of how readily they react with oxygen, water, steam, dilute acids and salt solutions, and is useful for predicting reactions.

Each of these reactions is a redox reaction, in which the metal is always the reducing agent. A more reactive metal will displace a less reactive metal from solution. In other words, a more reactive metal gives up its electrons more readily and is a stronger reducing agent. Reactivity of metals tends to increase down a group of the periodic table.

A mnemonic using the first letters of the names of the metals can be used to remember the beginning of this list:

Please	Potassium
Stop	Sodium
Calling	Calcium
My	Magnesium
Aunt’s	Aluminium
Zebra	Zinc
In	Iron
This	Tin
Lesson	Lead

FIGURE 14.7 Potassium reacting with water

TABLE 14.2 Reactivity series of metals.

Element	Appearance of metal	Reaction with oxygen	Reaction with water or steam	Displacement of metals	Reaction with dilute acids
K	Dull; stored under oil	Oxidises in air at room temperature to give oxides	Hydrogen formed from cold water	Any metal in this series displaces any other that occurs lower in the series (i.e. change the metal ion back to the metal). The further apart the metals are, the easier the displacement.	Hydrogen formed violently
Na					
Ca	Generally dull				Hydrogen formed with cold acid
Mg			Hydrogen formed with steam		
Al					
Zn					
Fe					
Sn		Oxidises when heated in air or oxygen to give oxides	No reaction at Bunsen burner temperatures		
Pb					No visible reaction
Cu	Generally shiny		No reaction		No reaction
Hg					
Ag		No reaction			
Au					

Note: aluminium metal often appears to be less reactive than the reactivity series indicates because it has a coating of aluminium oxide and this protects it from reacting further with oxygen. This coating can be thickened by a process called anodising.

In general, a reactive metal is:

- a stronger reducing agent
- more easily oxidised or corroded
- able to donate electrons readily
- able to form a positive ion easily
- more likely to be found naturally as a compound rather than a pure metal
- more difficult to extract from its ore and so requires more energy.

FIGURE 14.8 Reactivity of metals with hydrochloric acid; from left to right: calcium, magnesium, zinc, copper

14.4.2 Metal displacement reactions

The reactivity series can be established experimentally. Metals can be combined with oxygen, water, acids or other metal solutions. The most direct method to determine the relative reactivity of metals is to use displacement reactions.

To understand how displacement reactions may be used, consider the example earlier in this section, where copper displaced silver ions from solution to form silver metal and copper ions according to the following equation.

$$Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$$

Silver metal does not displace Cu from solution. This demonstrated difference in reactivity forms the basis of the displacement method of determining relative reactivity. Samples of each metal are immersed in separate solutions of metal cations. These cations accept electrons from any other more reactive solid metal atoms that are introduced, forming solid metal atoms themselves. The more reactive metal, therefore, loses electrons and forms ions in solution.

SAMPLE PROBLEM 6

Use the reactivity series to consider what happens when a piece of lead is dipped into a solution of silver nitrate.

a. Write an ionic equation.
b. Identify two half-equations for the reaction.
c. Determine the oxidising agent and reducing agent.

Teacher-led video: SP6 (tlvd-0581)

THINK	WRITE
a. Recognise that lead is above silver in the activity series. Because lead is more reactive than silver, a displacement reaction will occur between lead metal and silver ions. Lead displaces silver from solution, producing silver metal. The ionic equation omits the NO_3^- ions.	$Pb(s) + 2Ag^{+}(aq) \rightarrow Pb^{2+}(aq) + 2Ag(s)$
b. The half-equation for oxidation has lead and its conjugate. The half-equation for reduction has silver ion and its conjugate.	$Pb(s) \rightarrow Pb^{2+}(aq) + 2e^{-}$ $Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$
c. The substance that is oxidised, Pb, is the reducing agent. The substance that is reduced, Ag^{+}, is the oxidising agent. **TIP:** Remember to call this the silver ion and not just silver.	Pb is oxidised to Pb^{2+} so it is the reducing agent. Ag^{+} is reduced to Ag so it is the oxidising agent.

PRACTICE PROBLEM 6

Use the reactivity series to consider what happens when a piece of zinc is dipped into a solution of lead nitrate.

a. Write an ionic equation.
b. Identify two half-equations for the reaction.
c. Identify the oxidising agent and reducing agent.

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Weblinks Reaction of metals with oxygen and water
Displacement reactions

14.4 EXERCISE

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1. Refer to the activity series of metals in table 14.2 and name the metals that best meet the following.
   (a) Do not react at all with oxygen
   (b) Will displace tin ions but will be displaced by zinc
   (c) React with cold water to form hydrogen gas
   (d) Are generally dull and oxidise when heated with oxygen to give oxides
   (e) Displace hydrogen from cold water but not reactive enough to be stored under oil
   (f) Oxidise when heated in oxygen but are otherwise generally shiny
2. (a) In which part of the periodic table would you find the most reactive metals? Suggest a reason this is the case.
   (b) Explain the trend in terms of reactivity down a group of the periodic table.
3. (a) State three observations that you would expect to see if iron was placed in a copper sulfate solution and left overnight.
   (b) State the products.
4. Use the reactivity series to determine if a reaction occurs when the following are mixed. Write an equation for any reaction that occurs and identify the oxidising agent and the reducing agent.
   (a) $Mg(s) + Cu^{2+}(aq)$
   (b) $Al(s) + Pb^{2+}(aq)$
   (c) $Cu(s) + Fe^{2+}(aq)$
   (d) $Ag(s) + Zn^{2+}(aq)$
5. Use the reactivity series to predict what reaction, if any, occurs when the following reagents are mixed. Write half-equations and overall equations for any reaction that you predict.
   (a) $Pb^{2+}(aq)$ and $Fe(s)$
   (b) $Sn^{2+}(aq)$ and $Zn(s)$
   (c) $Al^{3+}(aq)$ and $Fe(s)$
   (d) $Zn^{2+}(aq)$ and $Al(s)$
6. Sharni used the reactivity series to predict that a strip of aluminium metal would react with copper(II) sulfate. When she tried the experiment, she observed no reaction.
   (a) Explain her observations.
   (b) What steps would she need to take to prove that her prediction was correct?
7. Write a full equation showing the reaction of sodium with water. State the oxidising agent and the reducing agent.
8. Name the product of the reaction of aluminium with oxygen. Write a balanced equation showing this reaction.
9. Identify the products and write the equations for the following reactions.
   (a) Magnesium and hydrochloric acid
   (b) Magnesium and steam
   (c) Magnesium burned in air

10. Matthew recommended that the next set of coins for Australia be made from gold, Oliver recommended the use of magnesium and Chris suggested that copper would be the best choice. Do you agree with any of these recommendations or do you have a completely different proposal? Include a discussion of the factors that need to be considered when producing coins.

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14.5 Corrosion

KEY CONCEPT

- The causes and effects of a selected issue related to redox chemistry

14.5.1 Types of corrosion

Although the ability of metals to act as reducing agent and be oxidised can be useful in cases like batteries, in other situations redox reactions can be quite a problem. **Corrosion** is the oxidation of metals by materials in their environment. When a metal corrodes, it suffers damage as a result of chemical change. Two types of corrosion are possible:

FIGURE 14.9 Car with corrosion above the wheel

1. **Dry corrosion** occurs when a metal reacts directly with oxygen. Some metals are very reactive and must be stored under oil. Potassium and sodium combine readily with oxygen, so they undergo dry corrosion very quickly. Iron undergoes dry corrosion, but at normal temperatures its rate of dry corrosion is very slow.
2. **Wet corrosion** occurs when a metal reacts directly with oxygen in the presence of water. Water can speed up some corrosion reactions, such as that of iron, which corrodes much more quickly in a moist environment. Corrosion returns metals to their natural states — the original ores. Except for gold and platinum, which are virtually unreactive, most metals are found in the ground as oxidised metals or ores. Since corroded metal often loses its structural purpose and attractiveness, this oxidation reaction has a great economic impact. In Australia, billions of dollars are spent each year preventing and treating corrosion.

Corrosion is the breakdown of a substance, often a metal, due to its reaction with chemicals in their environment.

14.5.2 Wet corrosion process

The wet corrosion of iron is commonly known as **rusting**. Iron rusts when a water drop containing dissolved impurities lands on it. Oxidation usually takes place at a 'stress site' on the iron, such as a dent in a car or a scratch on the surface. Iron is oxidised, releasing electrons.

$$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$$

These electrons travel through the metal and are accepted by oxygen, usually at the point where the edge of a water drop is in contact with the air. Oxygen can also come from dissolved oxygen in the water drop. Oxygen accepts electrons and is reduced.

$$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$$

Fe^{2+} and OH^- ions produced migrate toward each other and react to produce an insoluble green iron(II) hydroxide precipitate. The overall equation is as follows.

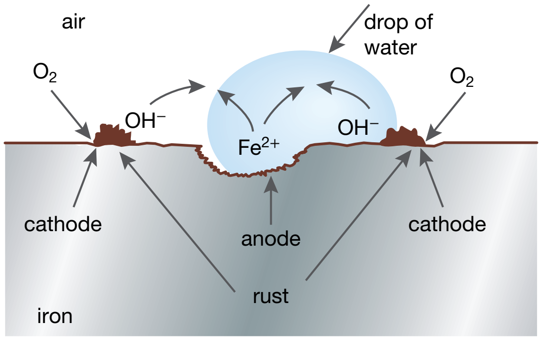
$$2Fe(s) + O_2(g) + 2H_2O(l) \rightarrow 2Fe(OH)_2(s)$$

Iron(II) hydroxide is very unstable in air and quickly reacts with oxygen to produce brown iron(III) hydroxide.

$$4Fe(OH)_2(s) + O_2(g) + 2H_2O(l) \rightarrow 4Fe(OH)_3(s)$$

This brown precipitate then partially dehydrates (loses water) to produce iron(III) oxide, $Fe_2O_3.xH_2O$, where x is less than 3. The degree of hydration (the value of x) affects the colour of the rust.

FIGURE 14.10 The corrosion of iron

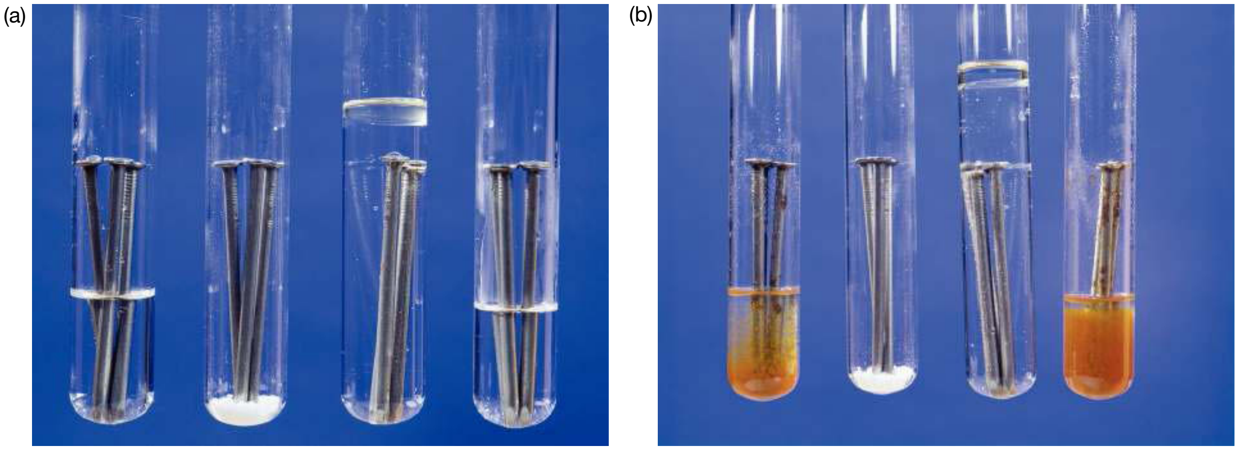


This process is responsible for the corrosion of cars, tools, bridges, buildings, machinery and virtually anything made of iron and steel that is not protected in some way. This is significant because iron is our most commonly used metal. Because it is soft and flaky, rust is easily dislodged, which accelerates further rusting. Water is vital in the process; without a 'salt bridge', iron does not rust. Cars last much longer in dry areas; however, the dissolved salts in the air in seaside areas accelerate rusting. The dissolved ions increase the conductivity of the moisture and, therefore, encourage corrosion. Other factors increase corrosion, including the presence of acids and pollutants such as sulfur dioxide and nitrogen dioxide, and imperfections in the metal surface.

Testing the conditions for rusting

The conditions for rusting can be tested experimentally, as demonstrated in figure 14.11. In the first image (a), from left to right, iron nails are placed in water and air, a drying agent, distilled water topped with oil, and salty water. The results after several weeks are shown in the second image (b). In the test tube with water and air (far left), rust (hydrated iron oxide) is seen; this process has been accelerated in the test tube with salt water and air (far right). The nails exposed to air but no water and exposed to water but no air (oxygen) have not rusted (middle test tubes).

FIGURE 14.11 Conditions for rusting; left to right: water and air, drying agent, distilled water, and oil and salty water.



Adverse effects of corrosion

Corrosion has many adverse effects, including the following.

- Metal oxides, or rust, have less tensile strength and less elasticity than uncorroded metal. Corroded buildings, bridges and machinery may develop weaknesses that result in malfunction or breakage.
- A severely corroded metal cannot conduct electricity. Products of corrosion formed in a car battery, between the terminals and the leads, may cause the electrical system to malfunction.
- Corroded copper pipes and hot water tanks leak as the copper compounds dissolve in the water flowing through them.
- Products of corrosion are brittle and flake off, resulting in holes in car mufflers or iron roofing. In car radiators, iron oxide flakes may cause blockages.
- A corroded metal occupies a larger volume than the original metal. Corroded nuts and bolts may jam machinery.

14.5.3 Corrosion protection

It is important to protect iron and steel from corrosion. Three main methods are used: surface protection (as described in topic 3), **alloying** and **electrochemical protection**.

Surface protection

If a metal surface can be prevented from coming into contact with oxygen or water, it will not corrode. An obvious way of achieving this is to coat the surface in some way. Commonly used surface coatings are plastic, paint, grease or oil, noble metals and sacrificial metals. Specific details of each of these methods may be found in topic 3.

A metal that produces its own surface protection is aluminium and it does this by corroding. It is an unusual metal because even though it is a reactive metal and reacts readily with oxygen in the atmosphere this produces an oxide layer that provides a thin, hard surface coating and prevents further reaction with the environment. This contrasts with iron where the rust flakes off exposing more iron to the air.

Alloying

Iron can be alloyed with small quantities of metals, such as chromium, nickel, manganese and molybdenum, to produce stainless steel. A wide variety of grades of stainless steel is produced for specific purposes. The most widely used stainless steel contains approximately 18% chromium and 8% nickel. This is called 18-8 stainless steel and is used for kitchen sinks and cutlery. The formation of a stable film of chromium (III) oxide, Cr_2O_3, provides a very strong surface protection for this alloy, although the metal may still be attacked by chloride ions. It is important, therefore, to choose an appropriate alloy for any specific use. For example, stainless steel containing molybdenum is more resistant to chloride ion attack and is, therefore, recommended for use in sea water; it is used in valves on scuba diving equipment.

Electrochemical protection

Electrochemical protection involves placing a more reactive metal in electrical contact with the metal that needs protection. It involves 'sacrificing' one metal to save another.

If a block of a metal higher in the reactivity series is connected to iron, that metal is corroded while iron remains intact. Underground pipes may be protected by attaching bags of magnesium scraps at intervals and replacing them when they have been corroded away. The hulls of ships can be protected by attaching blocks of zinc, which are sacrificed to protect the iron. This is sometimes referred to as **sacrificial protection** or **cathodic protection**. It involves the selection of an appropriate sacrificial anode. A **sacrificial anode** is a reactive metal that is used to prevent a less reactive metal from corroding.

A demonstration of cathodic protection can be seen in figure 14.12. Identical iron nails are introduced to two beakers of agar, each containing phenolphthalein indicator and potassium hexacyanoferrate, $K_4Fe(CN)_6$. The first nail (a) is unprotected, while the second nail (b) has magnesium ribbon wrapped around its centre. The pink colouring indicates rusting. For the second nail (b), the magnesium wrapped around the nail protects the nail from corrosion because magnesium is a more reactive metal than iron, meaning it will corrode preferentially, leaving the iron intact.

FIGURE 14.12 A demonstration of cathodic protection with (a) an unprotected nail and (b) nail with magnesium ribbon wrapped around its centre; the pink colour indicates rusting

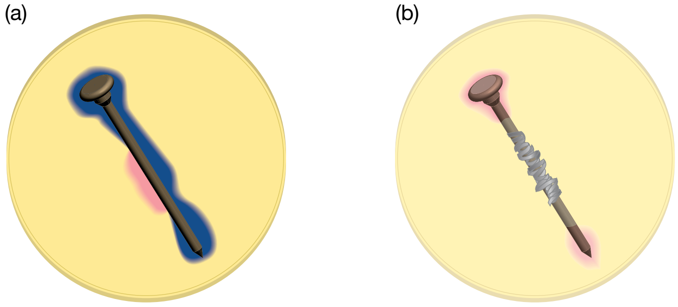


FIGURE 14.13 Corrosion of an iron pipe

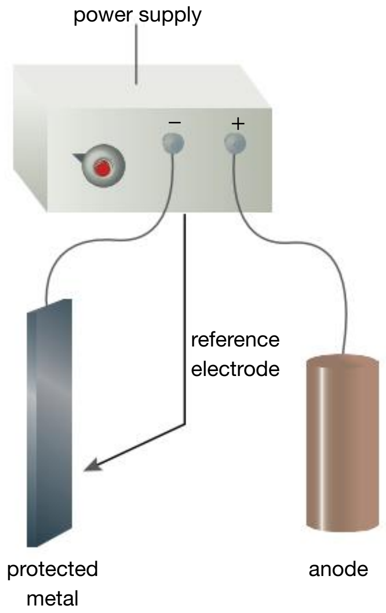
FIGURE 14.14 Sacrificial anode for cathodic protection on a pipe

Iron pipes exposed to the atmosphere rust very quickly. Even pipes buried beneath the ground rust. To prevent corrosion, blocks of a more reactive metal, such as magnesium, are bolted to aboveground pipes.

For underground pipes, bags of magnesium can be placed in holes in the ground and then connected to the pipe by a conductor. When the magnesium has corroded away, it is replaced.

Another method of cathodic protection is impressed-current cathodic protection. This is achieved by making iron the cathode that supplies electrons to the protected metal by connecting it to the negative terminal of a battery. A conductor such as graphite is connected to the positive terminal (see figure 14.15). The current is monitored by a reference electrode that keeps the potential difference between the cathode and the anode high enough to ensure protection. The formation of Fe^{2+} ions is inhibited by the negative impressed current on the iron.

FIGURE 14.15 Electrochemical protection



Car manufacturers treat newly formed car body shells so that they become resistant to corrosion and attractively coloured. This involves immersing each car body in a series of treatment and paint tanks containing zinc. The body then receives an electrocoat primer on the inside and outside surfaces to provide corrosion protection and prepare the surface for the coloured top-coat. Paint spraying is usually performed by automatic machines using electrostatic attraction. The car body is then heated to produce a hard, glossy finish.

FIGURE 14.16 Car bodies treated to provide corrosion protection

NAPOLEON, CANS AND CORROSION

A common example of corrosion prevention is found in your kitchen cupboard. Tin cans have been used as a method of storage of food items for about 200 years. The history of this method of food preservation is a fascinating one. In 1795 Napoleon needed a way of feeding his army food that did not spoil while they were away fighting, so the French government offered a large reward to anyone who could provide a solution to Napoleon's problem. After all, he was quoted as saying that 'an army marches on its stomach'.

In 1810 Nichols Appert, a sweet maker, ultimately won the prize when he proposed putting food in a container, sealing it and then heating it. It was said that he used his method to preserve a whole sheep. This process, however, was not suitable for all purposes because the material he used was glass and prone to shattering. He had considered using tinplate but the quality of the available material was poor.

At about the same time, Peter Durand patented the first cans for stopping food from becoming rancid. Again, the food was heated and this seemed to effectively preserve the food. Cans were easier to produce at that time than the thick glass used in Appert's method. These 'tins' were made of steel coated in tin to prevent corrosion and the lids were originally soldered on with lead to provide a good seal. Surprisingly, the can opener was not invented until 50 years later so it would have been interesting observing the various methods used to open the cans!

It was not long before lead was no longer used because it was found to be toxic. Today, cans are less likely to be coated in tin, which is not readily available. Today's cans are made from alloys and are lined with an epoxy coating. This plastic lining is sometimes bisphenol-A (BPA), which is a controversial compound. Many studies are investigating the safety of this material as well as the use of less contentious alternatives. The development of cans has made a variety of different foods available to us whenever we want them, and the ring top eliminates the need for finding a can opener in the bottom of a drawer.

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Digital documents Experiment 14.3 Corrosion of iron (doc-30863)

Experiment 14.4 Minimising corrosion (student design) (doc-30864)

14.5 EXERCISE

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1. (a) What is corrosion?
   (b) What type of reaction is involved?
2. Why do some metals corrode and others do not appear to?
3. Explain the difference between dry corrosion and wet corrosion. Give an example of each.
4. What is the difference between rusting and corrosion?
5. What conditions are required for rusting to occur? Describe the process of rusting, including chemical equations.
6. Describe an experiment to test if iron nails rust more quickly in tap water or salty water.
7. If iron is so susceptible to corrosion, why is it used so extensively?
8. Describe two examples where corrosion can cause problems in industry.
9. What sort of corrosion protection would you observe in the following examples?
   (a) A bike chain
   (b) A metal dish-draining rack
   (c) A metal roof
   (d) A ship's hull
10. Explain how sacrificial protection is used to prevent corrosion of pipelines and ships.

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14.6 Review

14.6.1 Summary

Redox reactions

- Redox reactions are very common reactions that occur in our bodies, batteries and during combustion.
- The word ‘redox’ comes from ‘reduction’ and ‘oxidation’.
- The early definition was that oxidation was the gain of oxygen and reduction was loss of oxygen.
- Oxidation–reduction (or redox) reactions are now known as those in which electrons are transferred from the reducing agent to the oxidising agent.
- Oxidation can be defined as the loss of electrons
- Reduction can be defined as the gain of electrons.
- Oxidation and reduction are complementary processes; one cannot occur unless the other occurs simultaneously.
- An oxidising agent can be defined as a substance that:
  - accepts electrons
  - undergoes reduction.
- A reducing agent can be defined as a substance that:
  - donates electrons
  - undergoes oxidation.
- Oxidation–reduction reactions can be represented as the sum of two half-equations — one for oxidation and one for reduction.
- A conjugate redox pair is made up of two species that differ by a certain number of electrons, and each has its own half-equation. These must be balanced to show the relative amounts of oxidising agent and reducing agent such that the number of electrons taken from the reducing agent is the same as the number of electrons accepted by the oxidising agent.

Oxidation numbers

- **Element**: oxidation number is 0; e.g. $\overset{0}{Zn}, \overset{0}{S}, \overset{0}{O_2}, \overset{0}{N_2}$
- **Monatomic ions**: oxidation number is the same as the charge on the ion; e.g. $\overset{+2}{Mg^{2+}}, \overset{-2}{O^{2-}}, \overset{+1}{Na^+}$
- **Hydrogen compounds**: the oxidation number of hydrogen is $+1$; e.g. $\overset{+1}{H_2}\overset{-2}{O}$, $C\overset{+1}{H_4}$.

  Exception: in metallic hydrides, the oxidation number is –1; e.g. $\overset{+1}{Na}\overset{-1}{H}$.
- **Oxygen compounds**: the oxidation number of oxygen is usually -2; e.g. $\overset{+1}{H_2}\overset{-2}{O}$

  Exceptions: in peroxide, the oxidation number of the oxygen is -1; e.g. $\overset{+1}{H_2}\overset{-1}{O_2}$. In compounds bonded to fluorine, the oxidation number of the oxygen is $+2$; e.g. $\overset{+2}{O}\overset{-1}{F_2}$
- **Fluorine** always has oxidation number of –1 because it is the most electronegative element.
- **Neutral molecule:** the sum of the oxidation numbers must equal zero.
- **Polyatomic ion:** the sum of the oxidation numbers must equal the charge on the ion.

- **Ionic compounds**: it is often easier to separate the compound into its ions first and calculate the oxidation numbers of each ion separately.
- **No hydrogen or oxygen**: the more electronegative atom is assigned an oxidation number, as if it were a monatomic ion.
- If an oxidation number increases, oxidation has occurred
- If an oxidation number decreases, reduction has occurred.

Reactivity series of metals

- Metals can be ranked in an approximate order of reactivity, called the reactivity (or activity) series of metals, according to how readily each reacts with oxygen, water, steam, dilute acids and salt solutions.
- Displacement reactions, where a more reactive solid metal displaces the ions of another metal from solution, can be used to determine this ranking.
  - The order of some selected metals in the reactivity series is, from most reactive to least reactive: K, Na, Ca, Mg, Al, Zn, Fe, Sn, Pb, Cu, Hg, Ag, Au.
  - More reactive metals displace the cations of less reactive metals from solution.

Corrosion

- Corrosion is the oxidation of metals by materials in their environment.
- There are two types of corrosion: wet and dry corrosion.
- Rusting is the corrosion of iron by oxygen in a moist environment.
  - Rust can be represented by the formula $Fe_2O_3.xH_2O$ where x is a number from 1 to 3.
  - Rusting requires the presence of oxygen, to act as the oxidising agent, as well as water and dissolved salts, to act as the electrolyte.
- Corrosion of iron occurs where the metal lattice has been deformed.
- Corrosion affects the properties of metals, which can be protected from corrosion by one of three methods:
  - surface protection, which may be achieved using plastic, paint, grease or oil, or through the application of metal coatings, such as noble coating and sacrificial coating
  - alloying
  - electrochemical protection, also known as sacrificial protection or cathodic protection, where a sacrificial anode of another more reactive metal (such as zinc or magnesium) is introduced.
- Common metal coatings include:
  - galvanised iron; this has a sacrificial coating of zinc, which is more reactive than iron. At a break in the sacrificial coating, protection of the steel is maintained.
  - stainless steel, which is produced by alloying and is corrosion resistant because of a protective layer of chromium(III) oxide.
- Aluminium is corrosion resistant because of a tightly bound impermeable layer of aluminium oxide.

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14.6.2 Key terms

alloying forming a substance with metallic properties that consists of two or more elements, at least one of which is a metal

cathodic protection method of protecting a metal from corrosion by connecting it to a power source and another metal that is less active. The electrons feed onto the metal being protected, making it a cathode so that it cannot be oxidised.

conjugate redox pair two species that differ only by a certain number of electrons

corrosion oxidation of metals by materials in their environment

displacement reaction chemical reaction in which a more reactive element displaces a less reactive element from its compound

dry corrosion corrosion of a metal in the air when water is not a reactant

electrochemical protection a method of protecting a metal from corrosion by connecting it to a power source and another metal that is less active. The electrons feed onto the metal being protected, making it a cathode so that it cannot be oxidised.

half-equation chemical equation showing either oxidation or reduction in a redox reaction

oxidation an increase in the oxidation number; a loss of one or more electrons

oxidation number a theoretical number that corresponds to the charge that atom would have if the compound was composed of ions

oxidising agent the substance that causes oxidation of another substance; it accepts electrons and is reduced

oxidant the substance that causes oxidation of another substance; it accepts electrons and is reduced; see *oxidising agent*

product chemical species obtained as the result of a chemical reaction

reactivity series of metals an ordered list of how readily the metals react with oxygen, water, steam, dilute acids and salt solutions

redox describes a reaction that involves the transfer of one or more electrons between chemical species

reducing agent the substance that causes reduction of another substance; it donates electrons and is oxidised

reductant the substance that causes reduction of another substance; it donates electrons and is oxidised; see *reducing agent*

reduction a decrease in the oxidation number; a gain of electrons

rusting wet corrosion of iron

sacrificial anode a reactive metal used to prevent a less active metal from corroding

sacrificial protection protection of a metal from corrosion by attaching another more active metal to it. The more active metal reacts preferentially with corrosive elements and protects the less reactive metal.

wet corrosion corrosion of a metal by reaction with oxygen in the presence of water

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Digital document Key terms glossary — Topic 14 (doc-30944)

14.6.3 Practical work and experiments

Experiment 14.2

Complex redox reactions

Aim: To react some common oxidising agents and reducing agents with one another and write their redox equations

Digital document: doc-30866
Teacher-led video: tlvd-0636

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Digital documents Practical investigation logbook (doc-30945)
Experiment 14.1 Simple redox equations (doc-30862)
Experiment 14.3 Corrosion of iron (doc-30863)
Experiment 14.4 Minimising corrosion (student design) (doc-30864)

14.6 Exercises

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14.6 Exercise 1: Multiple choice questions

1. Oxidation is best described as a process that involves what?
   - **A.** The gain of protons
   - **B.** The loss of oxygen
   - **C.** The gain of electrons
   - **D.** The loss of electrons
2. The process of black and white film photography depends upon a reaction initiated by light and represented by the following equation.

$$2Ag^+(aq) + 2Br^-(aq) \rightarrow 2Ag(s) + Br_2(aq)$$

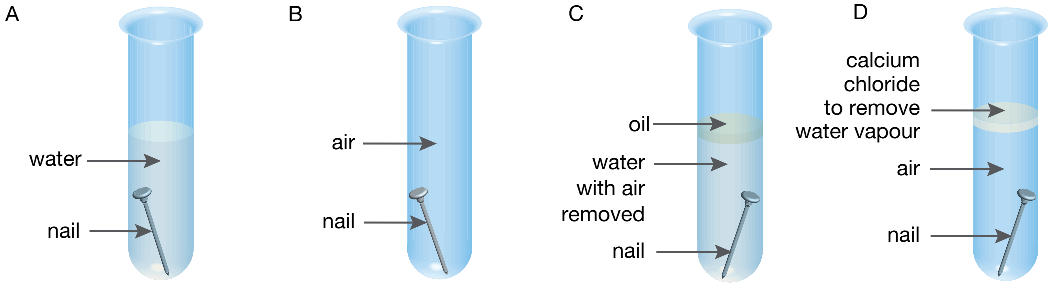
   Name the species that is oxidised in this reaction.
   - **A.** $Ag^+(aq)$
   - **B.** $Br^-(aq)$
   - **C.** $Ag(s)$
   - **D.** $Br_2(aq)$
3. For the reaction $Al \rightarrow Al^{3+} + 3e^-$, which of the following is *true*?
   - **A.** Aluminium loses electrons from its nucleus.
   - **B.** Al is reduced to Al^{3+}.
   - **C.** The electrons remain free in aqueous solution.
   - **D.** This reaction can occur only if another reaction is occurring simultaneously and using electrons.
4. Consider the following half-equation:

$$2Cl^-(aq) \rightarrow Cl_2(aq) + 2e^-$$

   What is this an example of?
   - **A.** An acid–base reaction
   - **B.** A precipitation reaction
   - **C.** A reduction reaction
   - **D.** An oxidation reaction
5. In which of the following reactions does $Fe^{2+}(aq)$ act as a reducing agent?
   - **A.** $Fe^{2+}(aq) + S^{2-}(aq) \rightarrow FeS(s)$
   - **B.** $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$
   - **C.** $Fe^{2+}(aq) + Mg(s) \rightarrow Fe(s) + Mg^{2+}(aq)$
   - **D.** $2Fe^{2+}(aq) + Br_2(aq) \rightarrow 2Fe^{3+}(aq) + 2Br^-(aq)$
6. A small piece of silver was placed in a solution containing both magnesium nitrate and copper(II) nitrate.

   Which one of the following will occur?
   - **A.** No reaction occurs.
   - **B.** The silver dissolves and only copper is precipitated.
   - **C.** The silver dissolves and only magnesium is precipitated.
   - **D.** A mixture of copper and magnesium forms on the silver.

7. What happens if a piece of shiny aluminium metal is placed in a $CuSO_4$ solution?
   - **A.** Sulfur dioxide gas is liberated.
   - **B.** No reaction occurs.
   - **C.** Copper oxide is formed on the aluminium metal.
   - **D.** Copper metal is deposited on the aluminium metal.
8. What would happen if you tried to store $FeSO_4$ solution in a container made of Cu metal?
   - **A.** The $FeSO_4$ solution would be stored quite safely.
   - **B.** The copper of the container would dissolve and Fe metal would form.
   - **C.** The copper container would dissolve and Fe^{2+} ions would form.
   - **D.** The copper container would dissolve and O_2 gas would form.
9. Which of the following metals could be used to make a container to store an aqueous copper sulfate solution?
   - **A.** Sn
   - **B.** Pb
   - **C.** Ag
   - **D.** Fe
10. Aluminium, although a fairly reactive metal, does not react with oxygen gas in air. Why not?
   - **A.** It is protected by small amounts of its own impurities.
   - **B.** Its oxide forms a hard protective layer on the metal.
   - **C.** Nitrogen and carbon dioxide gas in the air prevent oxidation.
   - **D.** It forms a hard protective layer by a reaction with sulfur in the air.
11. The following diagrams represent four tests in an experiment designed to find the conditions necessary for a steel nail to rust.



   Which diagram represents the most favourable conditions for rust to appear on the nail?
   - **A.** A
   - **B.** B
   - **C.** C
   - **D.** D
12. Why do iron rubbish bins coated with a complete layer of zinc not rust?
   - **A.** The zinc acts as a sacrificial metal, becoming oxidised to prevent the iron rusting.
   - **B.** Iron is a more reactive metal than zinc.
   - **C.** The zinc combines with the iron to form a new compound that does not rust.
   - **D.** Particles of rust are unable to stick to the zinc surface.

14.6 Exercise 2: Short answer questions

1. Define oxidation and reduction reactions in terms of the transfer of oxygen.
2. Define oxidation and reduction reactions in terms of the transfer of electrons.
3. What is an oxidising agent?

4. Use the electron transfer approach to identify the oxidising agent, reducing agent, direction of electron transfer and the direction of oxidation and reduction in the following equations.
    a. $Sn^{2+}(aq) + 2Ce^{4+}(aq) \rightarrow Sn^{4+}(aq) + 2Ce^{3+}(aq)$
    b. $Cd(s) + Pb^{2+}(aq) \rightarrow Cd^{2+}(aq) + Pb(s)$
    c. $2Al(s) + 3Cl_2(g) \rightarrow 2Al^{3+}(aq) + 6Cl^-(aq)$
    d. $Mg(s) + S(s) \rightarrow Mg^{2+}(aq) + S^{2-}(s)$
    e. $2Na(aq) + 2H_2O(l) \rightarrow 2Na^+(aq) + OH^-(aq) + H_2(g)$
5. Balance the following equations using half-equations.
    a. $Br^-\,(aq) + SO_4^{2-}\,(aq) \rightarrow SO_2\,(g) + Br_2\,(l)$
    b. $Al(s) + Cl_2(g) \rightarrow AlCl_3(s)$
    c. $I_2(s) + H_2S(g) \rightarrow I^-(aq) + S(s)$
    d. $Cu(s) + HNO_3(aq) \rightarrow Cu^{2+}(aq) + NO(g)$
    e. $CuO(s) + NH_3(g) \rightarrow N_2(g) + Cu(s)$
    f. $Cr_2O_7^{2-}\,(aq) + CH_3CH_2OH(g) \rightarrow CH_3COOH(aq) + Cr^{3+}(aq)$
6. What is a displacement reaction? Explain using an example.
7. Where are the most reactive metals found in the periodic table?
8. Which of the following reactions is more likely to happen if the reactants are mixed?

$$Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)$$

$$Pb(s) + Zn^{2+}(aq) \rightarrow Pb^{2+}(aq) + Zn(s)$$

9. Develop overall redox equations by first writing the oxidation and reduction equations for the following reactions. Label the oxidising agent and reducing agent in each case.
    a. Zinc solid is placed in a copper sulfate solution.
    b. An iron nail is placed in a solution of tin(II) nitrate.
    c. A small piece of lead is placed in silver nitrate solution.
    d. A small piece of aluminium is reacted with zinc chloride solution.
10. Tin displaces copper metal from $CuSO_4$ solution and cadmium displaces tin metal from $SnCl_2$ solution. The three metals and their cations are as follows: Sn^{2+}/Sn, Cu^{2+}/Cu, Cd^{2+}/Cd.
    a. Ignoring spectator ions, write an ionic equation for the two displacement reactions described.
    b. Use the information given to deduce the relative order of metals from most reactive to least reactive.
    c. State whether you would expect the following reactions to occur.
        i. $Sn^{2+}(aq) + Cu(s) \rightarrow Sn(s) + Cu^{2+}(aq)$
        ii. $Cd^{2+}(aq) + Cu(s) \rightarrow Cd(s) + Cu^{2+}(aq)$
        iii. $Sn(s) + Cd^{2+}(aq) \rightarrow Sn^{2+}(aq) + Cd(s)$
        iv. $Cd(s) + Cu^{2+}(aq) \rightarrow Cd^{2+}(aq) + Cu(s)$
11. Six different strips of metal, *A* to *F*, were placed in solutions of their metal nitrates to see if any reaction took place. The results are shown in the following table .

Metal	A nitrate	B nitrate	C nitrate	D nitrate	E nitrate	F nitrate
A	–	✓	✓	✓	×	✓
B	×	–	×	×	×	×
C	×	✓	–	✓	×	×
D	×	✓	×	–	×	×
E	✓	✓	✓	✓	–	✓
F	×	✓	✓	✓	×	–

✓ = metal displaced × = no reaction

    a. Deduce the order of reactivity of the metals *A* to *F* and list them in order, beginning with the most reactive.
    b. Identify the six metals.

12. Explain, using equations, why the corrosion of iron to form rust is a redox process.
13. Salt is often spread on roads in cold countries to melt the ice. Explain why cars rust quickly in these countries.
14. Explain what would happen if the following coatings of a steel object are damaged.
    a. Plastic
    b. Paint
    c. Tin
    d. Zinc
15. a. What does the term 'noble coating' mean?
    b. Why is it undesirable to coat screws with tin as a protection from corrosion?
16. An iron sculpture was produced to commemorate the anniversary of the founding of a small town. To stop it rusting, the mayor attached it by a wire to a block of zinc that was stored underground out of sight.
    a. Explain how the mayor's action would prevent corrosion.
    b. What would happen if a block of lead were to be attached to the iron rather than the block of zinc?
    c. List two other ways in which corrosion of the statue could be prevented.

14.6 Exercise 3: Exam practice questions

Question 1 (4 marks)

State which of the following are redox reactions:

a. $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$ **1 mark**

b. $HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$ **1 mark**

c. $Cu^{2+}(aq) + Mg(s) \rightarrow Cu(s) + Mg^{2+}(aq)$ **1 mark**

d. $2H_2O(l) + 2F_2(g) \rightarrow 4HF(aq) + O_2(g)$ **1 mark**

Question 2 (15 marks)

Balance the following equations using the half-equation method. Identify the oxidation reaction, the reduction reaction and the balanced equation.

a. $Fe_2O_3(s) + CO(g) \rightarrow Fe(s) + CO_2(s)$ **3 marks**

b. $P_4(s) + HNO_3(aq) + H_2O(l) \rightarrow H_3PO_4(aq) + NO(g)$ **3 marks**

c. $Cr_2O_7^{2-}(aq) + H_2C_2O_4(aq) \rightarrow Cr^{3+}(aq) + CO_2(g)$ **3 marks**

d. $Cu(s) + HNO_3(aq) \rightarrow Cu^{2+}(aq) + NO_2(g)$ **3 marks**

e. $PbS(s) + H_2O_2(l) \rightarrow PbSO_4(s) + H_2O(l)$ **3 marks**

Question 3 (6 marks)

A clean piece of tin metal is placed into a solution of lead nitrate and left for a period of time. It is then observed that the tin has become coated with a layer of dull lead metal.

a. Write the half-equation for the oxidation reaction that has occurred. **1 mark**

b. Write the half-equation for the reduction reaction that has occurred. **1 mark**

c. State the oxidising agent for this reaction. **1 mark**

d. State the reducing agent for this reaction. **1 mark**

e. If the tin is removed, cleaned of lead, and then placed into a solution of magnesium chloride, would a further redox reaction occur? Explain. **2 marks**

Question 4 (5 marks)

Four hypothetical metallic elements, *A, B, C* and *D*, form soluble nitrates with the formulas $A\text{NO}_3$, $B(\text{NO}_3)_2$, $C\text{NO}_3$ and $D(\text{NO}_3)_3$.

Strips of each of the four metals were immersed in solutions of the other metal nitrates and the following observations were recorded.

- Metal *B* underwent reaction in all solutions.
- Metal *A* reacted only with CNO_3.

a. State the formulas of the metal cations in the soluble nitrates. **2 marks**

b. Deduce the position of *D* in the series by listing the three metals *A, B* and *C* in decreasing order, with respect to their relative reactivities. **1 mark**

c. List the four metals in order from most reactive to least reactive. **1 mark**

d. Which metal cation is the weakest oxidising agent? **1 mark**

Question 5 (8 marks)

Experts estimate that over 20% of the iron produced in the world is used to replace that lost by corrosion.

a. List the substances needed for the rusting of iron to occur. **2 marks**

b. Name three other substances that speed up the rusting of iron. **3 marks**

c. Explain why iron objects are less likely to rust in the following environments. **3 marks**

  i. In a desert

  ii. In outer space

  iii. On the ocean floor

14.6 Exercise 4: studyON Topic Test online only

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AREA OF STUDY 1 How do substances interact with water?

OUTCOME 1

Relate the properties of water to its structure and bonding, and explain the importance of the properties and reactions of water in selected contexts.

PRACTICE EXAMINATION

STRUCTURE OF PRACTICE EXAMINATION		
Section	**Number of questions**	**Number of marks**
A	20	20
B	7	30
	Total	**50**

Duration: 50 minutes

Information:

- This examination consists of two parts, you must answer all question sections
- Pens, pencils, highlighters, erasers, rulers and a scientific calculator are permitted.
- You may use the VCE Chemistry Data Book for this task.

on Resources

Weblink VCE Chemistry Data Book

SECTION A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

1. Water makes up the largest proportion of sweat, which the body releases as a cooling mechanism. Which of the following reasons can explain this cooling effect?

A. The density of water decreases as it gets warmer.
B. Water is a good solvent for many of the ionic substances found in the body.
C. The electrical conductivity of water increases as it dissolves more ions.
D. Water has a particularly high latent heat of vaporisation for such a small molecule.

2. New bonds are formed when water freezes to produce ice. What are these bonds between the molecules predominantly?

A. Hydrogen bonds and dispersion forces
B. Covalent bonds and hydrogen bonds
C. Ion-dipole bonds and dispersion forces
D. Ionic bonds and ion-dipole bonds

3. Why is water able to dissolve ionic solids like NaCl?

A. Neutral molecules of NaCl react with water to produce ions.

B. Ion–dipole bonds form between the ions of the lattice and water molecules.

C. Ion–dipole bonds form between the metal cations and the non-metal anions.

D. Water is able to form hydrogen bonds with Na^+ ions and Cl^- ions.

4. Which of the following equations represents the dissociation of solid calcium chloride, $CaCl_2$, when it dissolves in water?

A. $CaCl_2(s) \xrightarrow{water} Ca(s) + Cl_2(g)$

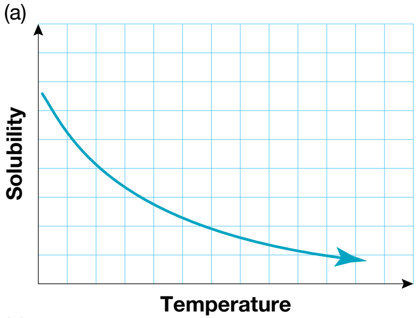
B. $CaCl_2(s) \xrightarrow{water} Ca^{2+}(aq) + Cl_2(g)$

C. $CaCl_2(s) \xrightarrow{water} Ca^{2+}(aq) + 2Cl^-(aq)$

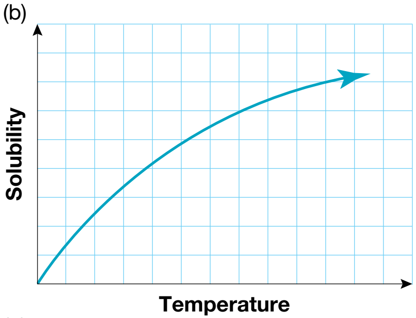
D. $CaCl_2(s) \xrightarrow{water} Ca(aq) + 2Cl^-(aq)$

5. Which of the following curves best represents the solubility of gases, like CO_2, in water?

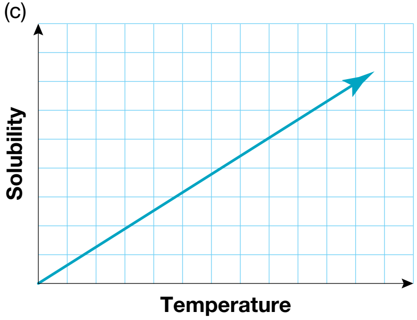
(a)



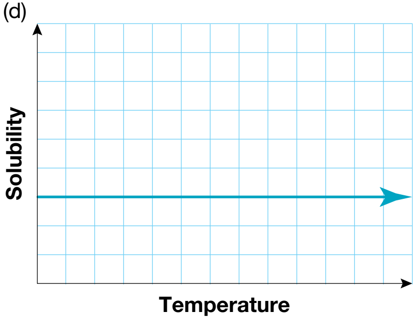
(b)



(c)



(d)



6. Which of the following lists contains only water-soluble substances?

A. HCl, CH_3OH and K_2SO_4

B. CH_4, NaCl and HNO_3

C. NH_4Br, $AgNO_3$ and Ag

D. $C_6H_{12}O_6$, NH_3 and C_3H_8

7. When a solution of sodium hydroxide is added to a solution of copper(II) nitrate, a blue gelatinous precipitate is formed. What happened in this reaction?

A. Copper(II) sulfate forms the precipitate and both sodium and nitrate ions are spectator ions.

B. Sodium nitrate forms the precipitate and hydroxide ions are spectator ions.

C. Copper(II) hydroxide forms the precipitate and sodium ions are one of the spectator ions.

D. Copper metal is deposited and sodium and nitrate ions are spectator ions.

8. When solutions of lead(II) nitrate and potassium iodide are mixed a precipitate forms. What is the balanced ionic equation for this reaction?

A. $Pb^+(aq) + I^-(aq) \rightarrow PbI(s)$

B. $Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$

C. $KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + KNO_3(s)$

D. $K^+(aq) + NO_3^-(aq) \rightarrow KNO_3(s)$

9. In which of the following reactions is the bolded species acting as a Brønsted–Lowry acid?

A. $\mathbf{NH_3}(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

B. $HCOO^-(aq) + \mathbf{H_2O}(l) \rightleftharpoons HCOOH(aq) + OH^-(aq)$

C. $H_2O(l) + \mathbf{F^-}(aq) \rightleftharpoons OH^-(aq) + HF(aq)$

D. $2\mathbf{HCl}(aq) + Mg(s) \rightarrow MgCl_2(aq) + H_2(g)$

10. Which of the following are NOT acid–base conjugate pairs?

A. HCO_3^- and CO_3^{2-}

B. NH_4^+ and NH_3

C. H_3O^+ and OH^-

D. HBr and Br^-

11. The following table lists some chemical species that have been classified according to their acid–base properties. Which classifications are correct?

	Acid	Base	Amphiprotic
I.	H_3PO_4	PO_4^{3-}	HPO_4^{2-}
II.	HCl	$Ca(OH)_2$	H_2O
III.	CH_4	NH_3	NH_4^+
IV.	HCO_3^-	$HClO_4$	CH_3COOH

A. Classifications (I) and (II) only

B. Classifications (I) and (III) only

C. Classifications (II) and (III) only

D. Classifications (II) and (IV) only.

12. Which of the following lists the solutions from lowest to highest pH?

A. 1.0 mol L^{-1} NaOH, 1.0 mol L^{-1} CH_3COOH, 1.0 mol L^{-1} HCl, 2.0 mol L^{-1} HCl

B. 1.0 mol L^{-1} CH_3COOH, 1.0 mol L^{-1} NaOH, 1.0 mol L^{-1} HCl, 2.0 mol L^{-1} HCl

C. 1.0 mol L^{-1} HCl, 2.0 mol L^{-1} HCl, 1.0 mol L^{-1} CH_3COOH, 1.0 mol L^{-1} NaOH

D. 2.0 mol L^{-1} HCl, 1.0 molL^{-1} HCl, 1.0 mol L^{-1} CH_3COOH, 1.0 mol L^{-1} NaOH

13. Two acidic solutions were compared. One was a solution of sulfuric acid, while the other was a solution of a carboxylic acid, ethanoic acid. Both solutions were found to have the same hydronium, H_3O^+, concentration. As a result, what is also true?

A. Both acids have the same molar concentration.

B. Both acids have an equal tendency to react with water.

C. Both solutions have the same pH.

D. Both solutions have most of their acids ionised.

14. When nitric acid reacts with sodium hydrogen carbonate, what is the salt formed in the reaction?

A. Sodium chloride

B. Sodium carbonate

C. Sodium nitrate

D. Sodium sulfate

15. The following four equations represent four different reactions.

I. $HI(aq) + KOH(aq) \rightarrow KI(aq) + H_2O(l)$

II. $Mg(s) + S(l) \rightarrow MgS(s)$

III. $H_2O(l) + HCl(aq) \rightarrow H_3O^+(aq) + Cl^-(aq)$

IV. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

Which of the following statements is true?

A. Reactions (I) and (IV) are both acid–base reactions.

B. Reactions (II) and (IV) are both redox reactions.

C. Reaction (I) is redox and reaction (III) is an acid–base reaction.

D. Reaction (IV) is neither acid–base nor redox.

16. For the rection

$$Pb^{2+}(aq) + Cd(s) \rightarrow Cd^{2+}(aq) + Pb(s)$$

which of the following statements is correct?

A. Cd(s) is oxidised and $Pb^{2+}(aq)$ is the reducing agent

B. Cd(s) is oxidised and $Pb^{2+}(aq)$ is the oxidising agent

C. Cd(s) is reduced and Pb(aq) is the oxidising agent

D. Cd(s) is reduced and $Pb^{2+}(aq)$ is the reducing agent

17. If an element is described as an oxidising agent, then it

A. accepts protons during a reaction.

B. is reduced when it accepts electrons from an reducing agent.

C. donates electrons to a reducing agent.

D. gains electrons from a reducing agent and is oxidised.

18. Which of the following is a balanced *reduction* half-equation?

A. $2H^+(aq) + e^- \rightarrow H_2(g)$

B. $Ni(s) \rightarrow Ni^{2+}(aq) + e^-$

C. $Cl_2(g) \rightarrow 2Cl^-(aq) + 2e^-$

D. $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

19. Several small pieces of zinc, Zn, are added to a solution of cobalt(II) sulfate, which is pink in colour. The equation for this reaction is:

$$Zn(s) + CoSO_4(aq) \rightarrow ZnSO_4(aq) + Co(s).$$

Which of the following observations would be made?

A. A deposit of cobalt metal would form and the zinc pieces would dissolve.

B. The zinc pieces would dissolve and the cobalt(II) sulfate solution would become a stronger pink colour.

C. Cobalt metal would dissolve and the zinc pieces would increase in size.

D. A precipitate of zinc sulfate would form and the cobalt(II) sulfate solution would become a paler pink colour.

20. Which of the following methods will NOT minimise the corrosion of iron?

A. Coating the surface with an impenetrable layer of plastic or enamel

B. Forming an alloy of the metal with non-corrosive metals like chromium

C. Placing a less reactive metal like nickel in electrical contact with the iron

D. Applying a light coating of zinc to the surface

SECTION B

Question 1 (4 marks)

1. a. The density of water varies with temperature. It is at its highest at 4 °C. At temperatures below 0 °C, however, water has a lower density. Explain why this is the case. **1 mark**

b. The density of water at 25 °C is 0.997 g mL^{-1}. Use this value to calculate the mass (in grams) of 10.0 L of water. **1 mark**

c. Given that the specific heat capacity of water is 4.18 J g^{-1} $°C^{-1}$, what amount of heat energy in Joules would be required to heat 10.0 L of water from 25 °C to 100 °C ? **1 mark**

d. What will happen to the density of water as its temperature rises from 25 °C? Give reasons for your answer. **1 mark**

Question 2 (5 marks)

2. Write equations with water to demonstrate the following:

a. Nitric acid, HNO_3, is a strong monoprotic acid. **1 mark**

b. H_2S is a weak diprotic acid. **2 marks**

c. The hydrogen sulfite ion, HSO_3^-, is an amphiprotic substance. **2 marks**

Question 3 (2 marks)

3. Explain the following chemical properties.
   - **a.** Methanoic acid, HCOOH, is a molecular substance, yet when it is dissolved in water, the resulting solution is able to conduct electricity. **1 mark**
   - **b.** Citric acid is found in many foods and is able to be consumed, yet spilling hydrochloric acid on your skin would result in a chemical burn. **1 mark**

Question 4 (3 marks)

4. **a.** A solution of hydrochloric acid (HCl) was prepared by diluting 0.50 mol L^{-1} hydrochloric acid stock solution.
   - **i.** Calculate the pH of 0.50 mol L^{-1} hydrochloric acid. **1 mark**
   - **ii.** What is the final concentration of hydrochloric acid if 100.0 mL of water is added to 20.0 mL of 0.50 mol L^{-1} hydrochloric acid? **1 mark**

   b. Determine the pH of a 0.0010 mol L^{-1} solution of the alkali, NaOH. **1 mark**

Question 5 (6 marks)

5. Nearly all solutions of transition metal compounds form precipitates when mixed with aqueous potassium carbonate.

   a. i. Write the full and the ionic equations for the reaction between solutions of iron(II) nitrate and potassium carbonate. Include symbols of state. **2 marks**

   ii. List the formulas of the spectator ions in this reaction. **1 mark**

   b. i. Write the full equation for the reaction between hydrobromic acid (HBr) and zinc carbonate powder. **1 mark**

   ii. List two things that would be observed during this reaction. **1 mark**

   c. Write the ionic equation for the reaction between sulfuric acid and nickel(II) hydroxide solid. **1 mark**

Question 6 (7 marks)

6. **a.** Concentrated nitric acid, HNO_3, is able to undergo redox reactions with many metals, including less reactive metals like copper. This reaction produces copper(II) ions and nitrogen dioxide gas among the products.
   - **i.** Derive the two half-equations for this reaction. **2 marks**
   - **ii.** Label the half-equation that is the reduction reaction. **1 mark**
   - **iii.** Use the two half-equations to write the overall ionic equation for this reaction. **1 mark**

   b. Nitrogen dioxide molecules often combine (dimerise) to form dinitrogen tetroxide, N_2O_4. What is the oxidation number of nitrogen in the following?
   - **i.** HNO_3 **1 mark**
   - **ii.** N_2O_4 **1 mark**

   c. In some reactions with metals, nitric acid produces nitrogen monoxide instead of nitrogen dioxide. Balance the following equation using the half-equation method.
   __ $HNO_3(aq)$ + __$V(s)$ + __$H^+(aq) \rightarrow$ __$NO(g)$ + __$V^{2+}(aq)$ + __H_2O **1 mark**

Question 7 (3 marks)

7. The reactivity series of metals is listed from highest to lowest:
   K, Na, Ca, Mg, Al, Zn, Fe, Ni, Sn, Pb, Cu, Ag, Au.
   The following experiments were set up to investigate this series using displacement reactions with metals.

   a. In the first experiment, three iron nails were polished and each was placed into a 1 mol L^{-1} solution of one of the following compounds: nickel(II) sulfate, zinc sulfate and tin(II) sulfate.
   - **i.** Predict the products of any potential reaction(s). **1 mark**
   - **ii.** Write the ionic equation for one predicted reaction. **1 mark**

   b. In a second experiment, three nails made of a different metal from iron were tested in the same solutions. This time none of three nails were observed to react. Give a possible metal used to make these nails. **1 mark**

ASSESSMENT TASK — SCIENTIFIC INQUIRY ESSAY

In this task, you will write a scientific inquiry essay on an issue related to the properties and reactions of water. This essay will relate the properties of water to a real-life issue.

- Pens, pencils, highlighters, erasers, rulers and a scientific calculator are permitted.
- You may use the VCE Chemistry Data Book for this task.

Total time: 55 minutes (5 minutes reading, 10 minutes planning, 40 minutes writing)

Total marks: 50 marks

WATER AND THE ENVIRONMENT

Introduction

Science writing is not always in the form of academic research papers. The general public receives the majority of their information about science through the popular media. Science journalists play an important role in presenting up-to-date scientific information and engaging the public with important scientific issues and developments.

Task

This task is a problem-solving assessment in the form of a science essay. You will discuss the real-life applications of chemical concepts, their related issues, and the effect on our everyday lives. The task is to write an essay based on **one** of the following topics.

- The relationship between acid rain and the environment
- The global issue of ocean acidity
- Metal corrosion effects on our environment

Please use the following assessment tasks as a guide to sequencing your essay. The assessment will be based on the following.

1. Presentation of the issue.
2. Causes of the issue.
3. Use of relevant graphs, reactions, equations and diagrams; for example, if referring to ocean acidity, chemical equation examples should be used to demonstrate this chemical process.
4. Effects of the issue, including discussion of the current state of the issue.
5. Solutions to the issue, including opinions or suggestions as to how the problem could be resolved using chemistry.

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AREA OF STUDY 2
HOW ARE SUBSTANCES IN WATER MEASURED AND ANALYSED?

15 Water sample analysis

15.1 Overview

Numerous **videos** and **interactivities** are available just where you need them, at the point of learning, in your digital formats, learnON and eBookPLUS at www.jacplus.com.au.

15.1.1 Introduction

Like many countries around the world, Australia's annual rainfall fluctuates highly, having an impact on surface water flows and sources for drinking water. Huge improvements have been made in the provision of drinking water around the world. For example, since 1990, 2.6 billion people have gained access to an improved drinking water source. An improved source is one that is designed to protect against contamination at the point of distribution or collection. The main concern in drinking water contamination is the presence of bacteria (such as *E. coli*) from faecal contamination, which has been detected even in protected water sources. Low-cost techniques for testing local water supplies are intrinsic to protecting the health of billions of people around the world. In Australia, the challenges of fluctuating rainfall are being met with the creation of additional water sources, such as desalinated water, to ensure that sufficient drinking water sources are available.

In this topic, you will explore the sampling protocols employed when testing for drinking water and some of the contaminants found in water supplies.

FIGURE 15.1 Since 1990, 2.6 billion people have gained access to an improved drinking water source.

KEY KNOWLEDGE

In this topic, you will investigate:

- the existence of water in all three states at Earth's surface, including the distribution and proportion of available drinking water
- sampling protocols, including equipment and sterile techniques for the analysis of water quality at various depths and locations
- the definition of a chemical contaminant and an example relevant to a selected water supply.

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Digital documents Key science skills (doc-30903)

Key terms glossary — Topic 15 (doc-30904)

studyon

To access key concept summaries and practice exam questions download and print the **studyON: Revision and practice exam question booklet** (doc-30905).

15.2 The states of water and sampling protocols

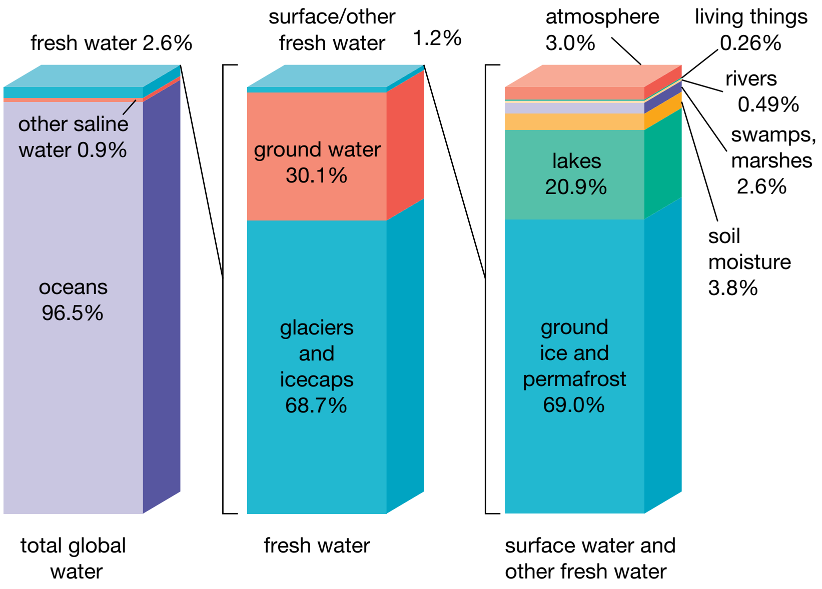
KEY CONCEPTS

- Existence of water in all three states at Earth's surface, including the distribution and proportion of available drinking water
- Sampling protocols, including equipment and sterile techniques for the analysis of water quality at various depths and locations

15.2.1 Using water

Nearly all the water that is used around the world every day, including that used for drinking, is not pure. This is because, as we have seen in previous topics, water is an excellent solvent. In the environment, water comes into contact with many substances that dissolve in it. As water falls through the atmosphere in the form of rain, gases from the atmosphere, both natural and pollutants, dissolve in it to some extent. When it travels over the Earth's surface, or soaks into the ground, naturally occurring salts also dissolve. As demonstrated in figure 15.2 and table 15.1, only ~2.6 per cent of all water on Earth is fresh water, and of this 2.6 per cent only 1.2 per cent is surface or other fresh water. This means only ~0.03 per cent of all water on Earth must sustain all land-based life, including more than 7 billion people.

FIGURE 15.2 Global distribution of water



As all water eventually flows into the oceans as part of the water cycle, this is precisely why the ocean is so salty and contains a wide range of dissolved substances. The major ions in sea water are shown in table 15.2.

The United Nations has estimated that water usage has grown at more than twice the rate of population increase in the last century and by 2025 an estimated 1.8 billion people will live in water scarce regions. Increased water usage is exacerbated by population growth and climate change, and by 2025 two-thirds of the world's population will live in water-stressed regions. Therefore, we need to ask, from a realistic viewpoint, whether the water we use is suitable for the purpose we have in mind. To answer such a question, we need to know not only *what* is dissolved in the water, but also *how much*. And we also need to know *how to measure* what the water contains.

TABLE 15.1 Approximate global distribution of water by percentage and state

	Location of water	State of water	Percentage of total water on Earth
Salt water 97.4%	Oceans	Liquid	96.5%
	Other saline water	Liquid	0.9%
Fresh water 2.6%	Glaciers and icecaps	Solid	1.72%
	Ground water	Liquid	0.753%
	Ground ice and permafrost	Solid	0.0207%
	Lakes	Liquid	0.00627%
	Soil moisture	Liquid	0.00114%
	Atmospheric water vapour	Gas	0.0009%
	Swamps and marshes	Liquid	0.00078%
	Rivers	Liquid	0.000147%
	Living things	Liquid and gas	0.000078%

Each state in Australia uses one or more sources of water depending on geography, available sources and population. Australia does not have a large number of natural lakes to assist in water storage, and as the driest inhabited continent, with highly variable rainfall and runoff, water management is a significant undertaking. In general, the sources of water in Australia are from:

- water catchments — protected and open
- lakes
- rivers and creeks
- groundwater (borewater)
- desalinated water
- recycled water.

About 15 670 billion litres of water was used for consumptive purposes across Australia in 2016–17. The biggest use of this water was agriculture (70%), followed by urban use (20%) and other industries (10%). Water in urban areas of Australia is sourced from a mains water supply, called **reticulated water**. Reticulated water is controlled by local water authorities and is carried via a pipe from a reservoir. These reservoirs may source their water from a few sources. For example, Melbourne's source of water is mainly from surface water assisted by desalinated water; Perth uses both groundwater and desalinated water; Adelaide relies on water from rivers and desalinated water.

TABLE 15.2 Major ions in sea water

Ion	Typical concentration (mg L^{-1})
Chloride, Cl^-	18 980
Sodium, Na^+	10 556
sulfate, SO_4^{2-}	2 649
Magnesium, Mg^{2+}	1 262
Calcium, Ca^{2+}	400
Potassium, K^+	380
Bicarbonate, HCO_3^-	140
Strontium, Sr^{2+}	13
Bromide, Br^-	65
Borate, BO_3^{3-}	26

In Melbourne, Melbourne Water is the authority that treats and provides high-quality water across the Port Phillip and Westernport region. Figure 15.3 shows the sources of drinking water for Melbourne, including ten catchment reservoirs that are linked through a network of pipes. These reservoirs depend on rainfall on a catchment area or diversion from a river. To diversify the water supply system for a growing population, a rainfall-independent desalination plant was built. The plant assists with maintaining water storage volumes and recovery after drought periods, and is part of planning for increased water resources. Figure 15.4 demonstrates the long-term planning for Melbourne's water supply, which is based on modelling of climate change impacts and potential demand variations.

FIGURE 15.3 Melbourne water supply system (Melbourne Water)

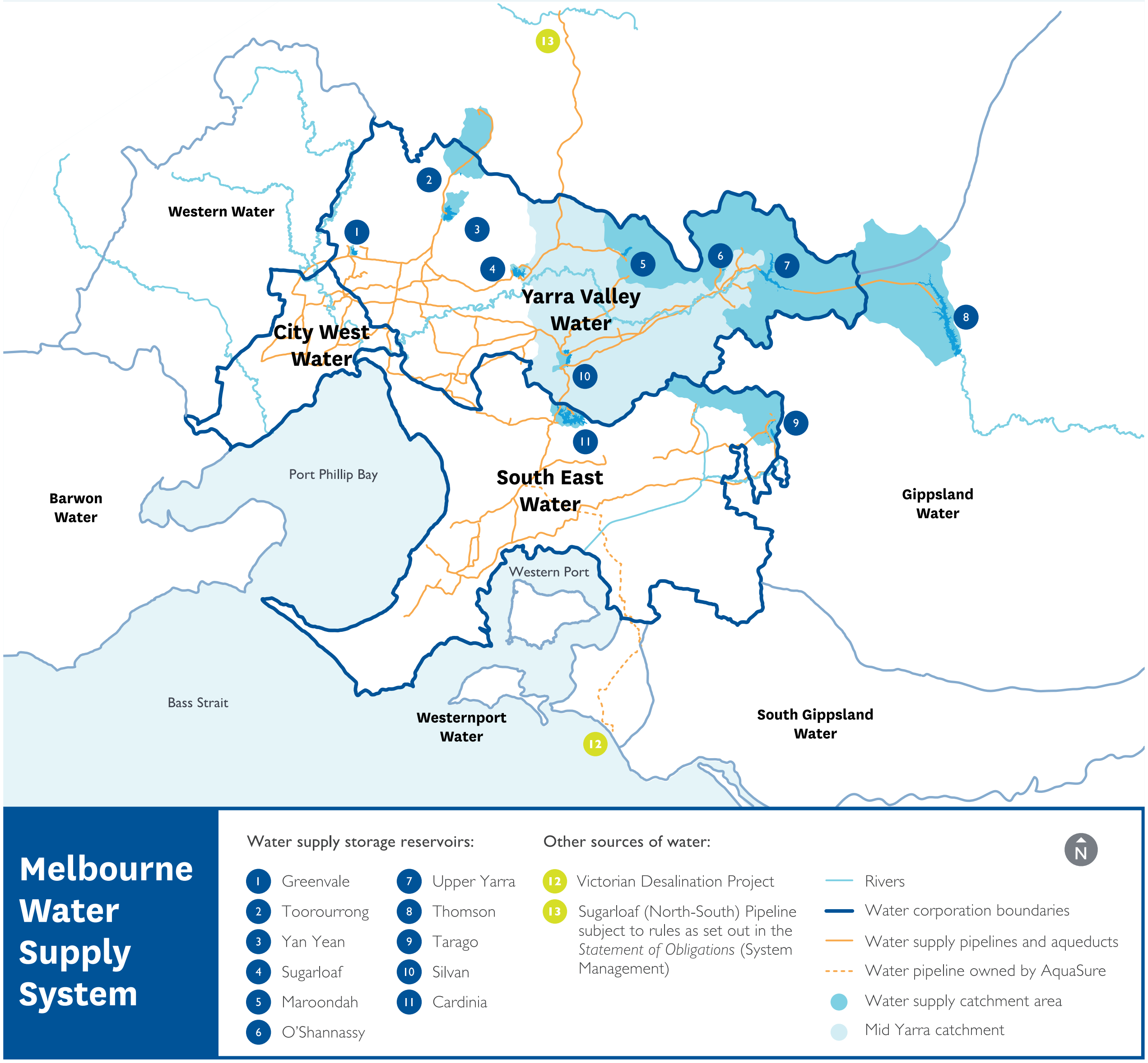


FIGURE 15.4 Long-term water supply and demand for Melbourne.

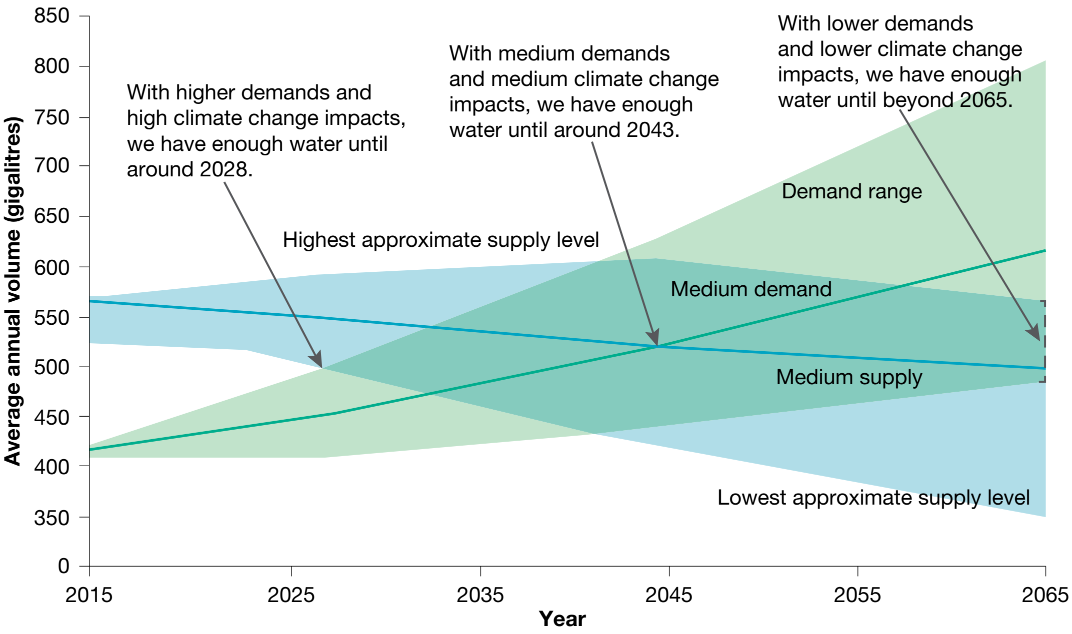


FIGURE 15.5 Thomson reservoir, the largest reservoir supplying water to Melbourne.

on Resources

Weblinks Melbourne's water supply

Water storage dashboard

15.2.2 Sampling protocols

Water sampling in the field can be a complex task because it depends on the water use and the type of analysis required. In Australia, to ensure consistency of sampling methods and result reporting, the Australian Government provides sampling standards and guidelines for sampling, and for transporting and preserving the sample (AS/NZS 5667.1:1998). These strict protocols exist for the collection of water samples in order to produce accurate analytical results. These apply before, during and after sampling takes place. Australia also has standards for water quality dependent upon its final intended use — drinking, recycling (for irrigation or animal use), swimming (contact), irrigation (agriculture), or treated wastewater (for discharge in the sea) — with a standard applied for each water purpose.

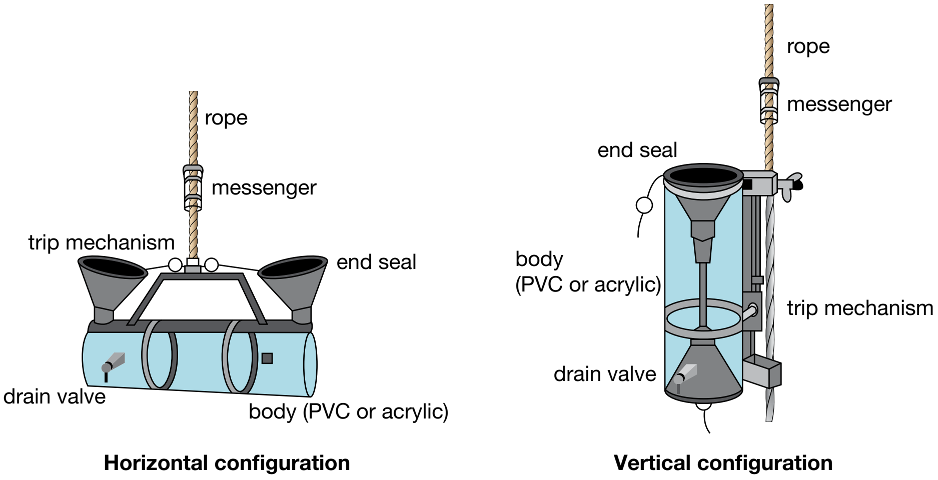
To obtain an accurate analysis of a water sample, it is important that a number of precautions are taken. These are discussed below.

Representative samples

The sample collected must be a **representative sample**. Considerations for a representative sample include sampling from a variety of locations and/or depths, or ensuring the water is well mixed. The sample should be representative of 'average conditions' for the particular location. It would not be appropriate, for example, to take samples during a flood, unless that was the specific reason for the analysis.

Surface water samples may be taken from 100 mm below the surface if the water is considered to be well mixed. Otherwise, several samples need to be taken at different depths that represent a specific 'layer' of water. In that case, a pump and a Van Dorn sampling instrument may be required (figure 15.6).

FIGURE 15.6 Van Dorn sampling instrument in vertical and horizontal configurations

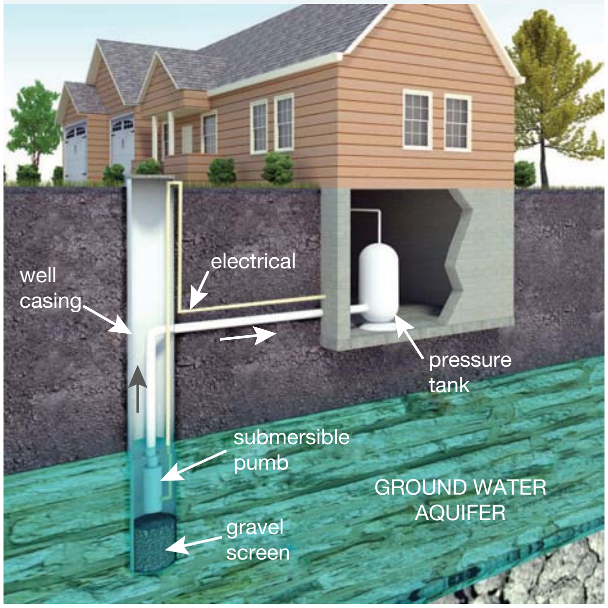


River water samples need to be taken at various locations that may be spread over a few kilometres. 'Grab' samples are usually taken to capture a representative sample for the location (figure 15.7).

Groundwater is typically sampled via the existing bores, using specific instruments by trained personnel (figure 15.8). The bore needs to be purged several times before samples can be considered representative. The volumes of the bore to be purged need to be calculated based on the diameter of the bore and depth.

FIGURE 15.7 Taking grab samples from river water

FIGURE 15.8 Simplified schematic of groundwater use at house scale.



Sampling equipment and procedures

The choices of sample equipment and sampling techniques help prevent contamination. The material of the sampling container must be considered. For example, if the analytes to be tested are organics, a glass bottle is required, because they do not interact with the glass. Conversely, for inorganic samples, a glass bottle is not suitable because the active sites from the glass can interact with the inorganic analytes. In this case, a polypropylene bottle may be used. Container lids should also be considered. Pre- and post-contamination can be minimised by specifying if containers must be completely filled, and avoiding contaminants during transportation including exhaust fumes and dust. This is important because the expected level of contaminants in the sample may be in parts per million or even parts per billion.

Cross-contamination may be prevented by thorough cleaning of sampling equipment and bottles between sampling days to avoid contamination from a prior sampling date.

Duplicate samples or even triplicate samples are required to increase the precision of the analysis. When results are analysed, the duplicates and triplicates need to agree within a range of 10 to 20% of each other.

Sample preservation

Sample preservation may involve cooling, freezing or adding chemicals. If chemicals are added, the preparation of ‘blanks’ is important; these contain the added chemical and distilled water only, rather than the sampled water. Once collected it is important that samples are analysed within certain prescribed times.

Finally, an appropriate analytical technique should be chosen. Prescriptions exist for analysis of particular contaminants, and are discussed in topics 17, 18 and 19. Additionally, such tests are often required to be carried out only by suitably qualified laboratories.

The level of sophistication required can be illustrated by the collection of samples for Fe^{2+} analysis. Many metals can be collected in glass, polyethene or polypropene containers, provided they are first washed with nitric acid to provide a pH between 1 and 2. However, for Fe^{2+}, hydrochloric acid must be used instead, because strong nitric acid is an oxidant and may convert some of the Fe^{2+} into Fe^{3+}. For the same reason, it is specified that the containers must be completely filled to exclude air, because oxygen gas is also an oxidant.

A number of tests are also done ‘in the field’. These include measurements of temperature, pH, dissolved oxygen, some anions such as fluoride and sulfide, and turbidity (cloudiness). Portable instruments are used for these tests, and these must be regularly calibrated against laboratory standards to maintain their accuracy.

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Weblink Sampling and analysis

15.2 EXERCISE

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1. Australia has the highest per capita water storage capacity in the world. Why?
2. What ways can water be sourced if reticulated water is not available?
3. Explain why it is necessary to have strict protocols for the collection of water samples for analysis.
4. Why do we need to take duplicate samples when sampling water?
5. Use the **Sampling and analysis** weblink in the Resources tab to access a list of procedures to be followed for the collection of water samples. Use appendix A of the document to list some of the procedures that need to be followed to collect water samples for testing the following.
   (a) Sulfate ions
   (b) Magnesium ions
   (c) Cyanide
   (d) Fe^{2+}
   (e) Total mercury
   (f) Chlorophyll

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15.3 Chemical contaminants

KEY CONCEPT

- The definition of a chemical contaminant and an example relevant to a selected water supply

In many ways, it is unfortunate that water is such a good solvent, because it means obtaining pure water is difficult and water is also easily polluted. For centuries, the oceans of the world have been a convenient and highly efficient water treatment plant, because almost anything we put in the sea is eventually dissolved or broken down into constituent atoms, molecules or ions. Natural processes include the addition of salts to rivers and oceans as rocks and soils are eroded.

The human impact on water sources is astounding. Globally, effluent has been poured into rivers, which eventually flow out to the sea. Much of this sewerage waste is not treated in any way. Oil is washed out of the holds of tankers (see topic 18 for a discussion of oil spills). Factories discharge waste directly into the sea.

The more sophisticated our technology becomes, the more complex are the pollutants that we discharge. This means that it takes longer for the sea to degrade the waste. Some waste is **biodegradable**. This means that it is broken down by natural processes. However, an increasingly larger proportion of waste is **non-biodegradable**.

15.3.1 Sources of contaminants in water

A **contaminant** can be defined as an unwanted substance that makes water unsuitable for an intended use. Both inorganic and organic contaminants enter water in many ways. Sources of contaminants include:

- *direct discharge* from factories and other sources in both surface water and sea water. This can also happen when poorly maintained sewerage systems overflow or when ships discharge ballast water inappropriately. Agricultural industries commonly contaminate surface and ground waters with nitrates from fertiliser run-off and farm animal wastes.
- *stormwater run-off*. As water flows over the surface, it may come into contact with surface contaminants and dissolve chemicals from them in the process. It may also flow through contaminated soil, dissolving soluble components and carrying soil particles along as a suspension.
- *contaminated air contact*. Components of contaminated air may either dissolve directly in water or react with it to produce a range of undesirable chemicals.
- *groundwater contamination*. This occurs when surface water comes into contact with contaminated soil and then soaks into the ground. It may also occur when leaching through old buried deposits such as landfill sites. Sometimes, the material being leached out is organic in nature and largely insoluble in water. In cases such as this, it may form 'pools' in underground reservoirs and become a source of long-term pollution as it dissolves very slowly over many years.

15.3.2 Contaminants: heavy metals

While **heavy metals** are natural components in the Earth's crust, they can contaminate the water through various means — both through natural contamination and as the result of human activity. They are contaminants of concern for human health because they do not degrade, they are cumulative in the human body and some of them are known to induce cancer and other health problems.

Humans are exposed to heavy metals mainly through drinking water and food consumption; therefore, close monitoring of the drinking water quality is paramount.

Limits for acceptable concentrations of heavy metals in water are included in almost all standards, but specifically in the drinking water standards in Australia. A selection of the guidelines for a few metals of concern is shown in table 15.3.

TABLE 15.3 Selection of heavy metals guideline values in drinking water

Heavy metal	Health guideline (mg/L)	Aesthetic guideline, (mg/L)	Source in drinking water
Arsenic	0.01		From natural sources and mining, industrial and agricultural wastes, including: manufacture of electrical components; wood preservatives and pesticides; gold mining; processing of ceramic materials
Cadmium	0.002		From industrial or agricultural contamination; contaminant in metals used to galvanise pipes and their subsequent erosion; metallurgical industries; improper waste disposal (e.g. nickel–cadmium batteries)
Chromium	0.05		From industrial or agricultural contamination of raw water or corrosion of materials in distribution system/plumbing; metallurgical processes such as alloying and electroplating; catalysts and oxidants; paint pigments

(Continued)

TABLE 15.3 Selection of heavy metals guideline values in drinking water *(Continued)*

Heavy metal	Health guideline (mg/L)	Aesthetic guideline, (mg/L)	Source in drinking water
Copper	2	1	From corrosion of copper pipes
Iron		0.3	Occurs naturally in water in small quantities; high concentrations stain laundry and fittings; iron bacteria cause blockages, taste/odour, corrosion
Lead	0.01		From dissolution from natural sources and domestic/industrial sources including household plumbing containing lead (e.g. pipes, solder); paint pigments; cable sheathing; ammunition; solder; batteries
Mercury	0.001		From mining and industrial wastes; fungicides; electrical equipment; batteries
Nickel	0.02		From prolonged contact of water with nickel-plated fittings
Zinc		3	From corrosion of galvanised pipes/fittings and brasses

Source: NHMRC, 2018, Australian Drinking Water Guidelines 6, Version 3.5.

15.3.3 Contaminants: persistent organic pollutants

Substances that persist in the environment because they cannot be broken down are non-biodegradable. On the other hand, those that can be broken down by natural means are biodegradable.

Persistent organic pollutants (POPs) are contaminants that are difficult to eradicate from the environment. Many have been used in large amounts in the past and are now considered a global concern. The Stockholm Convention of 2004 formally recognised an original twelve POPs, plus a further nine chemicals (or classes of chemicals) as posing a particularly unacceptable risk to the environment and to human health. This treaty aims to reduce and eliminate, where possible, the use of such chemicals.

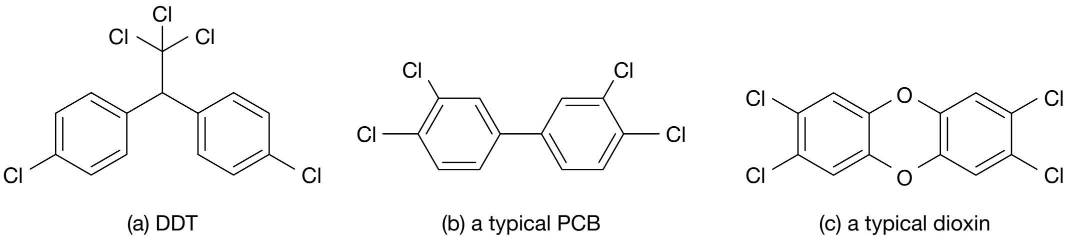
One of the most well-known chemicals on this list is DDT. DDT was an insecticide used in large amounts throughout the 1940s, 1950s and 1960s to control a number of insects, including the anopheles mosquito, the carrier of malaria. In this role, DDT was hugely successful and the amount used quickly escalated. The publication of Rachael Carlson's famous book *Silent Spring* in 1962, however, drew attention to a number of emerging concerns, including an apparent decline in the reproductive rates of a number of bird species. Birds, such as the American bald eagle, were producing eggs with very thin shells that broke as the parents sat on them. This in turn seemed to be associated with increased levels of DDT in the body fat of these birds. The use of DDT is now banned in many countries.

FIGURE 15.9 The American bald eagle was once on the endangered species list as its reproduction was compromised by the birds' ingestion of DDT.

Polychlorinated biphenyls (PCBs) and dioxins are two further groups of these POPs. PCBs have very specific uses in the electrical industry (in circuit boards) and are now being replaced with less dangerous alternatives. Dioxins, on the other hand, have no use at all but are often formed as unwanted by-products in other processes. Dioxins can even be produced by burning waste in domestic incinerators.

The structures of DDT, a typical PCB and a typical dioxin are shown in figure 15.10.

FIGURE 15.10 Structures of some persistent organic pollutants



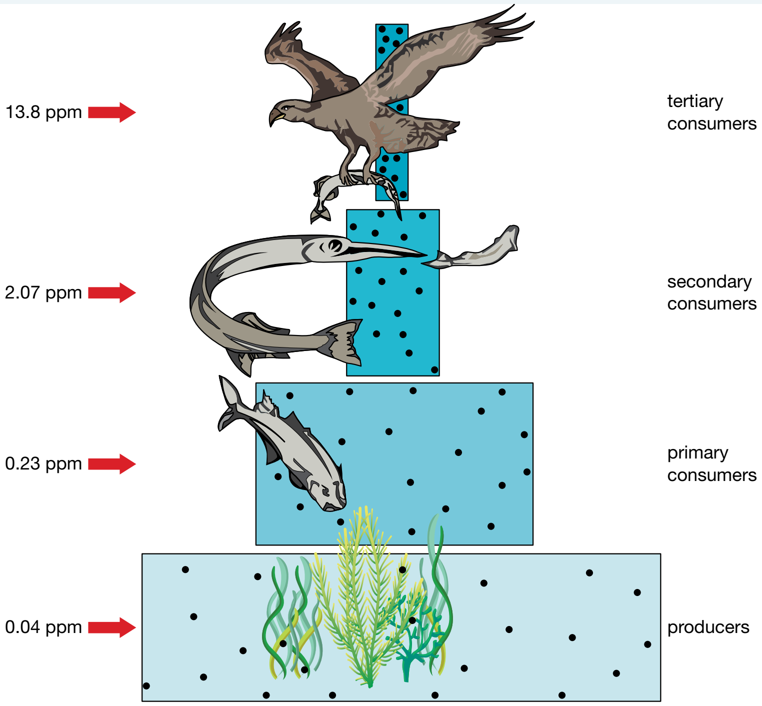
Biomagnification

Biomagnification is the process by which a substance becomes more concentrated in the tissues of organisms higher up the food chain. This results in the concentration of these substances in their bodies becoming higher than in the surrounding environment and the substances becoming impossible to eliminate. Biomagnification can occur when the substance involved is:

- persistent in the environment, even at low levels. This occurs when a substance withstands environmental processes that would normally break it down.
- fat soluble, rather than water soluble. Most substances that become biomagnified are non-polar in nature and, therefore, are soluble in the typically non-polar fatty tissues of animals. This non-polarity also means that they are water insoluble, a property that contributes to their persistence.
- difficult to eliminate. An animal may not be able to metabolise or excrete something once it has entered its body. If the substance is in the fatty tissues, this elimination is even harder.

Many cases of biomagnification have been well-documented, including mercury (especially in the form of methyl mercury), pesticides and DDT. Methyl mercury is discussed in further detail in topic 17.

FIGURE 15.11 Biomagnification of mercury resulting in levels at the top of the food chain considerably higher than the background environmental levels.



on Resources

Weblink Persistent organic pollutants (POPs)

15.3.4 Contaminants: microorganisms

FIGURE 15.12 *Escherichia coli* (*E. coli*) is a micro-organism commonly found in the bowel.

Fresh water contains many microscopic organisms such as phytoplankton, bacteria, algae and protozoa. These organisms may be harmful or beneficial to other forms of life. For example, the biological decay of plant and animal matter often produces nitrogen-rich urea and ammonia, both of which are very soluble. Bacteria may then react with these compounds, ultimately converting the nitrogen in them into free atmospheric nitrogen. This is a critical step in the nitrogen cycle, so it is important to life. On the other hand, disease-causing (pathogenic) bacteria may be present, especially if run-off has been in contact with animal waste, or poorly treated sewage has entered the water. By testing water for the presence of various bacteria, scientists can determine whether water is contaminated. The presence of *Escherichia coli* (*E. coli*) indicates that the water contains pathogenic bacteria.

Eutrophication

When conditions are right, phytoplankton (and especially blue-green algae or cyanobacteria) may reproduce explosively to seriously affect water quality. These conditions usually include still water, sunlight and an excess of nutrients required for growth. Two important such nutrients are nitrates and phosphates, which may be present in excess as a result of human activity. Domestic waste water, agricultural practices and sewage are the main sources of these excess nutrients. This out-of-control growth is often noticed as an ‘algal bloom’ and, besides being unsightly, can have serious consequences. When these organisms die, their subsequent decomposition seriously depletes the level of dissolved oxygen in the water. As a result of this, animals, especially fish, and even plants may die due to a lack of oxygen required for respiration. Additionally, all this decay may produce biotoxins, which can be a serious health hazard to any organism that consumes this water. This condition is called **eutrophication**.

15.3.5 Drinking water standards

The production and treatment of water for drinking is subject to standards to ensure public health, and the maximum permitted levels of some contaminants are summarised in table 15.4. Federal and state standards cover allowed levels of biological, chemical and radiological contaminants, as well as other issues such as pH and turbidity. Various acts and guidelines describe what the standards are and how they are to be monitored. These include:

- *Safe Drinking Water Act 2003* and *Safe Drinking Water Regulations 2005*, which were set and administered by the Victorian Department of Health
- Australian Drinking Water Guidelines 2011, which were developed by the National Health and Medical Research Council.

To ensure water quality is met, water authorities in collaboration with EPA continuously plan, sample and monitor the sources of our drinking water to identify any contaminants of concern. EPA monitors industrial discharges to rivers and ensures they are properly treated prior to discharge. Water authorities protect the reservoir areas by restricting access, and then provide sampling and monitoring of water quality. The water is then treated prior to being distributed via the reticulated system. This way, only the highest quality of water is provided to users.

Water quality concerns for Melbourne include low water flow during periods of drought. Low precipitation levels lead to a low volume of water in reservoirs, which in turn leads to having to access water potentially contaminated with lake sediments. To mitigate this risk of a low volume of water available to use, a desalination plant was constructed to provide a secondary source of water.

Every other state in Australia faces different challenges, depending on weather fluctuations and the specific source of water. However, the typical treatment steps in the water treatment are coagulation, flocculation, sedimentation, filtration and disinfection. Fluoridation is also an additional step employed in some areas.

TABLE 15.4 Maximum permitted levels of some contaminants in drinking water

Contaminant	Maximum drinking water level (mg/L)
Mercury	0.001
Lindane (an insecticide)	0.01
Lead	0.01
2,4-D (a herbicide)	0.03
Sodium	180 (aesthetic level only)

Source: NHMRC, 2018, Australian Drinking Water Guidelines 6, Version 3.5.

15.3 EXERCISE

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1. List the ways that chemicals can enter water bodies.
2. (a) Define the term ‘chemical contaminant’.
   (b) Give an example of a situation in which water might be unsuitable for one particular use but suitable for another use.
3. Define POPs and give two examples.
4. What characteristics would a substance have to have to lead to biomagnification?
5. Through what mechanism can POPs be transported away from the source of contamination?
6. What are the sources in drinking water for the following.
   (a) Lead
   (b) Nickel
   (c) Iron

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15.4 Review

15.4.1 Summary

The states of water and sampling protocols

- Water is such a good solvent that it is rarely pure. Fresh water constitutes only 2.6 per cent of the Earth’s total water but may still contain a number of dissolved substances.

- Strict protocols must be observed when collecting water samples for analysis so that the results obtained are accurate.

Chemical contaminants

- A contaminant is a substance that makes water unsuitable for an intended use.
- Contaminants enter waterways through a variety of methods, including direct discharge, stormwater run-off, contaminated air contact and contact with contaminated groundwater.
- A group of 21 substances or classes of substances have been identified as persistent organic pollutants (POPs). The use of these substances is now banned or severely restricted in many countries.
- DDT is an example of a chemical that was applied extensively as an insecticide but is now on the list of POPs.
- Compounds can be classified as biodegradable or non-biodegradable. Non-biodegradable substances may persist in the environment for an unacceptably long time.
- Biomagnification is the process by which levels of POPs build up in the fatty tissues of higher order organisms.
- Many pollutants dissolve in water because it is such a good solvent. Some waste is biodegradable and can be broken down by natural processes, but much is non-biodegradable. This poses environmental problems where species may be endangered due to water pollution.
- Eutrophication is caused by excessive levels of nutrients in water, which leads to explosive growth of microorganisms such as cyanobacteria. When these die, their subsequent decomposition leads to depleted levels of dissolved oxygen and the death of plants and fish.

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15.4.2 Key terms

biodegradable able to break down via natural processes over the medium term

biomagnification process by which a compound (such as a pollutant or pesticide) increases its concentration in the tissues of organisms as it travels up the food chain

contaminant the presence of unwanted chemicals, or chemicals present in a higher concentration than wanted that makes water unsuitable for its intended use

duplicate sample repeat of the original sample, taken under the same conditions, at the same time and the with same sampling equipment, in order to estimate the sample variability

eutrophication excessive levels of nutrients in water leading to explosive growth of microorganisms; subsequent death of these microorganisms can lead to depleted levels of dissolved oxygen and the death of plants and fish

heavy metal metal with high density or of high relative atomic weight and that is a risk to health or environment

non-biodegradable not able to break down in a natural environment

representative sample small sample with similar characteristics that accurately reflects the larger entity

reticulated water water provided through a network of pipes

on Resources

Digital document Key terms glossary — Topic 15 (doc-30904)

15.4 Exercises

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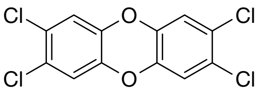
15.4 Exercise 1: Multiple choice questions

1. What is the use and phasing out of persistent organic pollutants governed by?
   - A. The Geneva Convention
   - B. The Stockholm Convention
   - C. The Helsinki Convention
   - D. The Dublin Convention
2. Which of the following statements about dioxins is true?
   - A. Dioxins have been used in the past as insecticides.
   - B. Dioxins are inorganic chemicals containing heavy metals.
   - C. Dioxins are restricted to areas where there are large chemical industries.
   - D. Dioxins are unwanted by-products that have no useful applications.
3. Biomagnification is the process whereby the concentration of a pollutant is increased
   - A. as it is exposed to sunlight and microbial action.
   - B. in organisms in each step up a food chain.
   - C. as remains undergo decomposition.
   - D. due to a population explosion of a particular species.
4. Despite their low solubility in water, many organic contaminants can persist in water bodies for a long time. Why is this?
   - A. They adsorb strongly to the surface of soil particles in river beds.
   - B. Other liquid impurities in the water dissolve them.
   - C. They are broken down by natural microbial action in the soil.
   - D. They accumulate in the bodies of aquatic organisms, which then die.
5. How is reticulated water delivered?
   - A. Via an underground bore
   - B. From a pump in a river
   - C. Via a network of pipes
   - D. From a desalination plant

15.4 Exercise 2: Short answer questions

1. What is the percentage of fresh water on the Earth?
2. Define the following.
   - a. A biodegradable substance
   - b. A non-biodegradable substance
3. Dioxins are a class of persistent organic pollutants. The figure shows the structure of a typical dioxin.



   - a. Write the molecular formula for this dioxin.
   - b. How are dioxins formed?
   - c. Explain why the expected concentration of dioxins in water would be very low.
   - d. Give two reasons dioxins accumulate in human fatty tissue

4. Undertake research to find an example of major contamination of ground water used for drinking water.
    a. What was the contaminant?
    b. What were the impacts?
    c. How was the contamination treated?
5. What are typical water treatments steps in a drinking water system?

15.4 Exercise 3: Exam practice question

Question 1 (11 marks)

A house in a remote area in Victoria needs to be connected to a water source as soon as possible. Reticulated water is not available. However, an option exists to source the water from a close-by river via an existing old pipe already connected to the house, or to use an old ground water bore close to the house. The owner employs a contractor to find out the water quality from both sources. The contractor drives on a rainy day for three hours from Melbourne to the house but forgets his sampling kit and bottles. The contractor takes a couple of clean plastic bottles from the owner and rushes to take a grab sample from the river. Then, the contractor comes back to the house, opens the bore tap and immediately takes a grab sample in a plastic bottle. He then drives straight back to Melbourne, to take the samples to the laboratory for analysis. However, the laboratory is closed, so he leaves the bottles with some instructions next to the sample drop-off area.

The laboratory results are shown in the following table. Some relevant guideline values are included for comparison.

Analyte	River water mg/L	Ground water, mg/L	Standard for drinking	Aesthetic guideline
Colour, HU	25	18	–	15
Total dissolved solids	100	3200	500	–
Arsenic	< 0.01	< 0.01	0.01	–
Cadmium	0.001	0.002	0.002	–
Copper	0.5	4	2	1
Chromium	0.01	0.08	0.05	–
Iron	0.2	4.5	–	0.3
Zinc	0.8	3.5	–	3

a. Identify five sampling protocols that have not properly been followed when collecting the samples.
5 marks

b. Identify three laboratory analysis protocols that have not properly been followed when analysing the data.
3 marks

c. Identify one contaminant and its water quality implications. Also consider contaminants not listed in the sample results. Discuss the likely contamination pathway for that contaminant. Propose a resolution.
3 marks

15.4 Exercise 4: studyON Topic Test

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16 Measurement of solubility and concentration

16.1 Overview

Numerous **videos** and **interactivities** are available just where you need them, at the point of learning, in your digital formats, learnON and eBookPLUS at www.jacplus.com.au.

16.1.1 Introduction

Most salts and polar gases are soluble in water. Their solubility is influenced by the temperature of the solution. This relationship between solubility and temperature is demonstrated using solubility curves, which are commonly used as quantitative and predictive tools in the scientific world. Solubility curves are introduced in this topic, and their use in calculating and predicting when a solute will dissolve or crystallise out of solution. The various units of solution concentration are also introduced.

Solubility curves allow the determination of how much solute can be dissolved in a solution. However, it is possible to increase the level of dissolved solute beyond what would be expected. This is called a supersaturated solution, and the process relies upon the relationship between solubility and temperature. Generally, as the temperature increases, the solubility of a substance increases; that is, more solute can be dissolved in a solution at higher temperatures. If the solution is cooled at a faster rate than precipitation, the solution will become supersaturated, and then precipitate once the temperature drops. Supersaturated solutions can be seen in nature in the Dead Sea. Applications of this include in drug delivery, allowing medicines with a very low solubility to be made into aqueous form, and delivered in very precise doses.

FIGURE 16.1 Salt crystals precipitating out from the salt saturated Dead Sea.

16.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- the use of solubility tables and experimental measurement of solubility in gram per 100 g of water
- the quantitative relationship between temperature and solubility of a given solid, liquid or gas in water
- the use of solubility curves as a quantitative and predictive tool in selected biological, domestic or industrial contexts
- the concept of solution concentration measured with reference to moles (mol L^{-1}) or with reference to mass or volume (g L^{-1}, mg L^{-1}, %(m/m), %(m/v), %(v/v), ppm, ppb) in selected domestic, environmental, commercial or industrial applications, including unit conversions.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of learning and assessment. Experiments and investigations, supported by a **Practical investigation logbook** and **Teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

on Resources

Digital documents Key science skills (doc-30903)
Key terms glossary — Topic 16 (doc-30915)
Practical investigation logbook (doc-30916)

studyon

To access key concept summaries and practice exam questions download and print the **studyON: Revision and practice exam question booklet** (doc-30917).

16.2 Measuring solubility and the use of solubility curves

KEY CONCEPTS

- The use of solubility tables and experimental measurement of solubility in gram per 100 g of water
- The quantitative relationship between temperature and solubility of a given solid, liquid or gas in water
- The use of solubility curves as a quantitative and predictive tool in selected biological, domestic or industrial contexts

Topic 12 introduced the concepts of solubility — a chemical property of a known amount of solid, liquid or gas to dissolve in a known amount of solvent. In this topic, the solvent is water. The substances to be dissolved in water are known as solutes.

16.2.1 Solubility of solids in water

What do soft drinks, coffee, tea, soup and cordial have in common? They are all mixtures and they all contain water. Some of them are heterogeneous mixtures (known simply as ‘mixtures’) and others are homogeneous mixtures called **solutions**. For example, when sodium chloride is dissolved in water, sodium chloride is the solute and water is the solvent. The end result is called the solution.

The solvent properties of water make it an excellent medium for chemical reactions. Water samples containing dissolved substances are called **aqueous solutions**.

Most salts are soluble in water to some extent (see an example of solubility of various salts in table 16.1). Solubility is influenced by the temperature of the solution. Most salts become more soluble at high temperature, but a few become less soluble as the temperature increases. Salts with a solubility greater than 0.1 M of water are classed as soluble and salts with a solubility less than 0.1 M of water are classed as insoluble.

A **saturated solution** is one that contains the maximum amount of solute for the volume of solution at a particular temperature. When less than the maximum amount of solute has been added to the solvent, the solution is called an **unsaturated solution**.

A **supersaturated solution** can be prepared by slowly cooling a saturated solution. When this is done, the solution contains more solute than it should at that temperature. The addition of a small crystal seed or even some dust causes the excess solute to crystallise. Increasing the temperature of a supersaturated solution will increase the solubility, leading to more of the solid to be dissolved and to obtain a saturated solution. Alternatively, more water could be added, so more solid can dissolve.

TABLE 16.1 The solubility of various salts

Salt	Formula	Solubility (g per 100 g water at 20 °C)
Barium chloride	$BaCl_2$	36.0
Barium sulfate	$BaSO_4$	0.000 24
Calcium chloride	$CaCl_2$	74.0
Calcium sulfate	$CaSO_4$	0.21
Copper(II) sulfate	$CuSO_4$	20.5
Copper(II) sulfide	CuS	0.000 03
Lead(II) sulfate	$PbSO_4$	0.004
Potassium chlorate	$KClO_3$	7.3
Potassium nitrite	KNO_2	300.0
Silver chloride	AgCl	0.000 000 1
Silver nitrate	$AgNO_3$	217.0
Sodium chloride	NaCl	36.0
Sodium nitrate	$NaNO_3$	87.0

Use of solubility curves in industry

Solubility curves are used in industry to manage the production of chemicals. For example, unnecessarily high temperatures provided to unsaturated solutions result in higher energy usage (and therefore higher costs) than required. This can be avoided by reducing the temperatures of unsaturated solutions to the saturation point determined from the solubility curves.

Solubility tables or graphs allow us to predict whether a salt is soluble or insoluble. An example of a solubility table was presented in topic 12, table 12.1, and was used to identify a likely precipitate from the mixing of two or more solutions. In many industries, knowledge of likely precipitates can be used to achieve a desired goal. In topic 17, we will see how gravimetric analysis, an important analytical technique, relies on the ability to form suitable precipitates. In a popular method for large-scale water purification, aluminium hydroxide precipitate is formed *in situ* by mixing alum (potassium aluminium sulfate, $KAl(SO_4)_2$), and limewater (calcium hydroxide, $Ca(OH)_2$) together. Freshly prepared aluminium hydroxide is a gelatinous precipitate that adsorbs and traps particles of dirt and bacteria from the impure water as it settles. Subsequent filtering easily removes the precipitate and these undesired particles. A solubility table would also highlight that one of the other possible products, calcium sulfate, is only slightly soluble; this would need to be taken into account when calculating the amounts of alum and limewater that need to be added.

16.2.2 Solubility of liquids and gases

As discussed in Topic 12, the solubility of solids, liquids and gases depends on the polarity of the solute. Liquids that dissolve in water (such as ethanol) are polar molecules and are said to be *miscible*. Liquids that do not dissolve (such as oil in water) are non-polar molecules and will form layers if left to settle. Such liquids are termed *immiscible*. With gases, the situation is the same. That is, polar gases, such as methane are more soluble in water than non-polar gases like oxygen. It should also be noted that gases usually behave in the opposite way to solids as far as temperature is concerned — that is, as the temperature increases, the solubility of a gas decreases.

Environmentally, this has the potential to become a significant problem for the Earth. As atmospheric temperatures increase due to higher levels of greenhouse gases (such as carbon dioxide), the upper layers of the oceans also increase in temperature. This leads to reduced amounts of gases such as oxygen and carbon dioxide being dissolved in the water and significant amounts of these gases being returned to the atmosphere. This would further increase the level of carbon dioxide in the atmosphere and would seriously hamper respiration and photosynthesis for marine organisms due to the lower amounts of dissolved oxygen and carbon dioxide that remain.

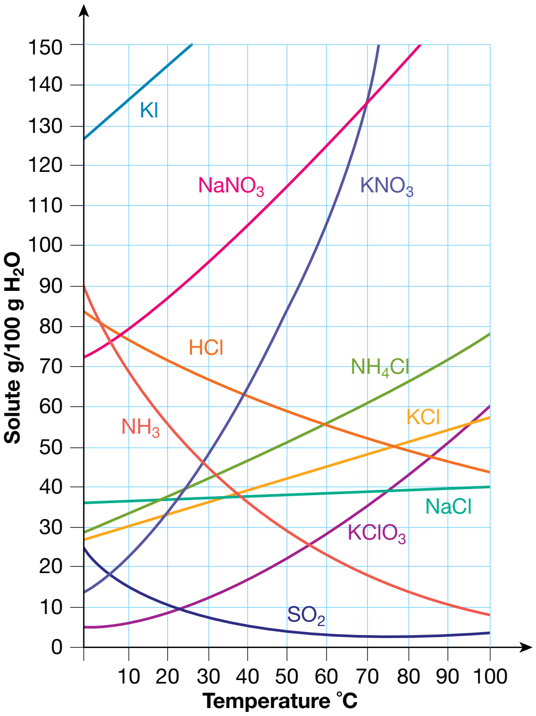
16.2.3 Solubility curves

Depending on the amount of solute present, solutions may be described as saturated, unsaturated or supersaturated.

The solubility of a substance in water depends on temperature. Most solids are more soluble as temperature increases. The solubility of soluble substances can be easily measured by experiment. A convenient unit used for this purpose is g/100 g of water. When this is done across a range of temperatures, the result may be drawn as a graph of solubility versus temperature. Such graphs are called **solubility curves**. These find extensive use in medical, pharmaceutical and industrial applications where compounds need to be isolated and purified.

In this context, 100 mL of water is the same as 100 g water. Several examples of solubility curves are shown in figure 16.2. Most substances in these examples increase their solubility as the water temperature increases, except for some gases.

FIGURE 16.2 Solubility curves for several substances



Important characteristics for interpreting solubility curves include the following:

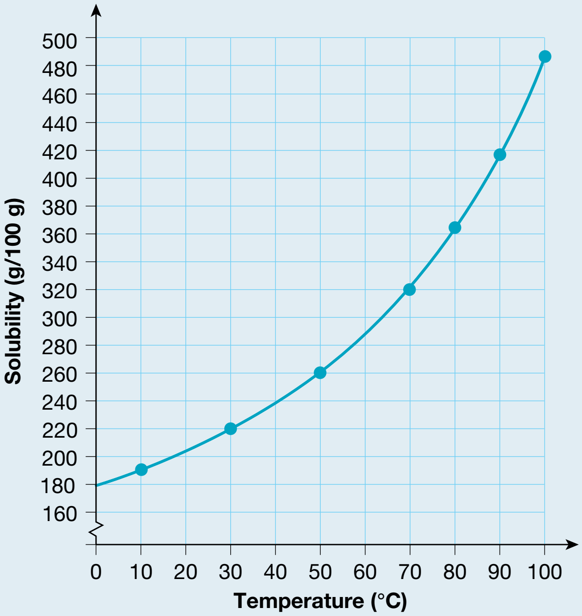
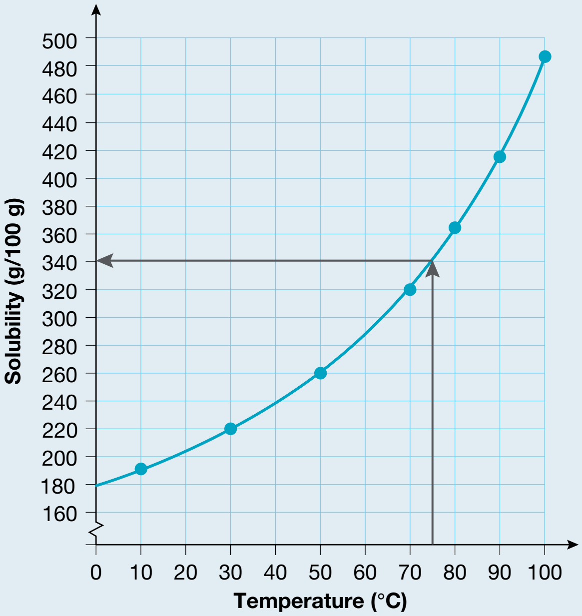
- All the points on the solubility curve represent the maximum amount of solute that can be dissolved in 100 g water at the specific temperature. Therefore, any point on the curve represents a *saturated solution*.
- Any point *above* the curve represents the amount of substance in a *supersaturated solution*, where only a part of the substance is dissolved, while some part is still found as solid.
- Any point *below* the curve represents the amount of substance completely dissolved, in an *unsaturated solution*.

SAMPLE PROBLEM 1

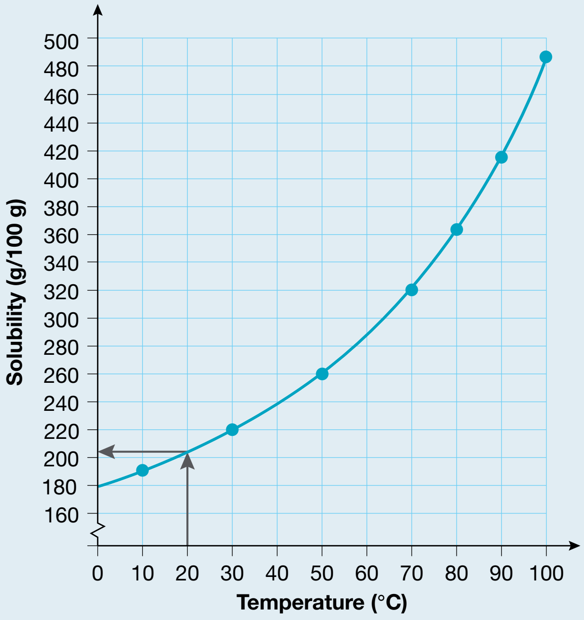
The solubility of a white crystalline substance known to be a type of sugar was determined over a range of temperatures. The results are shown in the following table.

Temperature (°C)	10	30	50	70	80	90
Solubility (g/100 g)	190	220	260	320	362	415

a. **Plot the solubility curve for this sugar.**
b. **What mass of sugar dissolves in 100 g of water at 75 °C?**
c. **What mass of sugar dissolves in 250 g of water at 20 °C?**
d. **What mass of water is required to dissolve 110 g of sugar at 50 °C?**

THINK	WRITE
a. Read values from the table and create a graph using appropriate scales, with the *x*-axis showing Temperature °C units and the *y*-axis showing g/100 g units. Plot the graph.	
b. Starting from 75 °C on the *x*-axis, draw a line to the curve, and then from the curve to the *y*-axis, to determine mass of sugar dissolved in 100 g water. Read the answer off the graph.	 At 75 °C, solubility = 340 g/100 g (from graph). Therefore at 75 °C, 340 g sugar dissolves in 100 g of water.

c. 1. Starting from 20 °C on the *x*-axis, draw a line to the curve, and then from the curve to the *y*-axis, to determine mass of sugar dissolved in 100 g water. Read the answer off the graph.



At 20 °C, solubility = 204 g/100 g.

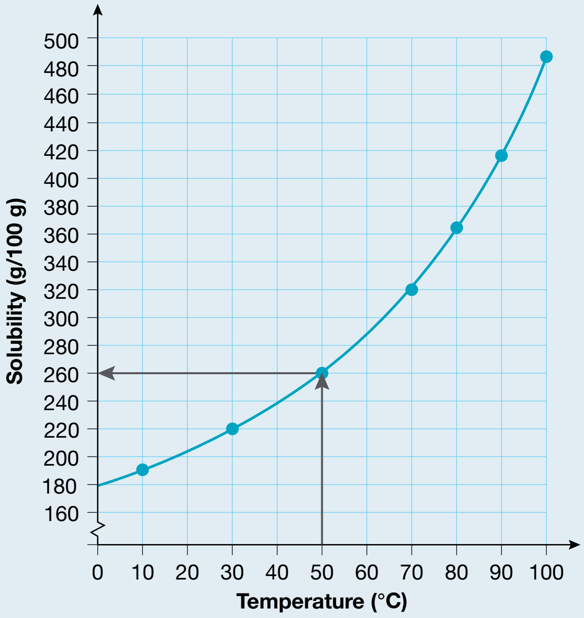
2. Determine the amount of sugar in 250 g of water.

For 250 g water:

$$204 \times \frac{250}{100} = 510\,\text{g}$$

Therefore, at 20 °C, 510 g sugar dissolves in 250 g water.

d. 1. Starting from 50 °C on the *x*-axis, draw a line to the curve, and then from the curve to the *y*-axis, to determine mass of sugar that can be dissolved in 100 g water.



At 50 °C, solubility = 260 g/100 g.

2. The actual mass of sugar required to be dissolved (110 g) at 50 °C is less than the maximum that can be dissolved (260 g). To determine how much water is required to dissolve 110 g sugar at 50 °C, ratios can be used.

Therefore, $100 \times \frac{110}{260} = 42.3\,\text{g}$ of water is required at this temperature.

PRACTICE PROBLEM 1

The solubility of NaCl was determined over a range of temperatures. The results are shown in the following table.

Temperature (°C)	0	10	20	30	40	50
Solubility (g/100 g)	35.7	35.8	36.0	36.3	36.6	37.0

a. **Plot the solubility curve for NaCl.**
b. **What mass of NaCl dissolves in 100 g water at 35 °C?**
c. **What mass of NaCl dissolves in 150 g water at 10 °C?**

SAMPLE PROBLEM 2

Discuss two analytical methods you can use under standard laboratory conditions (SLC) to obtain an unsaturated solution of $NaNO_3$ from 1.53 mols of $NaNO_3$/100 g (use figure 16.2).

THINK

1. As solubility curves are shown in g/100 g, the mass of $NaNO_3$ in the 1.53 mols of $NaNO_3$ must be determined.

   Recall the formula $n = \frac{m}{M}$ and calculate the mass of $NaNO_3$ per 100 g H_2O.

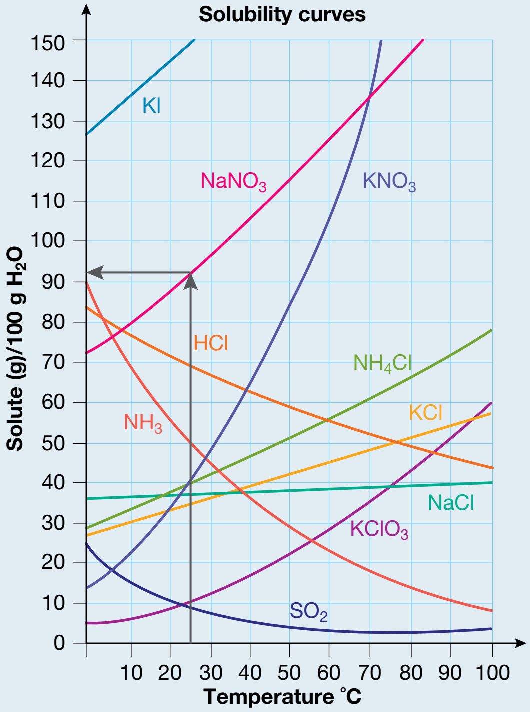
WRITE

$$n = \frac{m}{M}$$

$$\begin{aligned} m &= n \times M \\ &= 1.53 \times (23 + 14 + 3 \times 16) \\ &= 1.53 \times 85 \\ &= 130\text{ g } NaNO_3/100\text{ g } H_2O \end{aligned}$$

2. Use figure 16.2 to determine solubility of $NaNO_3$ at 25 °C (laboratory conditions) and identify the state of the solution.

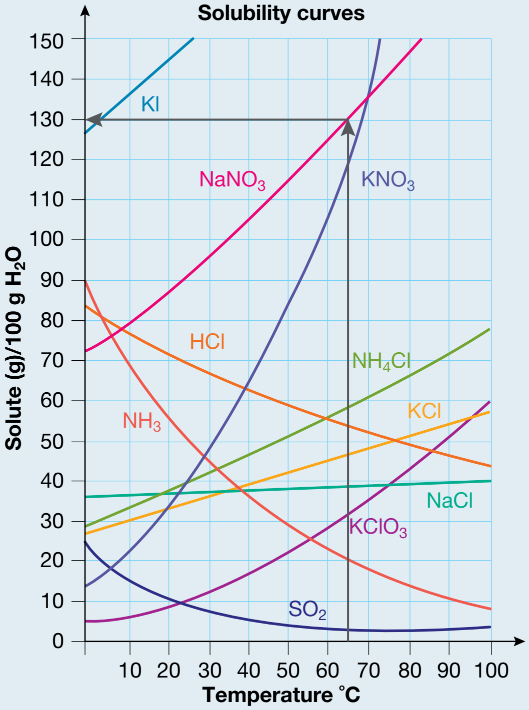
   From 25 °C on the *x*-axis, draw a line until it intersects with the solubility curve for $NaNO_3$. From that point draw a horizontal line to intersect with the *y*-axis. This will determine the amount of grams solute per 100 g H_2O.



3. Identify the state of the solution.

At 25 °C, solubility of $NaNO_3$ is around 93 g $NaNO_3$/100 g H_2O.
Therefore, a sample of 130 g $NaNO_3$/100 g H_2O will form a supersaturated solution.

4. Answer the question. Since the solubility graphs represent mass of solute with per 100 g H_2O versus temperature, think of changes to either the temperature or the volume of water to obtain an unsaturated solution. (In this case, the concentration was not required to be kept constant.)
   a. Option 1: consider the effect of a change in temperature on the solubility curve. To obtain an unsaturated solution of $NaNO_3$, the temperature would need to be increased to at least 65 °C, so all $NaNO_3$ can dissolve.
   b. Option 2: consider the effect of a change of water volume while maintaining temperature. Determine how much water is required to be added to the sample to obtain an unsaturated solution.



Option 1: The $NaNO_3$ sample would need to be heated to at least 65 °C.
Option 2: Since solubility of $NaNO_3$ at 25 °C is 93 g/100 g H_2O, 130 g $NaNO_3$ would need:

$$130 \times \frac{100}{93} = 139.78 = 140\,\text{g water to become}$$

completely dissolved.

So, at least 40 g of H_2O must be added to the supersaturated solution.

PRACTICE PROBLEM 2

Using figure 16.2 and working at SLC, discuss a method to obtain an unsaturated solution of KCl from 0.700 mols of KCl/100 g.

16.2.4 Crystallisation

Crystallisation is a technique that is used to isolate substances and is based on the difference in solubility between two temperatures. As a hot solution is cooled, it eventually reaches a temperature at which it is saturated. Further cooling results in it becoming supersaturated — an unstable situation. Crystals then precipitate out of the solution and, the more it is cooled, the more crystals are formed. These crystals can then be collected easily by filtration. This technique can also be used to purify substances. Solubility curves are an important predictive tool for both these processes.

Sample problem 3 illustrates this technique.

Evaporative crystallisation

An example of natural crystallisation through evaporation takes place around hypersaline seas situated in a dry and warm climate. The Dead Sea is known to display formation of salt bodies growing from water during certain periods of the year (see figure 16.3). The salinity of the Dead Sea is 33.7 g/100 g water.

As seen in table 16.2, solubility of NaCl is only minimally affected by temperature increase. Climate characteristics around the Dead Sea (such as average temperatures between 32 °C and 40 °C for half of the year, very low humidity and a low 50 mm annual rainfall) lead to water evaporation to a point where crystallisation occurs. The crystals dissolve as soon as the water is replenished during the wetter months.

FIGURE 16.3 Salt formation at the Dead Sea

TABLE 16.2 Solubility of $AgNO_3$ and NaCl salts at various temperatures

$AgNO_3$		NaCl	
Temperature (°C)	Solubility (g of salt per 100 g of water)	Temperature (°C)	Solubility (g of salt per 100 g of water)
0	122	0	35.7
10	170	10	35.8
20	222	20	36.0
30	300	30	36.3
40	376	40	36.6
50	455	50	37.0

Reduced-pressure crystallisation

Crystallisation is used extensively in industry. However, because heating, especially on a large scale, is a major cost, a modification of the process described previously is often employed. If pressure is reduced, water boils at a lower temperature. It also evaporates faster under such conditions. By reducing the pressure, the water may, therefore, be ‘sucked’ out of solution, making the remaining solution more and more concentrated. Once the solution becomes saturated, removing further water from the solution results in the formation of crystals. Using reduced pressure to achieve crystallisation, rather than heating and then cooling, is more cost effective in some situations.

Domestic application

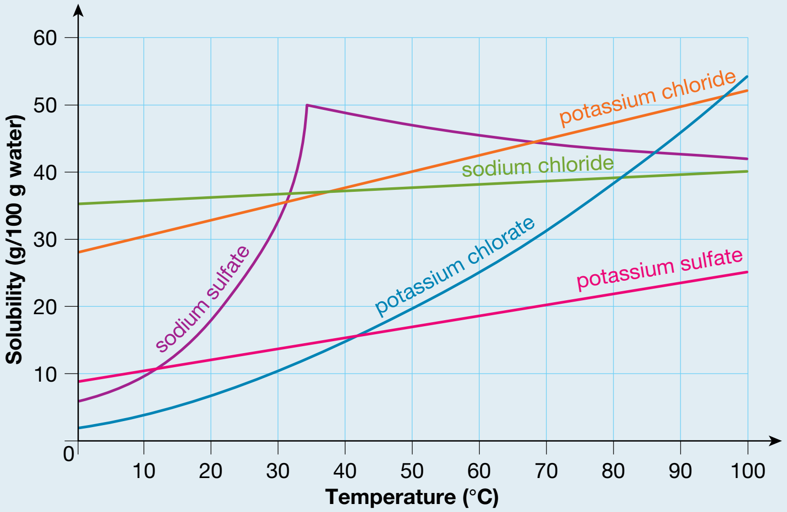
The process of crystallisation is also used to generate heat. One such reaction is the use of a supersaturated solution of sodium acetate and a metal (or crystal) that helps induce the crystallisation. This process provides instant heat, because the reaction is exothermic. This reaction can be used for single-use heat packs or reusable heat packs.

Solubility curves are an important predictive tool for both these processes. Sample problems 3 and 4 illustrate these techniques.

FIGURE 16.4 Reusable heat pack

SAMPLE PROBLEM 3

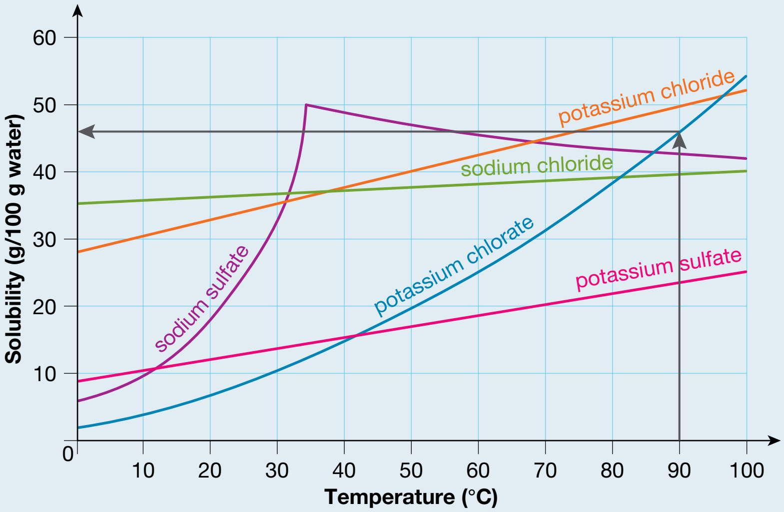
A solution of potassium chlorate ($KClO_3$) contains 40 g dissolved in 100 g of water at 90 °C. It is cooled to 20 °C. What mass of crystals form? Use the following figure to determine your answer.



THINK

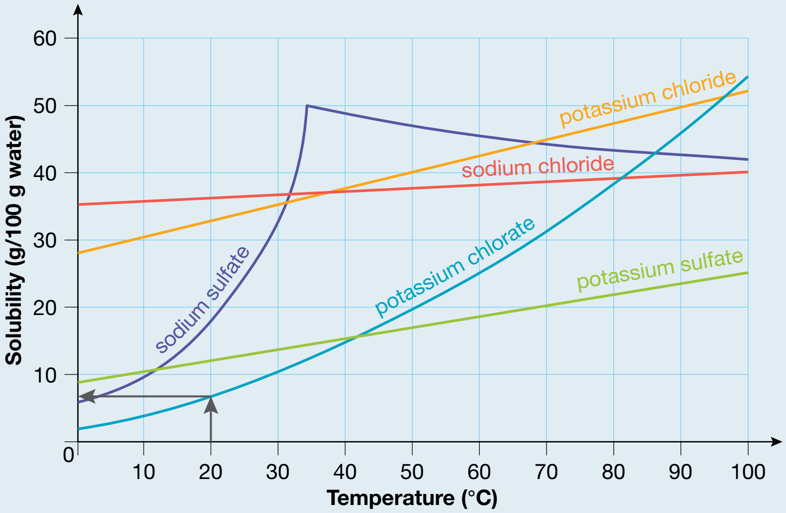
WRITE

1. Identify the curve for potassium chlorate. From 90 °C on x-axis, draw a vertical line to intersect the curve. Then draw a horizontal line to the y-axis to determine the solubility of $KClO_3$ per 100 g H_2O. Identify the state of solution.



At 90 °C, solubility equals 46 g/100 g. Therefore, all the 40 g of $KClO_3$ is dissolved, and the solution is unsaturated.

2. Repeat the same steps for 20 °C and determine the state of the solution.



At 20 °C, solubility equals 6 g/100 g.
Therefore, only 6 g can remain dissolved, and the solution is supersaturated.

3. Answer the question. Recall the mass of $KClO_3$ at the start and calculate the difference with maximum mass that can be dissolved at 20 °C.

The mass of crystals formed

$$m(KClO_3) = 40 - 6$$
$$= 34 \text{ g.}$$

PRACTICE PROBLEM 3

A mixture is known to contain 46 g of potassium chloride and 12 g of potassium sulfate. The mixture is dissolved in 100 g of water at 80 °C. Use the solubility curve in sample problem 3 to complete the following:

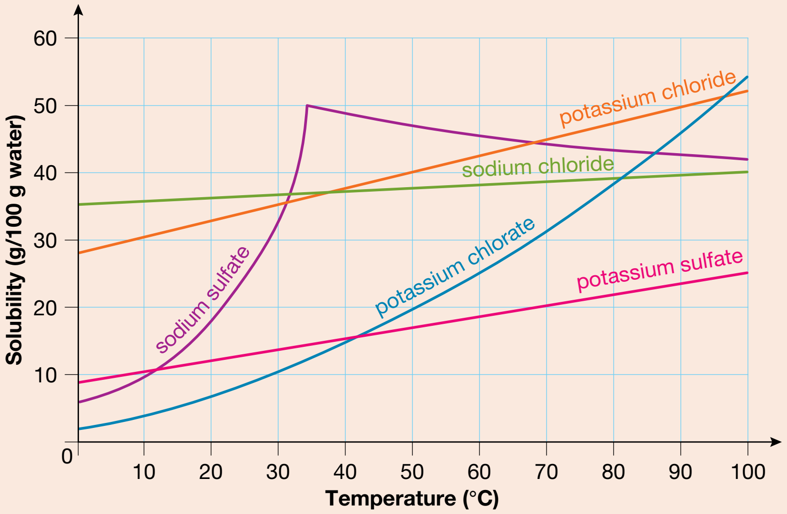
a. Do both these salts dissolve at this temperature?
b. The solution is cooled to 20 °C. Describe what you would expect to happen.
c. If you wish to obtain a pure sample of potassium chloride, why would it be inadvisable to cool the solution below 20 °C?

on Resources

Digital document Experiment 16.1 Determine a solubility curve (doc-30852)
Teacher-led video Experiment 16.1 Determine a solubility curve (tlvd-0639)

16.2 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.
Use the following solubility curves to answer questions 1 to 3.



1. Determine the amount of potassium chlorate that dissolves in 100 g of water at 35 °C.
2. What mass of potassium chloride dissolves in 50 g of water at 80 °C?
3. Which salt is most soluble at 40 °C?
4. Use the information in table 16.2 to draw solubility curves for silver nitrate and sodium chloride on the same set of axes.

Use figure 16.2 to answer questions 5 to 8.

5. 60 mL of a saturated solution of NH_4Cl contains 25 g of solute. What temperature is it at?
6. What mass of water would be needed to completely dissolve 300 g of $NaNO_3$ at 60 °C?
7. At what temperature is the solubility of HCl 27 g/50 g water?
8. How many mols of KNO_3 could dissolve in 40 mL of water at 35 °C?
9. Using the solubility values provided in sample problem 1, determine the following:
   (a) What type of solution will 260 g of sugar form at 55 °C?
   (b) What mass of sugar dissolves in 150 g water at 50 °C?
   (c) What mass of water is required to dissolve 245 g of sugar at 20 °C?
   (d) What type of solution would form from 250 g of sugar/100 g water at 20 °C?
10. The solubility of two gases is shown in the following table.

Solubility of gases at the specified temperature (g/100 g H_2O)			
Gases	0 °C	20 °C	50 °C
N_2(g)	0.0029	0.0019	0.0012
O_2(g)	0.0069	0.0043	0.0027

   (a) Identify the trend of solubility for nitrogen and oxygen gases.
   (b) Carbon dioxide (CO_2) is 89 times more soluble than N_2(g) at 20 °C and 28 times more soluble than O_2(g) at 50 °C. How much CO_2 could dissolve in 150 mL of water at 20 °C?
   (c) Discuss the trend in solubility of CO_2 versus temperature.

studyon

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16.3 Solution concentration

KEY CONCEPT

- The concept of solution concentration measured with reference to moles (mol L^{-1}) or with reference to mass or volume (g L^{-1}, mg L^{-1}, %(m/m), %(m/v), %(v/v), ppm, ppb) in selected domestic, environmental, commercial or industrial applications, including unit conversions

16.3.1 Units conversion

FIGURE 16.5 Solutions with decreasing concentration from left to right (the lightest coloured flask has $\frac{1}{16}$th of the concentration of the darkest flask).

Concentration is usually defined in terms of the amount of solute per volume of solvent. The greater the quantity of solute added to the solvent, the more concentrated the solution becomes. For example, the more sugar you dissolve in your cup of coffee, the higher the concentration of sugar in the coffee solution. When only a very small amount of solute is present, the solution may be called **dilute** (that is, its concentration is relatively low). Often, we use the terms ‘concentrated’ and ‘dilute’ comparatively. A concentrated solution of hydrochloric acid, for example, contains 37 g of HCl in 100 g of solution. If the same volume of the solution contained only 10 g of HCl, it would be called a dilute solution of hydrochloric acid. The unit of concentration depends on the units for the quantity of solution and the quantity of solute.

A concentrated solution may be diluted by the addition of more solvent.

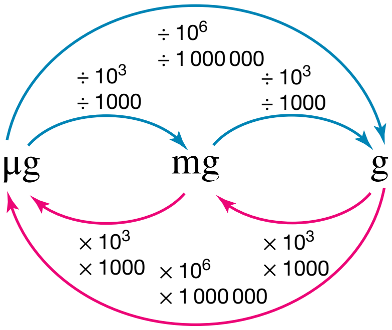
The concentration of a solution is a measure of how much solute has been dissolved in a certain volume of solvent. It is often expressed as a ratio of the solvent amount unit to the volume unit. A number of different units are typically used.

16.3.2 Common units of concentration

Common units of concentration include:

- grams per litre (g L^{-1})
- milligrams per litre (mg L^{-1})
- parts per million (ppm)
- micrograms per gram (μg g^{-1})
- micrograms per litre (μg L^{-1})
- parts per billion (ppb).

FIGURE 16.6 How to convert between different units of mass



16.3.3 Concentration with reference to mass or volume

The concentration of a solution can be represented in various units, and the conversion of those units becomes very important. Mass is measured in grams (g), milligrams (mg) and micrograms (*μ*g).

Volumes of liquids are measured in litres (L) and millilitres (mL), where 1000 mL = 1 L. Remember that 1000 L is equal to 1 cubic metre (1 m^3).

TIP: Comparison of metric prefixes can be found in table 6 of the VCE Chemistry Data Book.

Mass conversions	*Volume conversions*
1 g = 1000 mg = 1 000 000 μg	**$1\ \mathrm{mL} = 10^{-3}\ \mathrm{L}$**
$1\ \mathrm{g} = 10^{3}\ \mathrm{mg} = 10^{6}\ \mu\mathrm{g}$	**$1\ \mathrm{L} = 1000\ \mathrm{mL} = 10^{3}\ \mathrm{mL}$**
1 μg = 0.001 mg = 0.000 001	**$1000\ \mathrm{L} = 1\ 000\ 000\ \mathrm{mL} = 1\ \mathrm{m}^{3}$**
$1\ \mu\mathrm{g} = 10^{-3}\ \mathrm{mg} = 10^{-6}\ \mathrm{g}$	**$10^{3}\ \mathrm{L} = 10^{6}\ \mathrm{mL} = 1\ \mathrm{m}^{3}$**

Concentration in grams per L (g L^{-1}; mg L^{-1})

There are many ways to express concentration units. A common form is the ratio of the units of solute to the units of volume (for example, g L^{-1}, mg L^{-1}).

$$\textbf{Concentration } (c) \textbf{ in grams per litre } (\mathrm{g\,L^{-1}}) = \frac{\textbf{mass of solute in grams}}{\textbf{volume } (V) \textbf{ of solution in litres}}$$

$$\textbf{Concentration } (c) \textbf{ in milligrams per litre } (\mathrm{mg\,L^{-1}}) = \frac{\textbf{mass of solute in milligrams}}{\textbf{volume } (V) \textbf{ of solution in litres}}$$

When the concentration unit mg L^{-1} is required, the mass in grams is converted to mass in milligrams by multiplying by 1000. For smaller solute amounts, mg L^{-1} is usually more convenient to use.

SAMPLE PROBLEM 4

A fence-post preservative solution is prepared by dissolving 4.00 g of zinc chloride in enough water to make 2250 mL of solution. Find the concentration (c) of this solution in g L^{-1} and then in mg L^{-1}.

Teacher-led video: SP4 (tlvd-0585)

THINK	**WRITE**
1. Compare units given to units required. The volume of the solution is given in mL, while the concentration is required in g L^{-1}. We need to convert from mL to L.	Convert 2250 mL to a volume in litres. Since 1000 mL = 1 L: $V = \frac{2250}{1000} = 2.250\,\mathrm{L}$
2. Use the equation for concentration in g L^{-1}. $c = \frac{\text{mass of solute in grams}}{\text{volume of solution in litres}}$	$c = \frac{\text{mass of solute in grams}}{\text{volume of solution in litres}}$ $= \frac{4.00}{2.250}$ $= 1.78\,\mathrm{g\ L^{-1}}$
3. The result from the previous step is in g L^{-1} and needs to be converted to mg L^{-1}. Remember, 1 g = 1000 mg L^{-1}.	$c = 1.78 \times 1000$ $= 1780\,\mathrm{mg\ L^{-1}}$

PRACTICE PROBLEM 4

Find the mass of sodium bromide required to prepare 50 mL of a 0.40 g L^{-1} solution.

SAMPLE PROBLEM 5

What volume of solution is required to dissolve 125 mg of lithium chloride to give a concentration of 0.0500 g L^{-1}?

Teacher-led video: SP5 (tlvd-0586)

THINK	WRITE
1. Compare units given to units required. The mass of the solution is given in mg, while the concentration is required in g L^{-1}. We need to convert from mg to g.	Since 1000 mg = 1 g $\text{Mass of solute} = \frac{125\text{ mg}}{1000}$ $= 0.125\text{ g}$
2. Use the equation for volume of solution (L). $V = \frac{\text{mass of solute (g)}}{\text{concentration }\left(\text{g L}^{-1}\right)}$	$V = \frac{\text{mass of solute (g)}}{\text{concentration }\left(\text{g L}^{-1}\right)}$ $= \frac{0.125}{0.0500}$ $= 2.50\text{ L}$

PRACTICE PROBLEM 5

What volume of solution is required to dissolve 425 mg of sodium chloride to give a concentration of 0.100 g L^{-1}?

Concentration in ppm

When very small quantities of solute are dissolved to form a solution, the concentration can be measured in **parts per million (ppm)**. This can be written as mg L^{-1} or µg g^{-1}. The concentration of chemicals in the environment and trace elements in the soil are often expressed in ppm. This is a method of comparing the amount of solute, measured in millionths of a gram (micrograms or µg), to the mass of solution, measured in grams. So this is really a mass per mass ratio.

$$\textbf{parts per million (ppm)} = \frac{\textbf{mass of solute in micrograms}}{\textbf{mass of solution in grams}} = \frac{\textbf{mass of solute in milligrams}}{\textbf{volume of solution in litres}}$$

Our drinking water is treated with chlorine to kill bacteria. Too much chlorine is dangerous and too little does not kill the bacteria. A safe amount is about 1 part per million (ppm), or 1 microgram of chlorine in each gram of solution. (A microgram, µg, is 10^{-6} g, or 0.000 001 g.)

Concentration in ppb

For even smaller concentrations, parts per billion (ppb) is an appropriate unit to use. This is the same as µg kg^{-1} or µg L^{-1}.

$$\textbf{parts per billion (ppb)} = \frac{\textbf{mass of solute in micrograms}}{\textbf{mass of solution in kilograms}} = \frac{\textbf{mass of solute in micrograms}}{\textbf{volume of solution in litres}}$$

A number of organic water pollutants, for example, may have their allowable levels measured in this unit. Dieldrin, a chemical previously used against soil pests and termites, has been banned for many years but may still persist in the environment and, therefore, pollute water. Australian guidelines for drinking water state that the allowable level for dieldrin is 0.3 ppb.

SAMPLE PROBLEM 6

a. If 10 g of chlorine gas is dissolved in every 2 500 000 L of water, express the concentration of the chlorine water in (i) g L^{-1} and (ii) ppm (μg g^{-1}). (*Note:* 1 mL water = 1 g.)

b. Would this water be fit to drink?

Teacher-led video: SP6 (tlvd-0587)

THINK	**WRITE**
a. Compare units given to units required.	$V = 2.5 \times 10^6$ L
i. Use the equation for concentration in g L^{-1}. $c = \frac{\text{mass of solute in grams}}{\text{volume of solution in litres}}$	$c = \frac{\text{mass of solute in grams}}{\text{volume of solution in litres}}$ $= \frac{10\text{ g}}{2.5 \times 10^6}$ $= 4 \times 10^{-6}$ g L^{-1}
ii. Use the equation for concentration in ppm to find the mass of the solute and the mass of the solution. $c = \frac{\text{mass of solute in micrograms}}{\text{mass of solution in grams}}$	*Mass of solute:* Convert g to μg; mass of solute $= 10 \times 10^6$ $= 1 \times 10^7$ μg *Mass of solution:* Convert L to g: mass of solution $= 2.5 \times 10^6 \times 10^3$ $= 2.5 \times 10^9$ g Calculate the concentration in ppm. $c = \frac{1 \times 10^7}{2.5 \times 10^9}$ $= 4 \times 10^{-3}$ ppm
b. Recall that a safe amount of chlorine in water to kill bacteria is about 1 part per million (ppm), or 1 microgram of chlorine in each gram of solution. Determine whether water is fit to drink.	The water is unfit to drink because insufficient chlorine is present to kill the bacteria. (*Note:* when the solvent is water, mg L^{-1} also equals ppm.)

PRACTICE PROBLEM 6

A mass of 15 g of a substance is dissolved in 100 m^3 of water. Calculate its concentration in g L^{-1} and ppm.

SAMPLE PROBLEM 7

If 5.00 m^3 of river water was found to contain 1.250 mg of an organic pollutant, calculate its concentration in ppb.

Teacher-led video: SP7 (tlvd-0588)

THINK	WRITE
1. Compare units given to units required. ppb is μg L^{-1}. Convert mg to μg and m^3 to L.	Mass of solute = 1.250 mg = 1250 μg And 5.00 m^3 = 5000 L
2. Use the equation for concentration in μg L^{-1} (ppb).	$c = \frac{\text{mass of solute in micrograms}}{\text{volume of solution in litres}}$ $= \frac{1250\,\mu\text{g}}{5000\,\text{L}}$ $= 0.25\text{ ppb}$

PRACTICE PROBLEM 7

Calculate the concentration in ppb of 0.3 g of a drug found in 1 m^3 of waste water.

16.3.4 Concentration in percentage by mass or volume

Concentration in percentage by mass (%m/m)

Solutions of concentrated acids and some household cleaners often carry a label showing **%(m/m)**; for example, sulfuric acid, 98%(m/m). The unit %(m/m) expresses the concentration of the solute as a percentage of the mass (or weight) of solution. So, 98% (m/m) sulfuric acid means there is 98 grams of sulfuric acid solute per 100 g of solution. We can calculate this mass percentage using the following equation.

$$\textbf{Concentration in \% (m/m)} = \frac{\textbf{mass of solute in g}}{\textbf{mass of solution in g}} \times \frac{\textbf{100}}{\textbf{1}}$$

SAMPLE PROBLEM 8

Find the concentration in %(m/m) obtained when 18.5 g of hydrochloric acid is dissolved in 50.0 g of solution.

Teacher-led video: SP8 (tlvd-0589)

THINK	WRITE
1. Compare units given to units required for the formula for %(m/m). $\%(\text{m/m}) = \frac{\text{mass of solute in g}}{\text{mass of solution in g}} \times \frac{100}{1}$ No need to convert units.	18.5 g HCl and 50.0 g solution.

2. Calculate %(m/m) by applying the formula for %(m/m).	$\%(\text{m/m}) = \frac{\text{mass of solute in g}}{\text{mass of solution in g}} \times \frac{100}{1}$ $= \frac{18.5}{50.0} \times \frac{100}{1}$ $= 37.0\%\ (\text{m/m})$

PRACTICE PROBLEM 8

Calculate the concentration in %(m/m) when 24.5 g of copper sulfate is dissolved in 50 g of solution.

Concentration in percentage by mass/volume (%m/v)

An oven cleaner may have on its label ‘active ingredient 1%(m/v) sodium hydroxide’. The expression **%(m/v)** expresses the concentration as mass of solute in grams per 100 mL of solution. We can use the following formula to perform calculations.

$$\textbf{Concentration}\ \%\ (\textbf{m/v}) = \frac{\textbf{mass of solute in g}}{\textbf{volume of solution in mL}} \times \frac{\textbf{100}}{\textbf{1}}$$

SAMPLE PROBLEM 9

Find the %(m/v) of a physiological saline solution consisting of 4.30 g of sodium chloride dissolved in 0.500 L of aqueous solution.

Teacher-led video: SP9 (tlvd-0590)

THINK	WRITE
1. Compare units given to units required for the formula for %(m/v). $\%(\text{m/v}) = \frac{\text{mass of solute in grams}}{\text{volume of solution in mL}} \times \frac{100}{1}$ Need to convert L to mL.	$V = 0.500\ \text{L} \times 1000$ $= 500\ \text{mL}\ \ \text{solution}$ $m = 4.30\ \text{g}\ \ \text{NaCl}$
2. Calculate %(m/v) by applying the formula.	$\%\ (\text{m/v}) = \frac{\text{mass of solute in grams}}{\text{volume of solution in mL}} \times \frac{100}{1}$ $= \frac{4.30}{500} \times \frac{100}{1}$ $= 0.860\%\ (\text{m/v})$

PRACTICE PROBLEM 9

Calculate the %(m/v) of 200 mg of potassium chloride in 50.0 mL of solution.

Concentration in percentage by volume (%v/v)

The percentage by volume, **%(v/v)**, expresses the concentration unit for liquid solutes as volume of liquid solute in mL per 100 mL of solution. For example, a one-litre bottle of liqueur is labelled 40% alc/vol. This means 40 mL of ethanol (alcohol) is present in each 100 mL of liqueur. Calculations can be performed using the following equation.

$$\textbf{Concentration \% (v/v)} = \frac{\textbf{volume of solute in mL}}{\textbf{volume of solution in mL}} \times \frac{\textbf{100}}{\textbf{1}}$$

SAMPLE PROBLEM 10

If a standard glass (0.200 L) of a particular brand of beer contains 9.80 mL of ethanol, calculate the concentration of the beer in %(v/v).

Teacher-led video: SP10 (tlvd-0591)

THINK	WRITE
1. Compare units given to units required for the formula for %(v/v). $\%(\text{v/v}) = \frac{\text{volume of solute in mL}}{\text{volume of solution in mL}} \times 100\%$ We need to convert L to mL.	$V = 0.200\,\text{L} \times 1000$ $= 200\,\text{mL solution}$ $V = 9.8\,\text{mL solute}$
2. Calculate %(v/v) by applying the formula.	$\%\,(\text{v/v}) = \frac{\text{volume of solute in mL}}{\text{volume of solution in mL}} \times \frac{100}{1}$ $= \frac{9.80}{200} \times \frac{100}{1}$ $= 4.90\%\,(\text{v/v})$

PRACTICE PROBLEM 10

The label of an organic vanilla essence used for cooking shows a 35%(v/v) alcohol content. Tests in the lab found that a 25 mL sample of the essence contained 8 mL organic alcohol. Is the label on the bottle correct?

16.3.5 Molar concentration

Concentration can be expressed in a number of ways. Chemists most commonly use molar concentration or **molarity**, which is the amount of solute, in moles, present in each litre of solution. The symbol for concentration is *c*. The unit of measurement for molar concentration is moles per litre, which can be expressed as M or mol L^{-1}. Therefore a 1.5 M solution has a concentration of 1.5 mol L^{-1}.

If 1 mol of solute is dissolved in a total volume of 1 L of water, the concentration of the solution is 1 M. If 0.5 mol of solute is dissolved in a total volume of 1 L of water, the concentration of the solution is 0.5 M. If, however, 0.5 mol of solute is dissolved in a total volume of 0.5 L of water, the concentration of the solution is 1 M.

Molar concentration:

$$\textbf{concentration}\ (c) = \frac{\textbf{quantity of solute (moles)}}{\textbf{volume of solution (L)}}$$
$$c = \frac{n}{V}\ (\text{M or mol L}^{-1})$$

TIP: **Table 3 of the VCE Chemistry Data Book contains the formulas $n = cV$ and $n = \frac{m}{M}$. These formulas can be manipulated to determine the concentration and subsequently the mass of solute.**

The mass of solute (m) can be determined using:

$$n = c \times V$$
$$m = n \times M$$

where:
V is volume (litres, L)
c is concentration (M or mol L^{-1})
M is molar mass.

In order to prepare a particular volume of solution of known concentration, the following five steps should be followed:

1. Calculate the number of moles of solute that are needed to obtain the correct concentration of solution for the volume of solvent to be used, according to the formula $n = c \times V$.
2. Calculate the mass of the solute needed, using the formula $m = n \times M$.
3. Partially fill a volumetric flask with water, and add the correct mass of solute.
4. Dissolve the solute.
5. Add water to the required volume.

The volume (V) required for a specific concentration can be determined using:

$$\textbf{Volume}\ (V) = \frac{\textbf{quantity of solute (moles)}}{\textbf{concentration (moles L}^{-1})} = \frac{n}{c}$$

***Note:* volume is calculated in litres.**

SAMPLE PROBLEM 11

Calculate the number of moles of sodium chloride needed to prepare 500 mL of a 0.0800 mol L^{-1} salt solution.

Teacher-led video: SP11 (tlvd-0592)

THINK	WRITE
1. List the known information, and then compare the units given to units required for the formula. $n = c \times V$. Convert mL to L.	$V = 500\ \text{mL} \div 1000$ $= 0.500\ \text{L}$ $c = 0.0800\ \text{mol L}^{-1}$
2. Calculate the number of moles (n) of NaCl needed.	$n = c \times V$ $= 0.0800 \times 0.500$ $= 0.0400\ \text{mol}$

PRACTICE PROBLEM 11

Calculate the concentration (molarity) of copper(II) sulfate, $CuSO_4$, in 1 L of solution containing 200 g $CuSO_4{\cdot}5H_2O$.

SAMPLE PROBLEM 12

What mass of NaCl would be required to prepare 0.250 L of a solution of 0.0800 mol L^{-1} solution?

Teacher-led video: SP12 (tlvd-0593)

THINK	WRITE
1. List the known information and then compare the units given to units required for the formulas that can assist. $c = \frac{n}{V}$ and $n = \frac{m}{M}$. No conversion required.	$V = 0.250\text{ L}$ $c = 0.0800\text{ mol L}^{-1}$
2. Calculate moles from concentration.	$c = \frac{n}{V}$ $n = cV$ $= 0.0800\text{ mol L}^{-1} \times 0.25\text{ L}$ $= 0.0200\text{ mol}$
3. Rearrange $n = \frac{m}{M}$ to calculate mass of NaCl.	$n = \frac{m}{M}$ $m = n \times M$ $= 0.0200\text{ mol} \times (23 + 35.5)$ $= 1.17\text{ g}$

PRACTICE PROBLEM 12

Calculate the mass of solute in 120 mL of 1.7 M $Mg(NO_3)_2$.

SAMPLE PROBLEM 13

What volume of water would be required to prepare a solution of 0.0800 mol L^{-1} with 5.00 g of NaCl?

Teacher-led video: SP13 (tlvd-0594)

THINK	WRITE
1. List the known information and then compare the units given to units required for the formulas that can assist. $C = \frac{n}{V}$ and $n = \frac{m}{M}$ No conversion required.	$c = 0.0800\text{ mol L}^{-1}$ $m = 5.00\text{ g}$

2. Calculate moles from the mass of NaCl given.

$$n = \frac{m}{M}$$
$$= \frac{5.00}{(23 + 35.5)}$$
$$= 0.0855 \text{ mol}$$

3. Use calculated moles to calculate V by rearranging formula $c = \frac{n}{V}$.

$$C = \frac{n}{V}$$
$$V = \frac{n}{c}$$
$$= \frac{0.0855 \text{ mol}}{0.0800 \text{ mol L}^{-1}}$$
$$= 1.07 \text{ L}$$

PRACTICE PROBLEM 13

A student needs to prepare a solution of 0.100 mol L^{-1} from 7.50 g of $AlCl_3$. What volume of water would they need to use?

16.3.6 Concentration of ions in solution

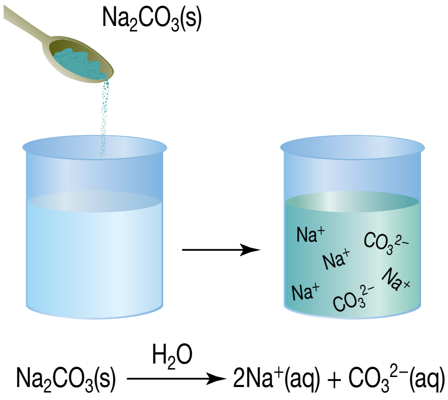
Many ionic substances are soluble. The process of dissolving involves the ionic lattice breaking up and the anions and cations dissociating from each other.

The concentrations of the resulting ions can be calculated. These concentrations are called **ionic concentrations** and they are designated by square brackets. For example:

$[Na^+] = 0.5 \text{ mol L}^{-1}$ (or M) means that the ionic concentration of Na^+ ions in solution is 0.5 mol L^{-1} (or M).

Ionic concentrations are calculated by finding the concentration of the solution, and then multiplying the mole ratio of ions in the dissociation equation.

FIGURE 16.7 Dissolving sodium carbonate brings about dissociation of the sodium and carbonate ions.



$$Na_2CO_3(s) \xrightarrow{H_2O} 2Na^+(aq) + CO_3^{2-}(aq)$$

SAMPLE PROBLEM 14

Sodium carbonate, Na_2CO_3, is often used as an ingredient in washing powders because it softens hard water. It dissolves readily in water. If 10.6 g of sodium carbonate is dissolved in 500 mL of water:

a. **what is the concentration of the solution?**

b. **what is the concentration of each ion in the solution?**

Teacher-led video: SP14 (tlvd-0595)

THINK	WRITE
a. 1. Calculate the number of moles by applying the formula $n = \frac{m}{M}$.	$n(Na_2CO_3) = \frac{m}{M}$ $= \frac{10.6}{106}$ $= 0.100$ mol

2. Calculate the concentration by applying the formula $c = \frac{n}{V}$.	$c\,(Na_2CO_3) = \frac{n}{V}$ $= \frac{0.100}{0.500}$ $= 0.200\ \text{mol L}^{-1}$
b. **1.** Write the dissociation equation.	$Na_2CO_3 \xrightarrow{H_2O} 2Na^+(aq) + CO_3^{2-}(aq)$
2. Look at equation to observe how many moles of ions are produced.	1 mol Na_2CO_3 produces 2 mol Na^+ and 1 mol CO_3^{2-}.
3. Since $c = \frac{n}{V}$ and the volume is constant, calculate concentration of each ion.	$c(Na^+) = 2 \times c\,(Na_2CO_3)$ $= 2 \times 0.200$ $= 0.400\ \text{mol L}^{-1}$ $c(CO_3^{2-}) = c\,(Na_2CO_3)$ $= 0.200 \times 1$ $= 0.200\ \text{mol L}^{-1}$

PRACTICE PROBLEM 14

In an $Al_2(SO_4)_3$ solution, the concentration of SO_4^{2-} is 0.050 mol L^{-1}. What is the concentration of solute?

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Digital document Experiment 16.2 Preparation of a solution of known concentration (doc-30853)

16.3 EXERCISE

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1. Calculate the concentration in g L^{-1} and mg L^{-1} for each of the following solutions.
   (a) 0.425 g barium nitrate in 25.0 mL of solution
   (b) 26.4 mg sodium bicarbonate in 55.0 mL of solution
2. How many grams of sodium chloride would you need to prepare 475 mL of a solution with a concentration of 0.200 g L^{-1}?
3. If you dissolve 200 mg of potassium chloride in sufficient solution to give a concentration of 50.0 g L^{-1}, what volume of solution is required?
4. Calculate the concentration in μg g^{-1} of a solution formed by dissolving 7.50 mg of aluminium sulfate in each 2.00 L of water processed by a water treatment plant (density of water = 1 g mL^{-1}).
5. A sample of pond water contained 88.00 μg g^{-1} of dissolved oxygen gas. Calculate the amount in grams of dissolved oxygen in 4.00 L of pond water.
6. A desirable level of dissolved oxygen in river water is 8 ppm. A sample of river water is found to have less than this, with 5.60 ppm oxygen. If 1.40 μg of dissolved oxygen was actually present in the sample, calculate the mass of water tested.
7. Calculate the concentration in %(m/m) for the following solutions.
   (a) 38.4 g of NaCl in 150.0 g of solution
   (b) 0.25 g benzalkonium chloride (active ingredient in an antibacterial kitchen cleaner) in 50.0 g of solution
   (c) 144 mg of sugar in 200 g of solution
8. Calculate the mass of sodium hydrogen carbonate in 250 g of a 0.500%(m/m) solution.

9. Calculate the mass of solution required to prepare a 6.00%(m/m) tile and household cleaner solution containing 15.5 g of sodium phosphate.
10. Calculate the concentration in %(m/m) when 24.5 g of copper sulfate is dissolved in 50 g of solution.
11. Calculate the %(m/v) of the following solutions.
    (a) 50 kg of sodium bromide in 250 L of solution
    (b) 0.025 g NaCl in 500 mL of solution
12. Calculate the mass of sodium hypochlorite required to prepare 375 mL of a 1.5%(m/v) solution of bathroom cleaner.
13. (a) Calculate the volume of solution required to dissolve 100 mg of plant food to make a concentration of 0.50%(m/v).
    (b) Calculate the volume of solution required to prepare a 10%(m/v) solution of sodium hydroxide from a 10 mg sample of sodium hydroxide.
14. Glycerol is a syrupy, sweet tasting liquid used in cosmetics and candy. If 10.0 mL of glycerol is dissolved in 0.455 L of aqueous solution, calculate the concentration as %(v/v).
15. Ethanoic acid is dissolved in aqueous solution to make a 5.00%(v/v) solution of vinegar.
    (a) What volume of ethanoic acid is in 0.750 L of vinegar?
    (b) What volume of vinegar contains 25.0 mL of ethanoic acid?
16. A particular brand of champagne has an alcohol concentration of 11.5%(v/v). What is the alcohol content of a 750 mL bottle?
17. Calculate the concentration (molarity) of copper(II) sulfate, $CuSO_4$, in:
    (a) 2.0 L of solution containing 250 g $CuSO_4 \cdot 5H_4O$
    (b) 700 mL of solution containing 750 g $CuSO_4 \cdot 5H_2O$
    (c) 125 mL of solution containing 50 g $CuSO_4 \cdot 5H_2O$.
18. Caustic soda, NaOH, is an ingredient in household drain cleaners. If a 250 mL bottle of cleaner contains 20 g of caustic soda, calculate the following.
    (a) The concentration of the solution
    (b) The concentration of each ion in the solution
19. Calculate the concentration of an $Mg(NO_3)_2$ solution in which the concentration of nitrate ions is 0.030 M.
20. Calculate the concentration of a solution of ammonium phosphate, $(NH_4)_3PO_4$, in which the concentration of ammonium ions is 1.25 M.

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16.4 Review

16.4.1 Summary

Measuring solubility and the use of solubility curves

- A solute is a substance that is dissolved in a solvent to make a solution. A dilute solution contains a small amount of solute. Volumes of water containing dissolved substances are called aqueous solutions.
- Solubility is the extent to which a solute can dissolve in a solvent and this depends on the temperature.
    - A saturated solution contains the maximum amount of solute that can be dissolved for the volume of solution at a given temperature.
    - An unsaturated solution contains less than the maximum amount of solute that can be dissolved.
    - A supersaturated solution is prepared by slowly cooling a saturated solution so that it contains more solute than normal at that temperature.
- A solubility curve is a graph of temperature versus solubility. It can be used to predict how much solute dissolves at a given temperature.
- Increasing temperatures in the upper levels of the oceans reduces the amount of dissolved carbon dioxide and oxygen, which will have serious implications for life in these regions.

Solution concentration

- Concentration is defined as the amount of solute per volume of solution. When more solute is added to the solution, a concentrated solution results. When more solvent is added to the solution, a dilute solution results.
- Concentration units can be expressed as $g\ L^{-1}$, $mg\ L^{-1}$, $g\ \mu L^{-1}$, parts per million (ppm), parts per billion (ppb), %(m/m), %(m/v) and %(v/v). Molar concentration (M) is an especially useful unit in chemistry.

Concentration units	Formula
$g\ L^{-1}$	$c = \frac{\text{mass of solute (g)}}{\text{volume of solution (L)}}$
$mg\ L^{-1}$	$c = \frac{\text{mass of solute (mg)}}{\text{volume of solution (L)}}$
ppm	$\text{ppm} = \frac{\text{mass of solute } (\mu g)}{\text{mass of solution (g)}} = \frac{\text{mass of solute (mg)}}{\text{volume of solution (L)}}$
ppb	$\text{ppb} = \frac{\text{mass of solute } (\mu g)}{\text{mass of solution (kg)}} = \frac{\text{mass of solute } (\mu g)}{\text{volume of solution (L)}}$
%(m/m)	$\%(m/m) = \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \times \frac{100}{1}$
%(m/v)	$\%(m/v) = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times \frac{100}{1}$
%(v/v)	$\%(v/v) = \frac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}} \times \frac{100}{1}$
Molar concentration, M or $mol\ L^{-1}$	$c = \frac{\text{quantity of solute } (n)}{\text{volume of solution } (V)}$; V in litres

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16.4.2 Key terms

aqueous solutions mixtures in which substances are dissolved in water
concentration the amount of solute that is dissolved in a known volume of solvent
dilute not concentrated
ionic concentrations concentrations of the anions and cations resulting from the dissolution of an ionic lattice; designated by square brackets
molarity concentration of a solute in a solution in units of M, or mol L^{-1}
parts per million (ppm) number of a particular component present within one million objects
saturated solution solution in which the maximum amount of solvent has been dissolved
solubility curves graphs of grams of a substance that will dissolve in 100 g of solvent at a particular temperature
solutions solute dissolved in a solvent
supersaturated solution solution in which a greater amount of solute is dissolved at a particular temperature than is predicted by a solubility curve temperature
unsaturated solution solution containing less than the maximum amount of solute that can dissolve
%(m/m) percentage mass per mass
%(m/v) percentage mass per volume
%(v/v) percentage volume per volume

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Digital document Key terms glossary — Topic 16 (doc-30915)

16.4.3 Practical investigations

Experiment 16.1

Determine a solubility curve

Aim: To obtain the solubility curve for potassium chloride, KCl
Digital document: doc-30852
Teacher-led video: tlvd-0639

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Digital documents Practical investigation logbook (doc-30916)
Experiment 16.2 Preparation of a solution of known concentration (doc-30853)

16.4 Exercises

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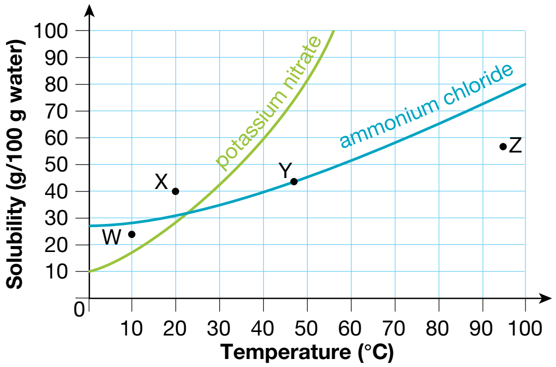
16.4 Exercise 1: Multiple choice questions

1. A saturated solution is best described as a solution in which
   - **A.** there is a small amount of solute in a large amount of solvent.
   - **B.** crystals of the solute can stay without growing or dissolving.
   - **C.** the addition of further solvent produces a precipitate.
   - **D.** there is a very large amount of solute in a small amount of solvent.

Questions 2–5 refer to the information shown in the following graph.

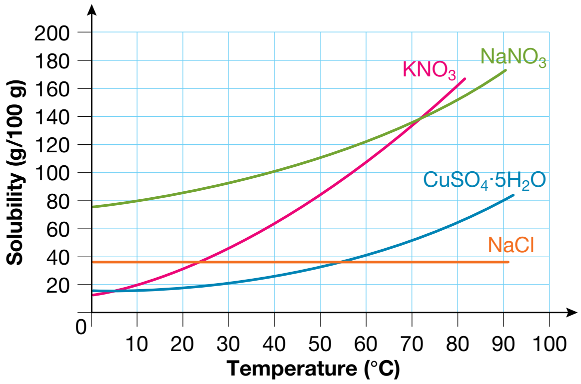


2. To completely dissolve 35.0 g of potassium nitrate in 50.0 g of water, what would the minimum temperature have to be?
   - **A.** 25 °C
   - **B.** 35 °C
   - **C.** 45 °C
   - **D.** 55 °C
3. Various solutions of ammonium chloride are represented by the letters W, X, Y and Z. Which letter represents a supersaturated solution?
   - **A.** W
   - **B.** X
   - **C.** Y
   - **D.** Z
4. Which letter corresponds to 11 g solution /25 g H_2O?
   - **A.** W
   - **B.** X
   - **C.** Y
   - **D.** Z
5. A total of 80 g of ammonium chloride was dissolved in 200 g of boiling water. If the solution were cooled slowly, when would the first crystals of ammonium chloride be expected to appear at?
   - **A.** 100 °C
   - **B.** 40 °C
   - **C.** 80 °C
   - **D.** 10 °C
6. The concentration of sodium chloride in a sample of sea water is 28.5 g L^{-1}. What would the mass of salt remaining be if 2.00 L of the sea water were evaporated to dryness?
   - **A.** 14.3 g
   - **B.** 57.0 g
   - **C.** 28.5 g
   - **D.** 42.8 g
7. What volume of water is required to dissolve 0.15 g of copper sulfate to make a solution of 250 mg L^{-1}?
   - **A.** 0.60 L
   - **B.** 1.7 L
   - **C.** 6.0×10^{-4} L
   - **D.** 1.6×10^{3} L
8. A recommended safe level of the heavy metal cadmium in drinking water is 0.01 ppm. What would the mass of cadmium in a litre of drinking water be (density of water = 1.0 g mL^{-1})?
   - **A.** 10 mg
   - **B.** 0.010 µg
   - **C.** 0.010 g
   - **D.** 0.010 mg

9. What is the %(m/v) of a 1 L solution of commercial bleach that contains 50 g of sodium hypochlorite?
   A. 50%(m/v)
   B. 2.0%(m/v)
   C. 5.0%(m/v)
   D. 0.050%(m/v)
10. A solution contains a mixture of two salts: potassium nitrate and potassium sulfate. If the concentration of potassium ions is 1.40 M and the concentration of nitrate ions is 0.600 M, what is the concentration of the sulfate ions?
   A. 0.400 M
   B. 0.800 M
   C. 0.900 M
   D. 1.800 M

16.4 Exercise 2: Short answer questions

1. An amount of 0.2 g of solute is dissolved in 500 L (500 kg) of aqueous solution. Express this in the following units.
   a. g L^{-1}
   b. ppm (μg g^{-1})
2. Consider the solubility curves shown and answer the questions.



   a. At what temperature is the solubility of $CuSO_4 \cdot 5H_2O$
      i. 30 g/100 g?
      ii. 40 g/100 g?
      iii. 20 g/100 g?
   b. At what temperature does 60 g of KNO_3 dissolve in 50 g of water?
   c. What is the solubility of the following solids at 35 °C?
      i. NaCl
      ii. $NaNO_3$
      iii. KNO_3
3. Distinguish between the terms ‘unsaturated’, ‘saturated’ and ‘supersaturated’.
4. Calculate the percentage by mass, %(m/m), of the solute in each of the following.
   a. A 45 g solution containing 5 g of potassium nitrate
   b. A 24.5 g solution containing 250 mg of ammonium nitrate
5. A solution of potassium chloride is 8.5%(m/m). An experiment requires 3.4 g of KCl. How many grams of the solution do you need for this experiment?
6. The label on a bottle says that it contains a 7.5%(m/m) solution of ammonium nitrate. The bottle contains 125 g of this solution. How many grams of ammonium nitrate are in the bottle?
7. A chemist dissolves 3.5 mL of ethanoic acid in water to give a total of 120 mL. What is the concentration in %(v/v)?
8. What volume of ethanol is required to prepare the following alcoholic solutions?
   a. 4.0 L of 10%(v/v)
   b. 350 mL of 40%(v/v)
9. A saline solution contains 2.5%(m/v). How much salt would be present in 2.0 L of this solution?

10. Calculate the mass of solute in each of the following solutions.
    a. 300 mL of 1.5 M NaOH
    b. 250 mL of 2.0 M H_2SO_4
    c. 17.5 L of 1.5 M Na_2CO_3
    d. 200 mL of 2.5 M $CuSO_4 \cdot 5H_2O$
11. Calculate the concentration (molarity) of solute in each of the following solutions.
    a. 250 mL of solution containing 17 g sodium bromide, NaBr
    b. 500 mL of solution containing 200 g magnesium sulfate, $MgSO_4$
    c. 1.5 L of solution containing 1500 g magnesium nitrate, $Mg(NO_3)_2$
    d. 100 mL of solution containing 1000 g sodium phosphate, Na_3PO_4
    e. 2.00 L of solution containing 500 g potassium carbonate, K_2CO_3
    f. 40.0 mL of solution containing 12.5 g aluminium chloride, $AlCl_3$
12. Calculate the concentrations of the following solutions.
    a. 4.0 g of sodium hydroxide in 200 mL of solution
    b. 12.6 g of sodium carbonate in 350 mL of solution
    c. 5.35 g of magnesium carbonate in 500 mL of solution
13. Calculate the number of moles of solute needed to prepare the following.
    a. 30 mL of 0.10 M $AgNO_3$
    b. 300 mL of 1.5 M $Mg(NO_3)_2$
    c. 230 mL of 0.40 M KCl
    d. 2.5 L of 0.2 M KNO_3
14. Calculate the mass of solute needed to make each of the following quantities of solution.
    a. 250 mL of a 1.5 M sodium bromide, NaBr, solution
    b. 500 mL of a 1.75 M magnesium sulfate, $MgSO_4$, solution
    c. 1.50 L of a 0.575 M magnesium nitrate, $Mg(NO_3)_2$, solution
    d. 100 mL of a 0.850 M sodium phosphate, Na_3PO_4, solution
    e. 2.00 L of a 0.00500 M potassium carbonate, K_2CO_3, solution
    f. 40.0 mL of a 2.30 M aluminium chloride, $AlCl_3$, solution
15. How many grams of CH_3COONa are obtained when 400.0 mL of a 0.500 M solution of CH_3COONa is evaporated to dryness?
16. What is the concentration of each of the following solutions?
    a. 58.5 g of H_2SO_4 dissolved in enough water to produce 2.00 L of solution
    b. 2.7 g of HCl dissolved in enough water to produce 500 mL of solution
    c. 4.04 g of KNO_3 dissolved in enough water to produce 150.0 mL of solution
    d. 234 g of sodium chloride dissolved in enough water to produce 6.00 L of solution.
17. Most common fertilisers contain nitrogen compounds. Ammonium nitrate, NH_4NO_3, is soluble in water and so is quickly taken up by the plant's root system. If a 1.5 L bucket contains 150 g of ammonium nitrate calculate the following.
    a. The concentration of the solution
    b. The concentration of each ion in the solution
18. Aluminium chloride, $AlCl_3$, is found in antiperspirants. A student wishes to do some tests on antiperspirant, so a jar of it is dissolved into a beaker containing 700 mL of water. The label on the jar states that it contains 200 g of aluminium chloride. Calculate the following.
    a. The concentration of the solution
    b. The concentration of each ion in the solution
19. Calcium hydroxide, $Ca(OH)_2$, is found in plaster and cement, and is used in the treatment of drinking water. If the concentration of OH^- of a solution of calcium hydroxide is 0.050 M, calculate the concentration of the calcium hydroxide.
20. In an $Fe_2(SO_4)_3$ solution, the concentration of SO_4^{2-} is 0.25 M. What is the concentration of solute?

16.4 Exercise 3: Exam practice questions

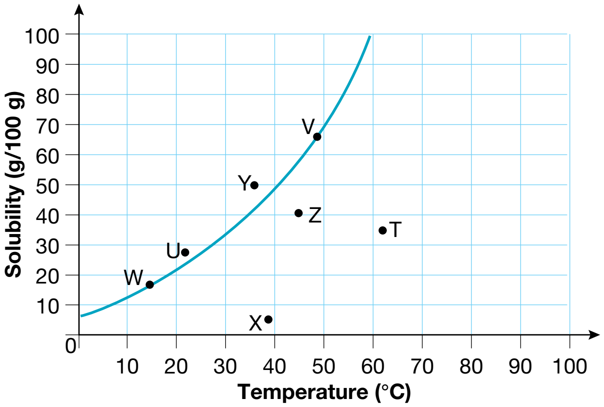
Question 1 (3 marks)

Various solutions of potassium nitrate are represented by the letters T, U, V, W, X, Y and Z on the solubility curve shown in the figure. Which letters represent

a. saturated solutions? **1 mark**

b. unsaturated solutions? **1 mark**

c. supersaturated solutions? **1 mark**



Question 2 (6 marks)

Calculate the concentration of the following aqueous solutions in (i) g L^{-1} and (ii) mg L^{-1} (ppm).

a. 3 g of solute in 100 mL solution **2 marks**

b. 0.004 g in 800 mL **2 marks**

c. 500 mg in 200 mL **2 marks**

Question 3 (7 marks)

It has been proposed that gold could be obtained from sea water. The concentration of gold varies from 0.1 to 2.0 mg per tonne.

a. Assuming that 1 tonne is equal to 1000 L of sea water, calculate the concentration of gold in parts per billion. **4 marks**

b. Assuming that sea water contains 1.0 mg of gold per tonne, calculate the volume of sea water that would need to be processed to produce a one-kilogram gold ingot. **2 marks**

c. Why do you think this proposal would be uneconomical? **1 mark**

16.4 Exercise 4: studyON Topic Test online only

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Access the Assignments section in learnON to begin creating and assigning assessments to students.

17 Analysis for salts in water

17.1 Overview

Numerous **videos** and **interactivities** are available just where you need them, at the point of learning, in your digital formats, learnON and eBookPLUS at www.jacplus.com.au.

17.1.1 Introduction

Water is an excellent solvent and as such is hardly ever found in a pure state. Water analysis is an important step in determining the type and amount of contaminants, to determine whether water is fit for use. This enables statutory bodies such as Victoria's Environment Protection Agency (EPA) to set allowable limits and to impose penalties in order to protect the environment. It also allows water that might be unsuitable for one purpose to be safely used for another, thus making more efficient use of precious water resources.

In this topic, we focus on the sources of dissolved salts in water — specifically, minerals, heavy metals and organo-metallic compounds. Analytical methods and instrumentation typically used to analyse for the dissolved contaminants are also explained, including gravimetric analysis, colorimetry, UV–visible spectroscopy and atomic absorption spectroscopy (AAS).

FIGURE 17.1 Knowing what is in a water supply may permit its use in some situations but not in others. This allows a more efficient use of water.

17.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- sources of salts found in water (which may include minerals, heavy metals, organo-metallic substances) and the use of electrical conductivity to determine the salinity of water samples
- the application of mass–mass stoichiometry to gravimetric analysis to determine the mass of a salt in a water sample
- the application of colorimetry and/or UV–visible spectroscopy, including the use of a calibration curve, to determine the concentration of coloured species (ions or complexes) in a water sample
- the application of atomic absorption spectroscopy (AAS), including the use a calibration curve, to determine the concentration of metals or metal ions in a water sample (excluding details of instrument).

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of learning and assessment. Experiments and investigations, supported by a **Practical investigation logbook** and **Teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

on Resources

Digital documents Key science skills (doc-30903)
Key terms glossary — Topic 17 (doc-30909)
Practical investigation logbook (doc-30910)

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To access key concept summaries and practice exam questions download and print the **studyON: Revision and practice exam question booklet** (doc-30911).

17.2 Chemical analysis

BACKGROUND KNOWLEDGE

- The differences between qualitative and quantitative analysis
- The method of chemical analysis is dependant upon the properties of the substance being analysed

17.2.1 Qualitative versus quantitative analysis

Chemical analysis is the process of determining what substances, and their abundances, are in a sample. Many methods are available to the analytical chemist and they can be divided into two broad groups: *qualitative analysis* and *quantitative analysis*.

Qualitative analysis

Qualitative analysis finds what substances are present. In **qualitative analysis**, the chemist is merely interested in *what* is present. In testing a urine sample from a racehorse or an athlete, for example, the mere presence of a banned drug (or its metabolised products) is all that is necessary. In the confectionery industry, a sample of imported food dye might be tested to see whether it contains chemicals that are banned in this country.

Quantitative analysis

Quantitative analysis measures the quantity (or how much) of substance present. A brewer of a low alcohol beer, for example, needs to know whether its alcohol content is below a certain limit. Health authorities might need to know whether the level of mercury in samples of fish is below the allowed level.

A logical sequence often used in many analytical procedures is to perform a qualitative analysis first to find out what is present, and then to perform a quantitative analysis to find the various amounts of substances present.

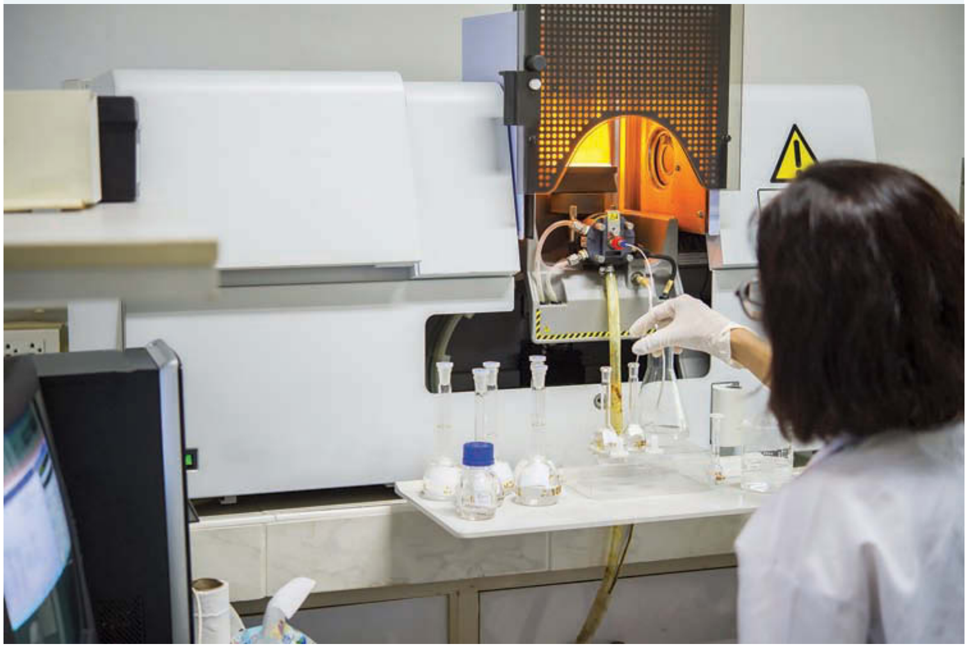
17.2.2 Methods of chemical analysis

The analytical chemist uses methods that range from sophisticated to very simple. Techniques such as gravimetric analysis (analysing by mass) and volumetric analysis (analysing with accurate concentrations and volumes of solution) may be used. These techniques may already be familiar to you through your laboratory work. More likely, however, the analyst will use advanced instrumental techniques such as atomic absorption spectroscopy (AAS) and various types of chromatography. Due to advances in technology, instruments for these techniques and have become cheaper, more powerful and more user friendly.

Deciding on an analytical method

When deciding on a method of analysis, the properties of the substance under investigation must be considered. A chemist, therefore, gives careful thought to the *way the substance reacts chemically.* Acid–base reactions are often important, while in other cases, redox reactions might be chosen as the basis for analysis. For example, to analyse a sample of oven cleaner containing the base sodium hydroxide, an acid might be used. However, a sample of bleach containing the oxidant sodium hypochlorite would most likely be analysed by reaction against a suitable reductant. In other cases, the most appropriate method for identifying and determining a substance may be to add another substance that forms a precipitate or, alternatively, something that produces a colour that can be matched against a set of standards (colorimetric analysis).

FIGURE 17.2 Scientists use atomic absorption spectroscopy (AAS) to analyse heavy metals (such as Pb, Ni, Cu, Fe, Mn) in solutions.



Similarly, careful thought is given to *physical properties* of a substance, such as melting and boiling points, colour and solubility. Many metals, for example, lend themselves to AAS due to their ability to conduct electricity. If chromatographic instruments are being considered, boiling temperature of a substance is important in distinguishing between gas chromatography and high-performance liquid chromatography.

The type of analysis used will be determined by the properties of the substance to be analysed. For example:
- **AAS can be used for metal ions.**
- **Liquid chromatography is used to analyse many organic molecules.**

Another important consideration for the analyst is the *degree of accuracy* required. This could mean the difference between choosing a tedious but accurate method and choosing a faster but less accurate one. Is a qualitative determination all that is needed, or is a quantitative one called for? An appreciation of the strengths and weaknesses of common analytical techniques is therefore necessary, so that the most appropriate method is chosen.

Standard tests to identify substances are also important. In the analysis of an unknown compound, for example, it would be important to establish that an evolved gas is carbon dioxide rather than oxygen or hydrogen.

All these considerations are important in designing a method of analysis. If no particular standard method suits, the chemical analyst may have to display initiative and creativity in designing a new, and maybe innovative, method to get the job done.

17.2 EXERCISE

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1. Classify the following analyses as either qualitative or quantitative.
   (a) A pregnancy test in which a chemical is added to a sample of urine and a colour change is sought
   (b) A chlorine test in which the colour of a chemical is compared with reference standards to estimate the chlorine level in a home swimming pool
   (c) Placement of a detector in the exhaust pipe of a car during a tune-up procedure to measure the level of carbon monoxide emissions
   (d) Testing for monosodium glutamate, MSG, in a sample of food claimed to be 'MSG free'
   (e) Using a pH paper indicator to measure acidity of a sample
   (f) Observing that a precipitation reaction occurs
   (g) Adding phenolphthalein indicator in a sample
   (h) Using electrical conductivity to test for salts
2. Oil tankers need to clean residual oil from their tanks before they load new cargo. The washings are supposed to be stored on board but unscrupulous captains sometimes dump these wastes at sea to reduce costs.
   In Victoria, there have been numerous cases of such wastes washing ashore and polluting beaches. Sometimes this has occurred in the vicinity of Phillip Island and has affected its colony of fairy penguins. In some instances, the analysis of such material has led to the prosecution of offending tanker captains.
   (a) Describe how qualitative analysis might assist in such cases.
   (b) Given that such oil is usually a complicated mixture, describe how quantitative analysis could lead to the identification of the ship causing such pollution.

Fully worked solutions and sample responses are available in your digital formats.

17.3 The sources of salts in water

KEY CONCEPT

- Sources of salts found in water (may include minerals, heavy metals, organo-metallic substances) and the use of electrical conductivity to determine the salinity of water samples

As discussed in topic 15, contaminants are unwanted substances found in a body of water. They may include minerals, heavy metals, organo-metallic substances and microorganisms (microorganisms are not included in the study design).

17.3.1 Minerals

The salts found naturally in water come from a wide variety of sources. Many minerals in the ground dissolve as the water flows over the ground or percolates through layers of rock underneath it. Sodium chloride is the most common salt found in water, and the fact that the oceans contain so much sodium chloride is evidence of this.

It should be remembered, however, that salts do not need to dissolve in large amounts to have a significant effect on how water may be used. Consider the salt calcium carbonate, $CaCO_3$, which dissolves from rocks such as limestone and chalk, as water flows over or through them. Calcium carbonate has a very low solubility (0.013 g L^{-1} at SLC). However, when it reacts with rain water, which contains dissolved carbon dioxide, calcium hydrogen carbonate (also known as calcium bicarbonate) is produced, which is much more soluble (~166 g L^{-1}). The equation for this reaction is:

$$CaCO_3(s) + CO_2(aq) + H_2O(l) \rightarrow Ca(HCO_3)_2(aq)$$

Dissolved calcium ions, Ca^{2+}, along with magnesium ions, Mg^{2+}, are a major cause of hardness in water, which makes it difficult to lather and also makes it unsuitable for use in boilers and other high-temperature applications.

Human activity can also affect the levels of dissolved salts. A good example of this is the increasing salinity levels in the Murray River. Until 100 million years ago, the Murray–Darling Basin was covered by a shallow sea. After this sea retreated, the salt deposits left behind were buried underground. Since the advent of extensive irrigation, the extra water soaking into the ground has caused water tables to rise. This has brought large amounts of this previously undisturbed salt to, or near to, the surface — from where it can enter rivers and streams, which eventually flow into the Murray River. The response to this problem has been the establishment of a number of salinity management schemes. Figure 17.4 shows measured salinity levels at Morgan in South Australia over 40 years, compared to modelled levels without a salinity strategy.

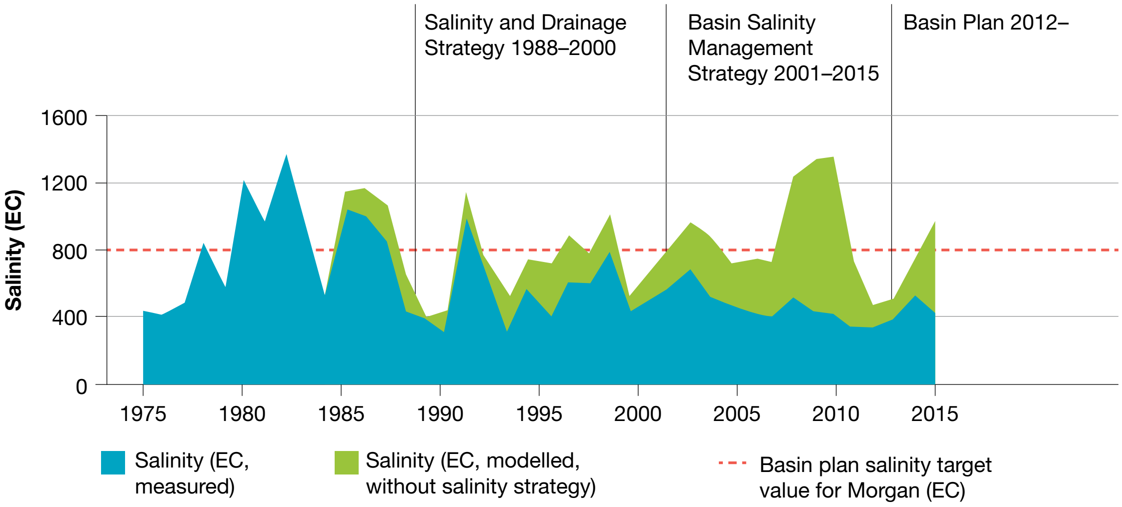
FIGURE 17.3 The label on a bottle of mineral water shows the many substances that are dissolved in it.

TYPICAL ANALYSIS mg/l

CALCIUM	55
MAGNESIUM	19
POTASSIUM	1
SODIUM	24
BICARBONATE	248
CHLORIDE	37
SULPHATE	13
NITRATE	<0.1
IRON	0
ALUMINIUM	0
DRY RESIDUE AT 180°C	280
pH AT SOURCE	**7.4**

BEST BEFORE END: SEE BOTTLE

FIGURE 17.4 Measured salinity levels at Morgan in South Australia against modelled salinity levels without a salinity strategy



Dissolved salts in water can come from natural sources as well as from human activity.

17.3.2 Heavy metals

Mercury, lead and cadmium are just some of the heavy metals that pose a major disposal problem. Inorganic and organic lead compounds are extremely poisonous and are difficult to dispose of. As detailed in subtopic 15.3, heavy metal contamination is commonly caused from industrial and agricultural activities due to corrosion of materials and inadequate disposal.

The dangers of some heavy metals were first exposed after a tragedy in Minemata Bay in Japan during the 1950s, where at least 60 people had been seriously affected by mercury poisoning prior to the cause being identified. Estimates of the cumulative death toll vary but it is generally agreed to be many thousands, and birth defects persist to this day. The source of the mercury was a chemical company that had been using mercury as a catalyst in the manufacture of acetaldehyde, a precursor to PVA (polyvinyl acetate), and was discharging large quantities of mercury waste into the sea. Bacteria converted the waste to toxic monomethyl mercury, often called methyl mercury, which was ingested by the fish, and subsequently by the humans in a process known as biomagnification (see section 15.3.3).

17.3.3 Organo-metallic substances

The conversion of mercury into toxic methyl mercury is an example of the formation of an organo-metallic substance. These substances form when a heavy metal forms a compound with an organic molecule. While inorganic heavy metals are often only sparingly soluble, organo-metallic heavy metal compounds are much more soluble and can, therefore, enter the food chain more easily, and hence be more environmentally dangerous. Methyl mercury is one such example. While dimethyl mercury is virtually insoluble (and a very dangerous poison), monomethyl mercury is essentially ionic and dissolves much more readily in water. Its structure is:

$$H_3C{-}Hg^+X^-$$

where X is an appropriate anion. You will note from this structure that a covalent bond exists between the mercury atom and the methyl group. This is an example of a more complex type of bonding that does not fit the models that were considered in Unit 1.

What makes this compound particularly dangerous is that it can also be produced when microorganisms in water combine with organic residues that may be present in the water.

17.3.4 Measuring salinity using electrical conductivity

Electrical conductivity provides a quick and inexpensive way of measuring the combined level of dissolved salts. Measurements are usually made in electrical conductivity units (ECs). Figure 17.4, outlining salinity levels at Morgan, shows salinity levels in EC units.

Salts are ionic compounds; when they dissolve in water, they dissociate to produce mobile ions. This means that they can conduct electricity to an extent that depends on how many ions are present. Electrical conductivity is measured using a handheld device (figure 17.5) and is, therefore, a quick and cheap method that can be used where the *total* salt content of a water sample is to be measured; it does not determine which particular salts are present.

FIGURE 17.5 An electrical conductivity meter

17.3 EXERCISE

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1. What is salt?
2. What is the most common salt found in water?
3. Give five example of elements that can be found dissolved in mineral water.
4. Calcium in mineral water is sourced from $CaCO_3$ from limestone. Given that calcium carbonate is almost insoluble in water, explain why calcium is found in mineral water.
5. Explain the sources of heavy metals in water.
6. What are the dangers posed by some heavy metals, such as mercury and lead, in drinking water?
7. What is an organo-metallic substance? Provide an example.
8. Explain why electrical conductivity provides a measurement of overall salt content rather than accurate values for particular salts.
9. What are the units for electrical conductivity?
10. Which solution will measure a higher electrical conductivity: 1M of $NaNO_3$ or 1M of Na_2CO_3? Explain.

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17.4 Analytical methods

KEY CONCEPTS

- The application of mass–mass stoichiometry to gravimetric analysis to determine the mass of a salt in a water sample
- The application of colorimetry and/or UV–visible spectroscopy, including the use of a calibration curve, to determine the concentration of coloured species (ions or complexes) in a water sample
- The application of atomic absorption spectroscopy (AAS), including the use a calibration curve, to determine the concentration of metals or metal ions in a water sample (excluding details of instrument)

17.4.1 Stoichiometry

The term used for calculations based on chemical equations is **stoichiometry**. Stoichiometry is the study and prediction of relative amounts (or ratios) of substances involved in chemical reactions.

The word ‘stoichiometry’ comes from two Greek words meaning ‘element’ and ‘measure’. Knowledge of stoichiometry is essential whenever quantitative information about a chemical reaction is required.

How is chemistry involved in the operation of an automotive airbag? Before being inflated, airbags contain a small amount of solid sodium azide, NaN_3. On impact with another vehicle or object, an electric spark is released, acting as a trigger. Very quickly, the sodium azide in the airbag decomposes to form solid sodium, Na, and nitrogen gas, N_2, according to the chemical equation:

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$$

The sodium then reacts with water vapour in air. The nitrogen gas produced in the reaction inflates the airbag in time to absorb much of the impact energy of the crash, which may otherwise have caused injury to the driver. The sodium azide reaction must produce just the right amount of nitrogen gas to inflate the bag to the correct pressure. How do designers know how much sodium azide to use so that the airbag is correctly inflated? They use stoichiometric calculations.

Solving stoichiometric problems is particularly important in fields such as food chemistry, drug chemistry, forensic science and in any industry where chemicals are manufactured. Stoichiometry may be thought of as the means of obtaining a recipe for a chemical reaction.

Mole ratios

Chemical equations show the relative number of molecules participating in a chemical reaction. They can also show the relative number of moles, the **mole ratio**, of each substance involved.

Chemically speaking, if we wish to say, 'One molecule of nitrogen gas reacts with three molecules of hydrogen gas to form two molecules of ammonia gas', we simply write:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

However, it would be impossible and impractical to make ammonia molecule by molecule. Chemists work with larger quantities called moles (symbol mol). The coefficients in a balanced chemical equation tell us the relative numbers of moles of reactants and products taking part in the reaction.

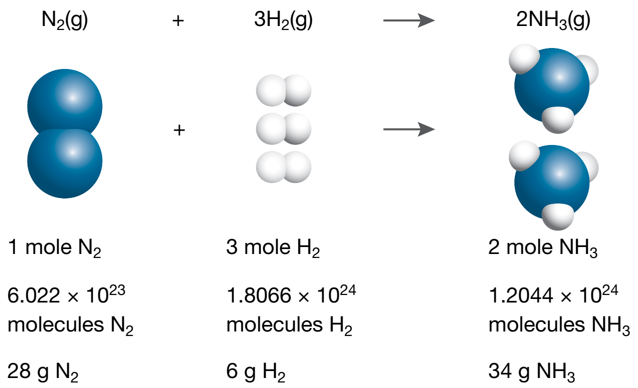
For example, one mole of nitrogen gas reacts with three moles of hydrogen gas to produce two moles of ammonia gas.

The equation actually indicates the mole ratio of reactants and products and the ratios never change. If one mole of nitrogen reacts with three moles of hydrogen, then two moles of nitrogen reacts with six moles of hydrogen, 0.5 mole of nitrogen reacts with 1.5 moles of hydrogen, and so on.

Assuming that plenty of nitrogen is available, how much ammonia is produced by three moles of hydrogen? The equation tells us: two moles of ammonia. Similarly, six moles of hydrogen produces four moles of ammonia.

An equation also tells us about the ratio between other quantities of the reactants and the products including the number of particles and the masses of the reactants and products (figure 17.6).

FIGURE 17.6 The balanced chemical equation for the formation of ammonia provides information about the ratios and abundances of products and reactants.



What a chemical equation *does not* tell us

An equation conveys no information about the rate of a reaction. It may be fast, as in the explosive oxidation of hydrogen to form water, or it may be slow, as in the oxidation or rusting of iron.

An equation does not tell us whether a reaction requires heat or gives off heat, or what temperature or pressure is needed.

Finally, an equation gives no details as to how the individual atoms or molecules are transformed from reactants to products. This information is essential for full understanding of how a chemical reaction takes place.

SAMPLE PROBLEM 1

Methanol, CH_3OH, is used as a fuel in some racing cars. The equation for the combustion of methanol in the car's engine is:

$$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(g)$$

How many moles of oxygen are needed to react with 4 moles of methanol?

🎞 **Teacher-led video:** SP1 (tlvd-0596)

THINK	WRITE
From the coefficients in the reaction we can see that two moles of methanol will react with three moles of oxygen.	The mole ratio of methanol:oxygen is 2:3. Therefore, four moles of methanol will require six moles of oxygen.

PRACTICE PROBLEM 1

Using the reaction from sample problem 1, $2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(g)$, how many moles of water are formed when one mole of oxygen reacts completely with methanol?

17.4.2 Mass–mass stoichiometry

Mass–mass stoichiometry involves solving a problem in which the mass of a reactant or product is given. You are then asked to calculate the mass of another reactant or product. Mass–mass stoichiometry requires the conversion of masses of substances to moles, or moles of substances to masses, through the application of the formula for determining the number of moles of a substance.

Mass–mass stoichiometry formulas:

$$\text{moles}\,(n) = \frac{\text{mass}\,(m)}{\text{molar mass}\,(M)}$$

$$\text{mass}\,(m) = \text{moles}\,(n) \times \text{molar mass}\,(M)$$

TIP: **The formula $n = \frac{m}{M}$ can be found in table 3 of the VCE Chemistry Data Book.**

Solving mass–mass stoichiometric problems in four steps

1. Write a balanced chemical equation for the reaction, identifying the known (given) and unknown (required) quantities of substance.
2. Calculate the number of moles of the known quantity of substance present.
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.
4. Calculate the quantity (mass) of the required substance.

SAMPLE PROBLEM 2

Some sulfur is present in coal in the form of pyrite, FeS_2. This substance is also known as 'fool's gold'. When pyrite burns, it pollutes the air with one of its combustion products, sulfur dioxide, and produces solid iron(III) oxide. What mass of iron(III) oxide is formed from the complete combustion of 183.5 g of pyrite?

Teacher-led video: SP2 (tlvd-0597)

THINK	WRITE
1. Write the balanced equation and identify known (given) and unknown (required) moles.	$4FeS_2(s) + 11O_2(g) \rightarrow 2Fe_2O_3(s) + 8SO_2(g)$ Pyrite is the known quantity and iron(III) oxide is the unknown quantity.
2. Calculate the number of moles of the known quantity of substance, FeS_2, using the formula $n = \frac{m}{M}$.	$n(FeS_2) = \frac{183.5}{55.8 + (2 \times 32.1)}$ $= \frac{183.5}{120.0}$ $= 1.529$ mol
3. Find the molar ratio, FeS_2:Fe_2O_3, from the equation, and use it to calculate the number of moles of the required substance, Fe_2O_3.	The ratio FeS_2:Fe_2O_3 is 4:2, which becomes 2:1. $n(Fe_2O_3) = \frac{2}{4} \times n(FeS_2) = \frac{n(FeS_2)}{2}$ $= \frac{1.529}{2}$ $= 0.7645$ mol
4. Calculate the required quantity (mass) of Fe_2O_3 formed by the reaction.	$m(Fe_2O_3) = n \times M$ $= 0.7645 \times (2 \times 55.8 + 3 \times 16.0)$ $= 122.0$ g

PRACTICE PROBLEM 2

When 2.864 g of potassium iodide reacts completely with a solution of lead nitrate, a yellow precipitate of lead iodide forms according to the following equation:

$$2KI(s) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$$

Calculate the mass of lead iodide that precipitates.

Calculations with limiting or excess reactants

Sometimes the amount of one of the reactants is not available in the required mole ratio (as per reaction). This means that the second reactant will not be used up completely in the reaction, because its consumption will be limited by the small amount of the first reactant.

The amounts of the products are limited by the amount of the reactant that is completely used up in the reaction. Let's return to the previous example in section 17.4.1:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Recall that a balanced equation indicates the mole ratios in which the reactants are used up and the products are formed. According to this equation, one mole of nitrogen gas reacts with three moles of hydrogen gas to produce two moles of ammonia gas.

If one mole of nitrogen and four moles of hydrogen are mixed and allowed to react, all the nitrogen and three moles of hydrogen are used up. This means that one mole of hydrogen is left over. The reactant that is completely used up (in this case nitrogen) is called the **limiting reactant**. Any unreacted reactants (hydrogen in this example) are called **excess reactants**. The amount of product formed by the reaction is limited by the amount of the limiting reactant. In this case, the addition of more hydrogen to the mixture would have no effect on the amount of product formed, because no more nitrogen is available to react with it.

In a limiting reactant calculation, first determine which substance is completely used up (the limiting reactant) and which is left over (the excess reactant). Calculations are then based on the moles of the limiting reactant.

Note: the same steps are used in questions asking to identify the excess reactant. However, any further calculations will still be using the limiting reactant.

SAMPLE PROBLEM 3

Magnesium and oxygen react to form magnesium oxide according to the equation:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

If 24.31 g of magnesium and 20.00 g of oxygen are available for reaction, determine the limiting reactant, the reactant in excess, and the mass of magnesium oxide formed.

Teacher-led video: SP3 (tlvd-0598)

THINK	WRITE
1. Convert the mass to moles.	$n(\text{Mg}) = \frac{m}{M}$ $= \frac{24.31}{24.3}$ $= 1.00\ \text{mol}$ $n(\text{O}_2) = \frac{m}{M}$ $= \frac{20.00}{32.0}$ $= 0.625\ \text{mol}$

2. Compare this mole ratio with the mole ratio in the equation to identify which reactant is limiting.	According to the equation, the ratio of Mg:O is 2:1. Therefore, if 1.00 mol of Mg were to react we would need 0.500 mol of O_2. We have 0.625 mol O_2, which is more than enough. Therefore, Mg is the limiting reactant and O_2 is the excess reactant.
3. Calculate the mass of MgO produced.	$m = n \times M$ (MgO) $= 1.00 \times 40.3$ $= 40.3$ g

---

PRACTICE PROBLEM 3

Aluminium burns in oxygen to produce an intense burst of light. Aluminium oxide is produced in the process. The equation for the reaction is:

$$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$

Calculate the mass of aluminium oxide produced from 5.00 g of aluminium and 10.0 g of oxygen.

17.4.3 Mass–concentration stoichiometry

Many chemical reactions involve interactions between solids and solutions. These reactions may be referred to as mass–concentration stoichiometry. Two solutions may be mixed to form a precipitate, or solids may dissolve in some solutions to form new products. Stoichiometric calculations that involve solids and solutions require the use of two formulas to calculate moles.

When working with solids, we use the formula:

$$n = \frac{m}{M}$$

When working with solutions, we use the formula:

$$n = c \times V$$

Solving mass–concentration stoichiometric problems in four steps

1. Write a balanced chemical equation for the reaction, identifying the known (given) and unknown (required) quantities of substance.
2. Calculate the number of moles of the known quantity of substance present.
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.
4. Calculate the quantity of the required substance.

Note: mass–concentration stoichiometry or mass–volume stoichiometry involves the same steps as in mass–mass stoichiometry, except that the formula used is $n = c \times V$.

SAMPLE PROBLEM 4

Art may be created on glass using hydroflouric acid, in a process called etching. The reaction that occurs is:

$$SiO_2(s) + 4HF(aq) \rightarrow SiF_4(g) + 2H_2O(l)$$

A quartz sculptor has a 500 mL container of 22.50 M hydrofluoric acid. What mass of quartz could be etched by the acid?

Teacher-led video: SP4 (tlvd-0599)

THINK	**WRITE**
1. Write the equation, identifying the known and unknown quantities.	$SiO_2(s) + 4HF(aq) \rightarrow SiF_4 + 2H_2O(1)$ *unknown* *known*
2. Calculate the number of moles of the known quantity of substance; HF.	$n\,(HF) = cV$ $= 22.50 \times 0.500$ $= 11.3$ mol
3. Find the molar ratio from the equation and use it to calculate the number of moles of SiO_2 required.	$n(SiO_2) : n(HF)$ is 1 : 4 $n(SiO_2) = \frac{1}{4}n(HF)$ $= 2.81$ mol
4. Calculate the required quantity of SiO_2, using $m = n \times M$.	$m = n \times M$ $= 2.81 \times (28.1 + 2 \times 16.0)$ $= 169$ g

PRACTICE PROBLEM 4

A piece of aluminium is placed in a beaker containing 500 mL of H_2SO_4 solution and hydrogen gas is evolved. Given that the initial mass of Al was 15.14 g and its final mass was 9.74 g, calculate the concentration of the acid.

17.4.4 Gravimetric analysis

Gravimetric analysis is a form of analysis by mass. It is a well-established technique that has been carried out in analytical laboratories for centuries.

At its simplest, it may involve drying a product to determine its moisture content. For example, we read about water being added to frozen chickens in order to increase the apparent weight at which they are sold. Gravimetric analysis can be used quickly to check whether this practice has occurred. In the area of agriculture, a soil scientist often needs to know the moisture content of a soil sample in order to determine the effectiveness of various soil conditioning procedures.

Gravimetric analysis can also be a much more sophisticated procedure in which knowledge of chemical reactions, solubilities and stoichiometry is used to determine the amount of a substance in a sample. Using chemical knowledge in this way makes gravimetric analysis a powerful tool that can be used by analytical chemists in a wide range of situations.

on Resources

Video eLesson Gravimetric analysis (eles-2491)

SAMPLE PROBLEM 5

A 23.1 g sample of soil is heated to 105 °C for 60 minutes. It is then weighed and the process repeated a number of times. A mass of 21.0 g is eventually obtained, which does not change upon further heating.

Calculate the percentage of water in this sample of soil.

THINK	WRITE
1. Identify the change in water mass.	The mass of water lost from the sample is $23.1 - 21.0 = 2.1$ g.
2. Calculate the percentage of water in the sample.	$\frac{2.1}{23.1} \times \frac{100}{1} = 9.1\%$

PRACTICE PROBLEM 5

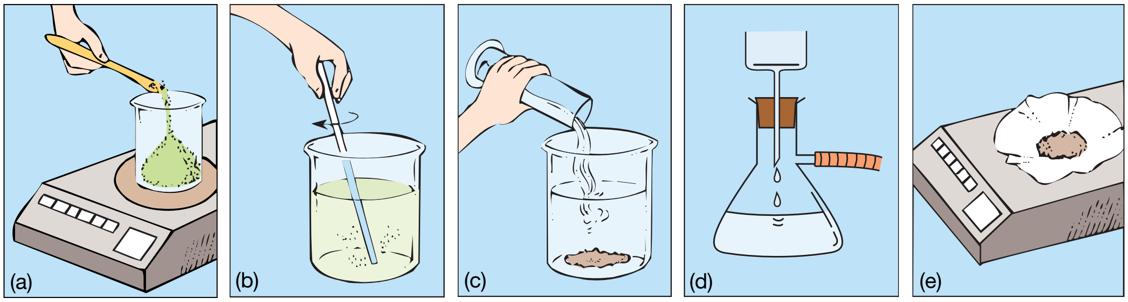
The soil sample from sample problem 5 was kept in inappropriate conditions and it re-absorbed moisture from air. A week later, its new mass was found to be 22.1 g. Calculate the percentage of water re-absorbed by the dried soil sample.

Extending the gravimetric method

A more sophisticated version of gravimetric analysis can be used to quantitatively determine a component in a solid mixture.

In this procedure, the sample being analysed is usually dissolved in water and then treated with a chemical that forms a precipitate. (It is this precipitate that removes the ions required from solution so that the amount of ions may later be determined.) This precipitate is then carefully collected (usually by filtration), dried and weighed to constant mass as described earlier. From this mass and the original mass of material dissolved, the component may be determined quantitatively.

FIGURE 17.7 The steps involved in gravimetric analysis



Gravimetric analysis often involves the following steps:

a. weighing the sample to be analysed
b. dissolving the sample in water
c. adding a suitable chemical to form a precipitate
d. filtering to collect the precipitate
e. repeated drying and weighing until a constant mass of precipitate is obtained

Simple stoichiometry is then used to deduce the amount of the component in the original mixture.

Note: sometimes the substance being analysed is already dissolved, as with water samples. In such cases, the method starts with adding a suitable chemical to form a precipitate.

This procedure is especially suitable for determining metal ions. It is also very useful for determining anions such as sulfate, which form some easily prepared insoluble salts (such as barium sulfate).

When designing a gravimetric procedure, knowing the solubilities of the precipitates likely to be produced is important. The solubility in water of some common compounds is shown in table 17.1.

TABLE 17.1 The solubility in water of compounds of common anions

Anion	Cations forming soluble compounds	Cations forming insoluble compounds
Nitrates	All	—
Chlorides Bromides Iodides	Most	Ag^+, Pb^{2+}($PbCl_2$ is moderately soluble in hot water.)
Sulfates	Most	Ba^{2+}, Pb^{2+}(Ag_2SO_4 and $CaSO_4$ are slightly soluble.)
Carbonates	Na^+, K^+, $NH_4{}^+$	Most
Phosphates	Na^+, K^+, $NH_4{}^+$	Most
Sulfides	Na^+, K^+, $NH_4{}^+$	Most (MgS, CaS, BaS, Al_2S_3 and Fe_2S_3 decompose in water.)
Hydroxides and oxides	Na^+, K^+, Ba^{2+}	Most ($Ca(OH)_2$ is slightly soluble.)

on Resources

Interactivity Gravimetric analysis (int-6357)

SAMPLE PROBLEM 6

A pharmaceutical company wishes to test the purity of some commercial barium chloride.

A sample of the commercial barium chloride was weighed and found to have a mass of 10.0 g. After dissolving in water, excess sulfuric acid was added to form a precipitate of barium sulfate. This precipitate was then filtered, dried and weighed. It was found to have a mass of 10.55 g.

Calculate the percentage purity of the barium chloride.

Teacher-led video: SP6 (tlvd-0601)

THINK	WRITE
1. Write the balanced equation	$BaCl_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(s) + 2HCl(aq)$

2. Calculate the moles of the known quantity of product.	$m(\text{BaSO}_4) = 10.55\ \text{g}$ $M(\text{BaSO}_4) = 233.4\ \text{g mol}^{-1}$ $\therefore n(\text{BaSO}_4) = \frac{m}{M}$ $\frac{10.\ 55}{233.4}$ $= 0.045\ 20\ \text{mol}$
3. Use the mole ratio from the equation to calculate the moles of BaCl_2 used in reaction	$n(\text{BaCl}_2) : n(\text{BaSO}_4) = 1{:}1$ $\therefore n(\text{BaSO}_4) = 0.04\ 520\ \text{mol}$
4. Calculate mass of BaCl_2.	$m = n \times M$ $M(\text{BaCl}_2) = 208.3\ \text{g mol}^{-1}$ $n(\text{BaCl}_2) = 0.045\ 20\ \text{mol}$ $\therefore m(\text{BaCl}_2) = 0.045\ 20 \times 208.3\ \text{g}$ $= 9.415\ \text{g}$
5. Calculate the percentage purity of the barium chloride.	$\therefore$ % barium chloride (original sample) $= \frac{9.415}{10.0} \times 100$ $= 94.2\%$ (3 significant figures)

PRACTICE PROBLEM 6

To determine the percentage of arsenic in a particular brand of pesticide, a 2.15 g sample is treated to precipitate all the arsenic as its sulfide, As_2S_3. If 0.353 g of precipitate is obtained, calculate the percentage of arsenic in the pesticide.

on Resources

Interactivity Common precipitation reactions (int-6358)

SAMPLE PROBLEM 7

A sample of groundwater from an abandoned lead battery manufacturing site was analysed gravimetrically for its lead content.

A 2000 mL sample was treated with sodium sulfate solution until no further precipitate was observed to form. After filtering and drying to constant weight, 1.21 g of lead sulfate was obtained.

Calculate the level of lead ions in the groundwater in mg L^{-1}.

Teacher-led video: SP7 (tlvd-0602)

THINK	WRITE
1. As we are interested in the Pb^{2+} ions only, an ionic equation is more practical than a full equation. We can therefore write the equation and list the known information.	$\text{Pb}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{PbSO}_4(\text{s})$ $m(\text{PbSO}_4) = 1.21\ \text{g}$ $M(\text{PbSO}_4) = 207.2 + 32.1 + 4 \times 16.0$ $= 303.3\ \text{g mol}^{-1}$

2. Calculate moles of $PbSO_4$.	$n(PbSO_4) = \frac{1.21}{303.3}$ $= 0.003\,99\text{ mol}$
3. Use the mole ratio from the equation to calculate moles of Pb^{2+}.	$n(Pb^{2+}):n(PbSO_4) = 1:1$ Therefore, $n(Pb^{2+}) = 0.003\,99\text{ mol}$
4. Calculate mass of Pb^{2+}. **TIP:** Remember to use the non-rounded value for the number of moles.	$m(Pb^{2+}) = 0.003\,99 \times 207.2$ $= 0.827\text{ g}$
5. Calculate Pb^{2+} in mg L^{-1}.	Since the amount analysed was 2000 mL (2 L), the concentration of Pb^{2+} equals: $\frac{0.827}{2} = 0.413\text{ g L}^{-1}$ $= 413\text{ mg L}^{-1}$

Alternative solution:

The workings shown rely on the assumption that all the lead in the original sample precipitates. Of course, this should be true of *any* properly designed gravimetric process!

When the molar mass of $PbSO_4$ is evaluated, it becomes clear that, in the 303.3 g that represents one mole, 207.2 g is due to lead.

The fraction of lead in $PbSO_4$ is therefore $\frac{207.2}{303.3}$.

It follows that in 1.214 g of $PbSO_4$ there are $\frac{207.2}{303.3} \times 1.21 = 0.827$ g of lead.

The rest of the calculation is the same as shown.

PRACTICE PROBLEM 7

The zinc content in a water sample was analysed as follows.

A 1000 mL sample was treated with a solution of ammonium hydrogen phosphate, $(NH_4)_2HPO_4$, to precipitate all the zinc ions present. A precipitate of NH_4ZnPO_4 was obtained, which was then decomposed by heating to produce 2.918 g of $Zn_2P_2O_7$.

Calculate the concentration of zinc in the sample in g L^{-1}.

Sources of error in gravimetric analysis

Although most gravimetric procedures have a clearly recognisable series of steps, the properties of the chemicals involved may necessitate some modifications to the method in particular circumstances.

A thorough appreciation of the method is also important in evaluating the limits of your results. If certain things happen during the analysis, it is important to predict how these may affect the final result. Table 17.2 shows some possible sources of error.

TABLE 17.2 Sources of error in gravimetric analysis

Action	Effect on analysis result	Reason
Insoluble materials not filtered out before forming the precipitate	Overestimated	The apparent mass of the precipitate will increase.
Not enough of the precipitate-formingchemical added	Underestimated	Not enough of the precipitate will form because some of the ions that are being analysed will remain in solution.
Forming a precipitate that is too soluble	Underestimated	Not all of the ions being analysed will be in the precipitate.
Forming extra precipitate due to the presence of other competing ions	Overestimated	Too much of the precipitate will form.
Weighing the precipitate before it is dry	Overestimated	The water present will increase the apparent mass of the precipitate.
Not rinsing the precipitate before drying it	Overestimated	As the precipitate dries, other soluble chemicals will begin to crystallise out of the small amount of solution still trapped in the precipitate. These will add to the mass.*
Adding too much of the precipitate-causing chemical	No effect	This is a necessary part of the method to make sure that all of the required ions are in the precipitate. The chemical must be in excess.
Using too much water for the initial dissolving	No effect	This is a practical consideration — the more water you have, the longer the filtering step.

**Note*: to avoid this situation, the precipitate should always be washed with a small amount of pure solvent before it is dried and weighed. Too much water or solvent could cause some of the precipitate to redissolve. In this case, the final result will be underestimated.

17.4.5 The application of colorimetry and UV–visibility in water analysis

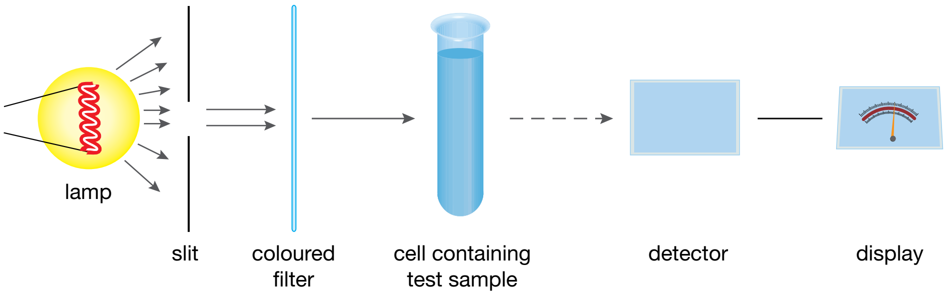
Colorimetry is a technique in which colour levels are compared with a set of standards; where the intensity of the colour depends on the component being analysed.

If a solution is naturally coloured, or can be coloured by adding certain chemicals, colorimetric analysis may be performed. Two common examples of this are the use of universal indicator to measure pH, where the colour produced is compared by eye with a standard chart, and estimating the level of chlorine in swimming pool water, where the sample is coloured by adding chemicals and the results compared with a chart.

More accurate results can be obtained by using an instrument called an **instrumental colorimeter**. This compares the colour in the test sample with the colours produced in samples of known concentration (standards) that have been treated identically to the test sample. From the readings produced, a graph of absorbance versus concentration (called a calibration curve) may be produced and the concentration in the sample may be read directly from this graph.

The basic design of a colorimeter is shown in figure 17.8.

FIGURE 17.8 The essential features of the colorimeter



In a colorimeter, the detector measures the amount of light that passes through the test cell and converts this into a reading that is displayed. While normal white light can be used, more accurate results are obtained if light that is *complementary* to the colour in the test cell is used. When light of a particular colour is removed from white light, the remaining light produces a different colour. The second colour is said to be complementary to the original colour. For example, if copper sulfate (blue–green) levels are being measured, the best colour to shine through such samples is red light, as shown in figure 17.9. Examples of complementary colours are summarised in table 17.3.

FIGURE 17.9 A blue–green substance strongly absorbs red light.

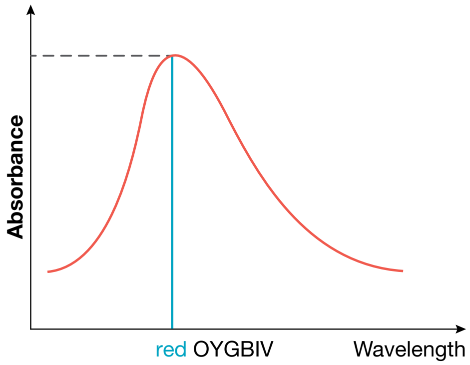


TABLE 17.3 Observed colours and their complementary colours

Colour observed	Colour absorbed	Frequency of absorbed colour ($\times 10^{14}$ Hz)
Green	Deep red	3.85–4.41
Blue–green	Red	4.41–4.83
Blue	Orange	4.83–5.17
Blue–violet	Yellow	5.17–5.45
Violet	Yellow–green	5.45–5.76
Deep red	Green	5.76–6.00
Red	Blue–green	6.00–6.38
Orange	Blue	6.38–6.81
Yellow	Blue–violet	6.81–7.14
Yellow–green	Violet	7.14–7.89

Colorimetry is relatively cheap and is obviously useful for measuring the concentration of a coloured species (or a colourless species from which a coloured derivative can be easily prepared). However, it is very important that care is taken to ensure that no other coloured species are present in a sample that could also absorb the particular colour of light being used. This method can produce results to within 1 to 2% accuracy.

UV–visible spectrometry and atomic absorption spectrometry are two further methods that can be used if more accurate results are required.

Example of colorimetry use for determining phosphate levels

Phosphates are important nutrients for plants but can become an environmental problem when they enter water systems in large amounts. As we have already seen, this can lead to eutrophication if certain other conditions are present as well. Sources of phosphate include fertilisers such as ammonium phosphates (typically $(NH_4)H_2PO_4$ and/or $(NH_4)_2HPO_4$) and superphosphate (a mixture of calcium dihydrogen phosphate, $Ca(H_2PO_4)_2$, and calcium sulfate, $CaSO_4$), which are applied in large amounts as part of current agricultural practices. Until recently, laundry detergents were another source of this nutrient; they were added to remove hardness and buffer the wash water to maintain a slightly alkaline pH. Environmental awareness by consumers, however, has now led to phosphate removal from virtually all such detergents.

Testing for phosphate levels is a relatively easy application of colorimetry. The process involves the addition of ammonium molybdate, a chemical that forms a blue compound if phosphate is present — and the more phosphate, the more intense the blue colour of the solution. If a set of standards containing known phosphate levels is produced in the same way as the unknown, the level in the unknown sample may then be determined. This can be done roughly by eye using colour matching, or by determining a calibration curve produced from an instrumental colorimeter.

17.4.6 Calibration curves

Creating a calibration curve is an important step in analytical chemistry to assist with the calibration of the instrument used. A calibration curve is used in many analyses, including UV–visible spectroscopy and atomic absorption spectroscopy (AAS).

This step involves analysis of several known concentrations (also known as **standard solutions**) of the substance, and then plotting a graph with the results and determining the line of best-fit. The calibration curve is then used to determine unknown concentrations of the substance from its calibration curve.

SAMPLE PROBLEM 8

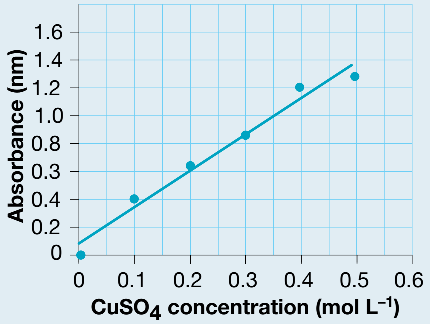
The absorbance of several solutions of $CuSO_4$ was determined through colorimetry and the results are shown in the table. Using a calibration curve, determine the absorbance of a solution of 0.23 mol L^{-1} $CuSO_4$.

Concentration $CuSO_4$ (M)	Absorbance (nm)
0	0
0.1	0.406
0.2	0.638
0.3	0.854
0.4	1.202
0.5	1.276

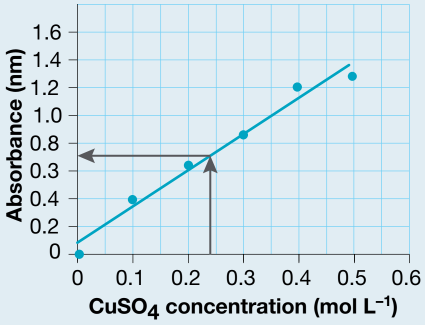
THINK

1. Plot the graph of absorbance versus concentration and the line of best fit.
The line of best fit represents the calibration curve for this experiment.

WRITE



2. Draw a vertical line from the *x*-axis to the calibration line and then from the line to the *y*-axis to determine the corresponding absorbance for a solution of 0.23 mol L^{-1} $CuSO_4$.



The absorbance of the 0.23 mol L^{-1} solution is 0.74 nm.

PRACTICE PROBLEM 8

With the data from sample problem 8, determine the absorbance for a sample of 0.05 mol L^{-1} $CuSO_4$.

17.4.7 The application of UV–visible spectroscopy in coloured water samples

UV–visible spectroscopy measures the absorption of light at a particular frequency. It is suitable for coloured solutions, solutions that can be made coloured, and solutions that absorb in the ultraviolet region.

In many ways, **UV–visible spectroscopy** is a more sophisticated development of instrumental colorimetry. The basic idea is the same — the amount of absorption is related to the concentration of the substance being tested. However, this technique is far more selective and, therefore, less likely to suffer interference from similarly coloured compounds. This is because light of a specific frequency is used. For example, although two compounds in a sample to be tested appear blue, one of these might absorb strongly at a frequency that the other does not. This frequency can then be used in the subsequent analysis to distinguish between them.

Just as many substances absorb light from the visible section of the electromagnetic spectrum (and, therefore, appear coloured), substances absorb radiation from the ultraviolet region of the spectrum. Our eyes are not able to detect this radiation, which means substances absorbing in this region may not necessarily appear coloured. UV–visible spectroscopy is, therefore, suitable for many colourless substances as well as coloured ones.

A UV–visible spectrophotometer is schematically illustrated in figure 17.10. Light is passed through a monochromator with a wavelength selection device, such as a prism. This selected wavelength then passes through the sample and is picked up by a photodetector and recorded by computer. The resulting spectrum is given as a graph of absorbance (A) versus wavelength (λ).

FIGURE 17.10 UV–visible spectrophotometer schematic

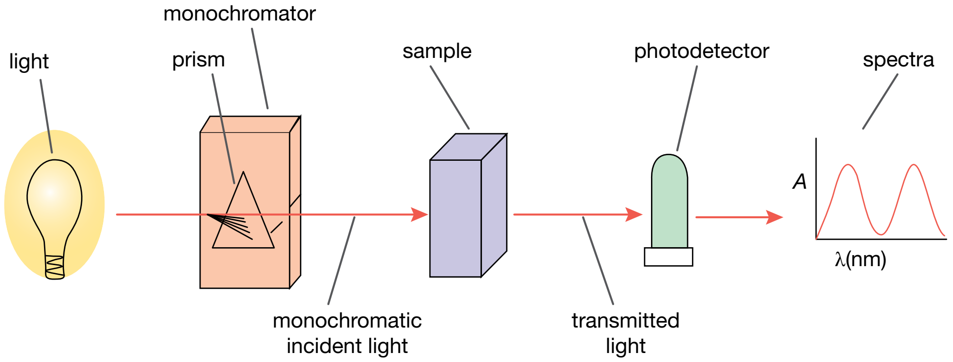


FIGURE 17.11 UV–visible spectrophotometer

The UV–visible spectrophotometer can be used both qualitatively and quantitatively.

For *qualitative analysis*, the sample to be analysed is dissolved using a suitable solvent. A spectrum is obtained by measuring the absorbance against a range of frequencies. This spectrum can then be compared to known spectra for the suspected substances in the sample. In practice, UV–visible spectroscopy is used only as supporting evidence in an identification, not as a qualitative analysis method in its own right.

In *quantitative analysis*, a pure sample of the substance to be measured would initially have its spectrum determined as already described. From this spectrum, a frequency would be chosen at which strong absorption occurs. The absorbance of the test sample at this frequency would then be compared to the absorbance of a set of standards of known concentration at this chosen frequency. For example, to measure the level of glucose in a sample of urine, the spectrum of pure glucose would be obtained and a suitable frequency chosen. Glucose samples of known concentration would then be tested at this chosen frequency and their absorbances noted. Finally, the sample of urine would be tested and the absorbance compared to that of the standards to obtain the glucose concentration. Note that in choosing a suitable frequency for such an analysis, care must be taken to choose a frequency at which the glucose in the urine is the only substance that absorbs.

SAMPLE PROBLEM 9

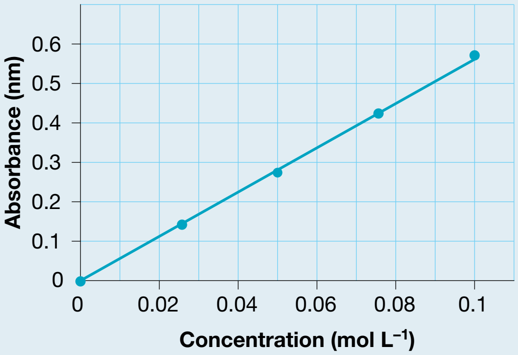
Standard solutions of $Cu(NO_3)_2$ were analysed with UV–visibility and their absorbance was determined, as shown. Derive the concentration of the unknown solution.

Concentration (mol L^{-1})	Absorbance (nm)
0.025	0.135
0.05	0.270
0.075	0.420
0.1	0.570
unknown	0.34

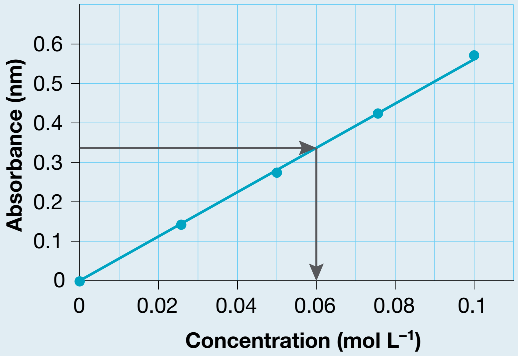
THINK

1. Plot the graph of absorbance versus concentration and the line of best fit.
   The line of best fit represents the calibration curve for this experiment.

WRITE



2. Determine the unknown concentration by drawing a horizontal line from the *y*-axis at the 0.34 nm absorbance value given to the calibration line, and from this to the *x*-axis to determine the corresponding concentration.



The unknown concentration is 0.06 mol L^{-1}.

PRACTICE PROBLEM 9

With data from sample problem 9, determine the absorbance of a sample of 0.04 mol L^{-1} $Cu(NO_3)_2$.

17.4.8 The application of atomic absorption spectroscopy for metals in water samples

Atomic absorption spectroscopy (AAS) is suitable for the detection of many metals and metalloids, and was developed by the CSIRO in Australia. Atoms of a particular element, when energised in a flame, absorb light from an emission lamp containing that same element.

The instrument used is called an atomic absorption spectrometer. It uses the absorption of light to measure concentrations of metal ions. It works on the principle that atoms absorb light if the frequency (and, therefore, the energy) of this light is of the correct value to promote an electron from its ground state energy level to a higher energy level.

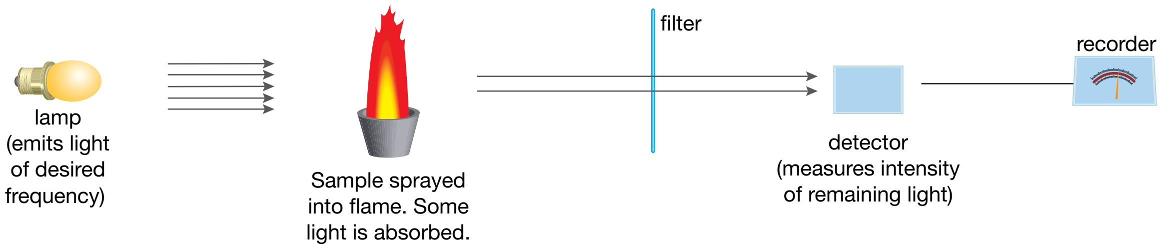
Solutions of known concentration are analysed first, followed by the solution being tested. These solutions are drawn into the flame. A lamp is used to shine light of a very specific wavelength through the flame towards a detector. The amount of light absorbed by the flame is measured and can be used to determine the concentration of the substance being analysed.

FIGURE 17.12 Allan Walsh of the CSIRO — the Australian inventor of the atomic absorption spectrometer

The flame is usually an air/acetylene, C_2H_2, mixture. However, for some analyses a hotter flame is required; in such cases nitrous oxide, N_2O, may be added to the gas mixture.

This technique is both very sensitive as well as being very selective. Concentrations of parts per million are easily measured, and, for some ions, parts per billion. The selectivity of the instrument allows a particular component of a mixture to be analysed without having to separate it from other components. For example, the proportions of sodium and potassium in a salt substitute can be measured without one ion affecting the other. This selectivity is due to the frequency of the light from the lamp. If, for example, a lamp is used that emits light at a frequency found in the sodium spectrum but not the spectrum of potassium, only the sodium atoms in the flame absorb the light. The sodium content can thus be measured with no interference from the potassium.

FIGURE 17.13 The atomic absorption spectrometer



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Interactivity Atomic absorption spectroscopy (AAS) (int-6356)

SAMPLE PROBLEM 10

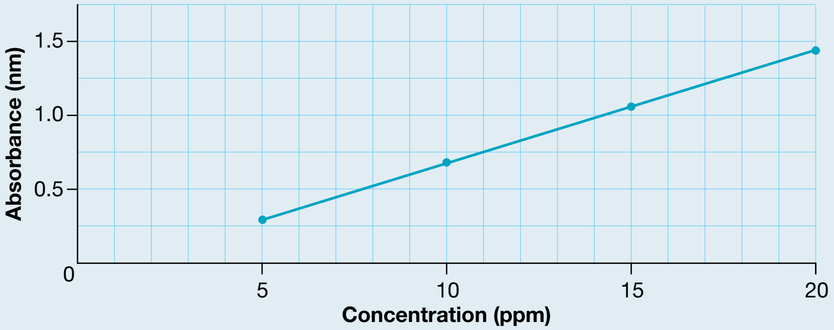
A sample of drinking water was collected for analysis of its sodium (ion) content.

Immediately prior to analysis by AAS, it was diluted by the addition of an equal volume of deionised water.

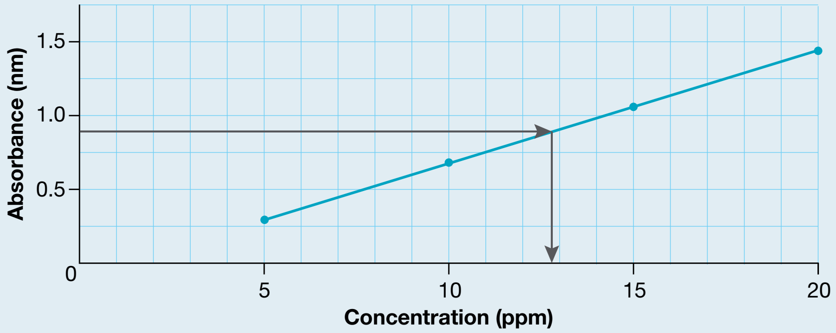
A set of sodium standards were analysed, followed by the diluted sample. The results of absorbance versus concentration of the sodium are shown in the following table.

Solution concentration (ppm)	Absorbance reading (nm)
5	0.367
10	0.712
15	1.110
20	1.470
test sample	0.989

Use the results to determine the level of sodium in the original sample.

THINK	WRITE
1. Plot the results from the table to create a calibration curve.	

2. Draw a line from the y-axis to intersect the calibration curve. (Always start from the given information; this time, the provided information is the absorbance of the test sample, of 0.989.) Then draw a vertical line down to the x-axis, to determine the concentration (ppm).



3. Remember the sample was diluted by a factor of two. Calculate the concentration in the undiluted sample.

Concentration (diluted sample) = 13.5 ppm.
Concentration (undiluted sample) = 13.5 × 2
= 27ppm

Note: 27 ppm is the same as 27 mg L^{-1}.

Such a level may need to be taken into account for people on low-sodium diets.

PRACTICE PROBLEM 10

A second sample of water suspected to contain a higher level of sodium was analysed on the same instrument as in sample problem 10, using the same standards. Due to its suspected higher level of sodium, it was diluted by adding 90 mL of deionised water to a 10 mL sample of the sampled water.

AAS produced a reading of 0.790 from the diluted sample.

What is the concentration of sodium (in mg L^{-1}) of this sample?

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Digital document Experiment 17.1 Zinc content of cornflakes using AAS (doc-30854)
Teacher-led video Experiment 17.1 Zinc content of cornflakes using AAS (tlvd-0641)

For quantitative use, both instrumental colorimetry and atomic absorption spectroscopy require calibration by taking readings from a set of known standards.

17.4 EXERCISE

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1. Silver tarnishes partly because of the presence of small amounts of hydrogen sulfide, H_2S (a gas that originates from the decay of food and smells like rotten eggs), according to the reaction:

$$4Ag(s) + 2H_2S(g) + O_2(g) \rightarrow 2Ag_2S(s) + 2H_2O(l)$$

   (a) How many moles of silver sulfide form from the complete reaction of 1 mole of silver?
   (b) How many moles of hydrogen sulfide react with 1 mole of silver?
   (c) How many moles of silver sulfide form from 3.5 moles of hydrogen sulfide?

2. Methane is the main component of the natural gas that we use as a fuel. When methane burns in air, the following reaction takes place:

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

   (a) How many moles of methane are needed to react with 1 mole of oxygen gas?
   (b) How many moles of oxygen are needed to react with 0.1 moles of methane?
   (c) How many moles of carbon dioxide are produced from 0.1 moles of methane?
   (d) How many moles of water are produced from 0.1 moles of methane?
   (e) How many moles of carbon dioxide are produced by 0.1 moles of oxygen gas?
   (f) How many moles of oxygen gas react completely with 0.25 moles of methane?
   (g) How many moles of water are produced from 8 moles of oxygen gas?

3. Calculate the mass of water that is produced when 2.8 g of methane is burned in air.
4. Propane burns in air according to the equation:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

   (a) What mass of water is produced by 6.5 g of oxygen?
   (b) What mass of O_2 reacts with 1.7 g of C_3H_8?
   (c) What mass of CO_2 is produced by 0.50 moles of propane gas, C_3H_8?
   (d) How many grams of propane gas are needed to produce 5.92 g of $CO_2(g)$?
   (e) How many kilograms of $CO_2(g)$ are released into the atmosphere when the entire contents of a 5.0 kg cylinder of propane are used at a barbecue?

5. A solution is made by dissolving an unknown amount of barium chloride in water. It is then added to a solution of potassium carbonate such that all the barium chloride reacts. The equation for this reaction is:

$$K_2CO_3(aq) + BaCl_2(aq) \rightarrow 2KCl(aq) + BaCO_3(s)$$

   (a) What information in the equation tells you that barium carbonate is a precipitate?
   (b) If 4.582 g of barium carbonate is formed, calculate the mass of barium chloride that was originally dissolved.

6. Consider the following balanced equation:

$$Cd(NO_3)_2(aq) + Na_2S(aq) \rightarrow CdS(s) + 2NaNO_3(aq)$$

   (a) Calculate the mass of CdS produced from 235 mL of a 0.178 M solution of Na_2S with excess $Cd(NO_3)_2$ present.
   (b) What is the significance of $Cd(NO_3)_2$ being present in excess?

7. In a laboratory experiment, a strip of magnesium weighing 2.56 g was placed into 200 mL of a hydrochloric acid solution. Bubbles of gas were observed and identified as hydrogen. The next day, no bubbles were observed in the beaker so the magnesium strip was removed from the beaker, dried and reweighed. Its mass was recorded as 0.350 g.
   (a) Why were no bubbles observed in the beaker on the second day of the experiment?
   (b) Calculate the concentration of the acid.
   (c) Identify the sources of error in this experiment.
   (d) Outline the safety precautions that should be taken in this experiment.
8. The calcium content of a sample can be determined gravimetrically by first precipitating the calcium ions present as calcium oxalate. When this precipitate is heated, it decomposes completely to produce calcium oxide.

In testing the purity of a calcium chloride sample, a chemist used gravimetric analysis and obtained 8.81 g of calcium oxide from an original 18.0 g sample of calcium chloride.

Calculate the percentage purity of the calcium chloride.

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17.5 Desalination

BACKGROUND KNOWLEDGE

- The process for the removal of salts and other unwanted substances from sea water and waste water through distillation, reverse osmosis and new technologies to create a water supply for human use

17.5.1 Desalination methods

Desalination is the process of removing salt and other unwanted substances from sea water and waste water. It is becoming an increasingly important method for producing water that is fit for humans to use in the future, with approximately thirty countries worldwide relying, at least in part, on desalinated water. In Australia, the drought of 1997–2009 created severe water shortages, and the resulting construction of desalination plants for seawater.

Two main methods of desalination are used around the world. These are **distillation** and **reverse osmosis**. Both of these methods use a large amount of energy. This results in the fresh water produced being higher in cost than fresh water from traditional sources. However, with further research and attention to energy-saving measures, it is anticipated that these energy requirements will be reduced, resulting in desalination becoming an economically viable alternative. New developments use advances in membrane technology and nanotechnology. These are much more energy efficient than the traditional large-scale methods, with some of them also showing promise for small-scale, ‘point-of-use’ applications.

FIGURE 17.14 Desalination plant in the Arabian Desert emirate of Dubai

Another problem is the disposal of the waste products. Removing salt and other substances from water means that a waste product is created in the form of residual water, in which is dissolved everything that was removed to make the fresh water. This waste has to be disposed of carefully. Although the residual water can be pumped back into the ocean, care must be taken to avoid causing localised areas of higher salinity and thus affecting the local marine flora and fauna. In inland areas, care must be taken to prevent such disposal contaminating existing fresh water supplies such as rivers, lakes and aquifers.

Distillation

You may recall from Unit 1 that distillation is a process that can separate the components of crude oil using a process known as *fractional distillation*. In more general terms, distillation can be used to separate a mixture of substances that have different boiling temperatures. If the boiling temperatures are similar, fractional distillation is required. However, if there is a large difference in boiling points, as is the case with ionic solutions, **simple distillation** will suffice.

For simple distillation in the laboratory, as shown in figure 17.15, the impure mixture is placed in the distillation flask and heated. As the temperature rises and the solvent begins to boil, its vapours rise and are led into the condenser. This then cools the vapours and results in condensation back into droplets of the pure solvent. As these droplets coalesce, they run down the condenser and into the collecting flask. The impurities are left behind in the distilling flask.

Distillation is used on a large scale to desalinate water. To save energy, it is nearly always performed in a series of stages, in a process called **multi-stage flash distillation**, with successive stages being at lower and lower pressures. This allows the water to be boiled at lower and lower temperatures and results in a significant saving in heating costs. A typical arrangement is shown in figure 17.16.

FIGURE 17.15 Simple distillation in the laboratory

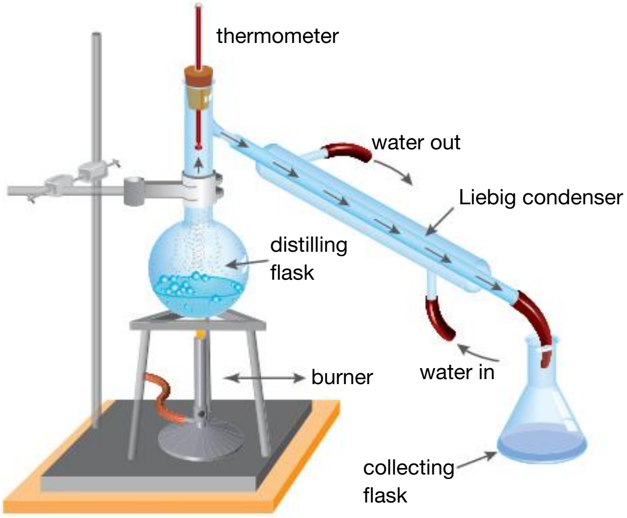
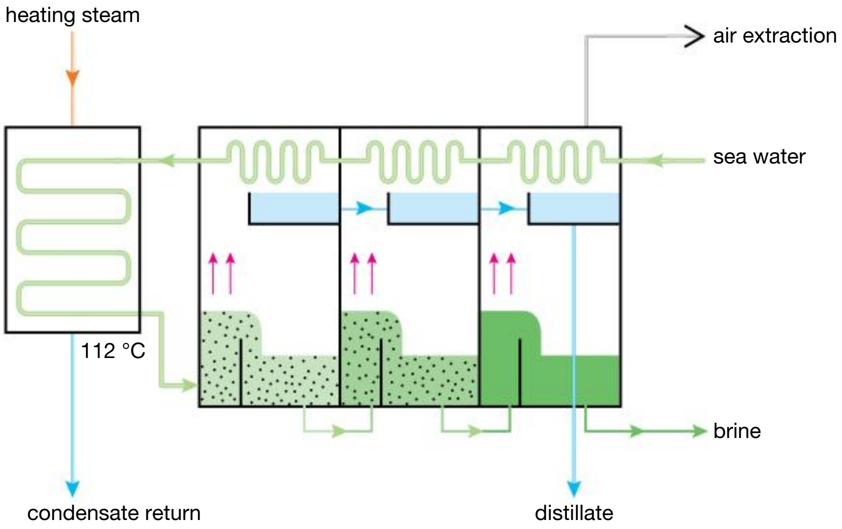


FIGURE 17.16 Multi-stage flash distillation saves energy by lowering the water pressure. This causes the water's boiling point to be reduced, meaning that less heating is required.



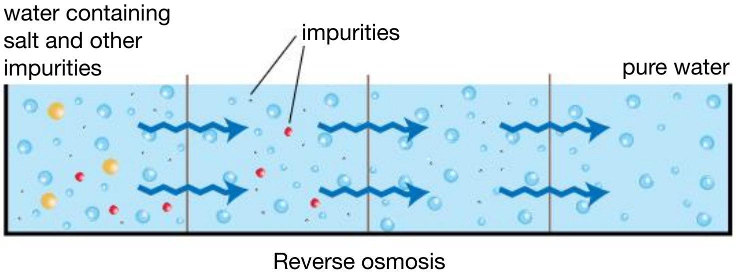
Note that the process involves a number of stages, each one cooler and at a lower pressure than the one before it. The steam in each stage is cooled by pipes containing the cold, incoming sea water to form the distillate (pure water). The incoming sea water is, in turn, heated by the steam through which it passes.

Reverse osmosis

Osmosis is the diffusion of a solvent through a membrane towards an area of higher concentration of dissolved substances. This is important in many biological situations as a means of moving water and is a spontaneous process that does not require energy. To purify water, however, we want the water to move away from a region of dissolved substances so that it can subsequently be collected as pure water. This process is called **reverse osmosis**, and it cannot occur without energy.

When used to desalinate water, this process forces pressurised water through a series of membranes. These membranes contain tiny holes (or pores) that allow the smaller water molecules to pass through but prevent larger particles of contaminants. To prevent clogging of these membranes, they are often used in series, with each successive membrane becoming more selective. Figure 17.17 shows a summary of this process.

FIGURE 17.17 Energy is applied to move water from an area of dissolved salts through a membrane to create pure water.



New developments in desalination and water purification

Recent developments in **nanotechnology** have suggested many new and exciting possibilities for increasing the supply of drinkable (potable) water and lowering its cost.

Nanofiltration membranes can be made from materials such as carbon nanotubes and work in a similar way to reverse osmosis. However, they operate at a lower pressure, resulting in energy and cost savings. They also appear easier to clean after becoming fouled due to blockage of their pores. Nanomaterials that could be used in desalination and water purification include carbon nanotubes, graphene, nanoparticles of metals (especially silver and gold) and nanoparticles made from various metal oxides.

Another technique under development is the use of nanocomposite materials to enable distillation of sea water in the presence of sunlight. These materials convert sunlight into heat energy and, when added in small amounts to sea water, can lead to a significant increase in the temperature of the water when compared with solar heating alone. This results in higher rates of evaporation, where the water vapour produced can subsequently be condensed to produce pure water. Nanomaterials made from graphene and nanosilver, and from graphene and nanogold, appear to be very suitable for this application. An additional advantage is that nanosilver has been shown to have antibacterial properties.

Nanomaterials can also be ‘functionalised’ by adding various chemical groups to promote adsorption and sterilisation of substances in impure water. Their high surface area means that they are highly efficient at adsorption or when acting as catalysts in reactions to remove impurities. This means that small-scale devices that operate at the ‘point of use’ may soon become available. Such devices could be used following natural disasters, for example, when large-scale water treatment and purification plants may be rendered inactive. They could also help produce potable water from impure supplies in many developing countries. An example of such a device currently under development is a straw-like device called a ‘waterstick’.

17.5 EXERCISE

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1. Consider the chemical, economic, social and ethical factors of desalination plants. Construct a table listing the advantages and disadvantages of desalination.

Fully worked solutions and sample responses are available in your digital formats.

17.6 Review

17.6.1 Summary

Chemical analysis

- Chemical analysis is the process of determining the substances present in a test sample.
- The choice of a particular technique depends on many factors, such as physical properties, chemical reactions of the substance and degree of accuracy required.
- Qualitative analysis is the process of determining which substances are present in a sample.
- Quantitative analysis is the process of determining how much of a substance is present.

The sources of salts in water

- Many pollutants dissolve in water because it is such a good solvent. Some waste is biodegradable and can be broken down by natural processes, but much is non-biodegradable. This poses environmental problems where species may be endangered due to water pollution.
- The salts dissolved in water samples come from a variety of sources, including natural sources as well as the result of human activity. They include a range of salts as well as heavy metal compounds and organometallic substances.
- Salts are ionic compounds that produce mobile ions when dissolved in water that can conduct an electrical current.
- Salinity (total level of combined salts) is measured using electrical conductivity.

Analytical methods

- Stoichiometry is the study of the relative amounts (ratios) of substances involved in chemical reactions.
- Mass–mass stoichiometry uses known masses of reactants or products to calculate masses of another reactant or product using mole rations; $n = \frac{m}{M}$
- Mass–concentration stoichiometry is used in reactions involving solids and solutions, and uses the known mass or concentration of reactants or products and mole ratios to calculate the mass or concentration of another reactant or product; solids use $n = \frac{m}{M}$, solutions use $n = c \times V$
- A four-step approach to solve stoichiometric problems is:
  1. Write a balanced chemical equation for the reaction, identifying the known (given) and unknown (required) quantities of substance.
  2. Calculate the number of moles of the known quantity of substance.
  3. From the equation, find the mole 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.
  4. Calculate the required quantity of the substance.
- Whenever two or more reactants are given in a chemical reaction, the limiting reactant must be identified before we can calculate the maximum amount of product that may form. A limiting reactant is completely used up in a chemical reaction. The other reactants are known as excess reactants and are not used up.

- The choice of instrument for a particular analysis depends on the properties of the material that is under analysis.
- Gravimetric analysis is analysis by mass.
- A simple method of determining moisture content is to gently heat a sample and record the mass lost as the water is driven off.
- A more sophisticated version of determining moisture content involves the formation of suitable precipitates, the amounts of which are stoichiometrically related to the amount of the component being analysed.
- This method involves weighing the sample to be analysed, dissolving the sample, forming the required precipitate, filtering the precipitate, and weighing the precipitate to constant mass.
- The exact details of a given method depend on the properties of the substances involved.
- If the properties of the substances are not carefully considered, or if mistakes are made, the final result will be affected. It is important to be able to predict how such situations will affect the final calculated result.
- Instrumental colorimetry can be used for coloured solutions. It relies on measuring how much light of a complementary colour is absorbed by a solution.
- Atomic absorption spectroscopy (AAS) and UV–visible spectroscopy both rely on the accurate measurement of radiation of a particular frequency.
- AAS is suitable for determining many metals.
- All the instruments mentioned in these analytical methods may be calibrated for quantitative analysis by obtaining readings from a number of standards and then plotting a calibration curve.

Desalination

- Desalination is the process of removing salt and other unwanted substances from water.
- Two common methods for doing this are distillation and reverse osmosis.
- New developments in membrane technology and in nanotechnology will offer further methods for desalination of water in the future.

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17.6.2 Key terms

atomic absorption spectroscopy analytical method for detecting metal species
desalination the process of removing salt and other unwanted substances from sea water and waste water to produce water for human use
distillation the process used to separate a mixture of substances due to differences in boiling point
excess reactant reactant left over when the limiting reagent has been used up
gravimetric analysis analysis by mass of precipitate
instrumental colorimeter instrument used to determine the concentration of coloured compounds in solution through comparison of a test sample against samples of known concentration through production of a calibrations curve (absorbance versus concentration)
limiting reactant the reactant that is completely used up in a reaction
mole ratio represents the relative number of moles of each substance involved in the reaction
multi-stage flash distillation an energy-efficient process to desalinate water that occurs in multiple stages of decreasing pressure allowing boiling to occur at successively lower temperatures
nanofiltration a filtration method using membranes with nanometre-sized pores
nanotechnology technology at a dimension of less than 100 nanometres

osmosis diffusion of a solvent through a membrane towards an area of higher concentration of dissolved substances
qualitative analysis use of experimental procedures to determine which elements are in a substance
quantitative analysis use of experimental procedures to determine the percentage composition of a compound or the percentage of a component in a mixture
reverse osmosis movement of a solvent under hydrostatic pressure through a membrane towards an area of lower concentration of dissolved substances; commonly used to purify water
simple distillation the process to separate two liquids of different boiling points
standard solutions a solution that has a precisely known concentration
stoichiometry the amount and ratio of reactants and products in a chemical reaction
UV–visible spectroscopy technique used to study compounds that absorb light in the UV–visible region

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Digital document Key terms glossary — Topic 17 (doc-30909)

17.6.3 Practical investigations

Experiment 17.1

Zinc concentration of cornflakes using AAS

Aim: To interpret some typical data obtained using atomic absorption spectroscopy

Digital document: doc-30854
Teacher-led video: tlvd-0641

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Digital document Practical investigation logbook (doc-30910)

17.6 Exercises

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17.6 Exercise 1: Multiple choice questions

1. Which of the following would be most likely to be a contaminant in a sample of surface water?
   - A. Cl^-
   - B. SO_4^{2-}
   - C. Ca^{2+}
   - D. Hg^+
2. A field technician collected a creek water sample (CW) and added two preserving agents (A and B) to her sample. In her field kit was also a container of deionised water (DW).

   The protocol also called for the preparation of a blank.

   Which substances should the blank contain?
   - A. A, B and CW
   - B. A, B and DW
   - C. A, B
   - D. DW only

3. A calibration curve is a graph of instrumental reading against
   - **A.** mass.
   - **B.** time.
   - **C.** volume.
   - **D.** concentration.
4. Before digital printing, sodium thiosulfate, $Na_2S_2O_3$, known as 'hypo' by photographers, was used to remove excess silver bromide, AgBr, in the liquid-based film-developing process according to the equation:

$$2Na_2S_2O_3(aq) + AgBr(s) \rightarrow Na_3Ag(S_2O_3)_2(aq) + NaBr(aq)$$

   In the *balanced* equation for the reaction between sodium thiosulfate and silver bromide
   - **A.** 1 mole of sodium thiosulfate solution produces 1 mole of sodium bromide solution.
   - **B.** 3 moles of sodium thiosulfate solution produce 2 moles of $Na_3Ag\ (S_2O_3)_2$.
   - **C.** 1 mole of sodium thiosulfate solution produces 2 moles of sodium bromide solution.
   - **D.** 1 mole of silver bromide produces 1 mole of $Na_3Ag\ (S_2O_3)_2$.
5. Phosphorus may be prepared from calcium phosphate according to the equation:

$$2Ca_3(PO_4)_2(s) + 6SiO_2(s) + 10C(s) \rightarrow P_4(s) + 10CO(g) + 6CaSiO_3(s)$$

   How much phosphorus can be produced if 1000 kg of calcium phosphate is used completely?
   - **A.** 50 kg
   - **B.** 100 kg
   - **C.** 500 kg
   - **D.** 400 kg
6. A gelatinous precipitate of iron(III) hydroxide may be prepared according to the equation:

$$FeCl_3(aq) + 3NaOH(aq) \rightarrow 3NaCl(aq) + Fe(OH)_3(s)$$

   A solution containing 16.23 g of $FeCl_3$ is mixed with 300 mL of a 1.00 M NaOH solution. Which of the following statements is correct?
   - **A.** $FeCl_3$ is in excess.
   - **B.** NaOH is in excess.
   - **C.** $FeCl_3$ is the limiting reagent.
   - **D.** Neither reagent is in excess.
7. We want to analyse a sample of salty water for its chloride ion content. This is to be done using gravimetric analysis. Which of the following chemicals could be added for the formation of a precipitate?
   - **A.** $CaCO_3$
   - **B.** $Ba(NO_3)_2$
   - **C.** K_2SO_4
   - **D.** $AgNO_3$
8. A sample of lawn food is analysed gravimetrically. As part of the procedure, a precipitate of barium sulfate is formed. However, when this is later weighed, it is still slightly damp. As a result of this, the calculated percentage of sulfate would
   - **A.** be increased.
   - **B.** be decreased.
   - **C.** be unchanged.
   - **D.** vary in a random manner.

9. An atomic absorption spectrometer was used to determine the amount of iron lost by peas in the cooking process.

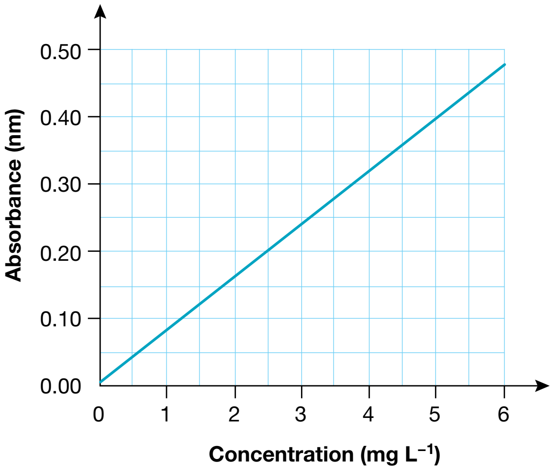
Using a suitable means of extraction, the iron was first extracted from a 5.0 g sample of uncooked peas to produce 100 mL of extract.

When tested in the instrument, the uncooked peas gave an absorbance reading of 0.20.

The AAS was then calibrated and its calibration curve is shown in the figure.

What is the mass of iron (in mg) in the original sample of uncooked peas?

A. 0.21
B. 0.25
C. 2.1
D. 2.5



10. Using the technique of AAS, we want to accurately determine the concentration of some copper(II) sulfate solution. The concentration of this solution is known to be about 0.02 M.

A set of standards is available, the concentrations of which range from 2 to 16 mg (of copper) per litre.

By what factor should the copper(II) sulfate solution be diluted prior to being drawn into the instrument?

A. 1
B. 10
C. 100
D. 1000

17.6 Exercise 2: Short answer questions

1. Classify the following analyses as either qualitative or quantitative.
   - **a.** Analysis of the propellant from a spray can by an instrument, to check the claim that it does not contain chlorofluorocarbons, CFCs
   - **b.** Adding a piece of ‘testape’ to a diabetic’s urine to estimate the glucose level present
   - **c.** Instrumental checking of the level of mercury in a sample of fish
   - **d.** Analysis of a hair-colouring preparation ‘containing less than 1.5% aromatic nitro amines’, to check the claim
2. A number of qualitative tests are used to identify various substances. Find out the tests that are used to identify the following.
   - **a.** Carbon dioxide gas
   - **b.** Starch
   - **c.** Oxygen gas
   - **d.** Protein in a food sample
   - **e.** Hydrogen gas
   - **f.** Water
3. A solution containing 5.10 g of barium nitrate reacts completely with a solution of sodium sulfate. The unbalanced equation for this reaction is:

$$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow NaNO_3(aq) + BaSO_4(s)$$

   - **a.** Balance the equation by inserting the necessary coefficients and identify the precipitate formed.
   - **b.** Calculate the mass of precipitate formed.
4. Barbecues burning charcoal briquettes are unsafe for indoor use because of the colourless, odourless, poisonous gas produced.
   - **a.** Find the $n(O_2)$ gas that reacts with 3.5 g of charcoal briquettes (assume pure C) to produce carbon monoxide.

b. If there is a plentiful supply of air, a safe colourless, odourless gas is produced. Find the mass of this gas produced if the same amount of charcoal is burned. You will need to write another equation.

5. In respiration, the equation for the reaction that produces energy in our bodies is:

$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$

If 8.90 g of glucose is used, determine the following.

a. The mass of oxygen needed

b. The mass of carbon dioxide produced

6. The silver used in jewellery and tableware becomes tarnished when exposed to air containing small amounts of hydrogen sulfide. The tarnish is a layer of silver sulfide.

$$4Ag(s) + 2H_2S(g) + O_2(g) \rightarrow 2Ag_2S(s) + 2H_2O(g)$$

Calculate the mass of the tarnish when 0.025 g of silver is reacted.

7. The metal tungsten, used to make the filaments for incandescent light bulbs, can be obtained from its oxide by reduction with hydrogen.

$$WO_3(s) + 3H_2(g) \rightarrow W(s) + 3H_2O(g)$$

a. What mass of tungsten can be obtained from 200 g of its oxide?

b. What mass of hydrogen is required?

8. As an emergency procedure, the *Apollo 13* astronauts used lithium hydroxide to remove carbon dioxide from the interior of their crippled spacecraft as it returned from the moon. Like all hydroxides, lithium hydroxide forms the appropriate metal carbonate when it reacts with carbon dioxide.

The equation for this reaction is:

$$2LiOH(s) + CO_2(g) \rightarrow Li_2CO_3(aq) + H_2O(l)$$

A possible alternative chemical for this process might have been the more readily available sodium hydroxide.

a. Calculate the mass of carbon dioxide that could be removed per kilogram of lithium hydroxide.

b. Write the equation for the reaction between sodium hydroxide and carbon dioxide.

c. From part b, calculate the mass of carbon dioxide that can be removed per kilogram of sodium hydroxide.

d. Use your answers to parts a and c to suggest a reason for the choice of lithium hydroxide rather than sodium hydroxide in a spacecraft.

e. Derive the ionic equations for both the reactions mentioned in this question.

9. Limestone is an important raw material for industry. When heated in a kiln, it decomposes to form quicklime, CaO, and carbon dioxide.

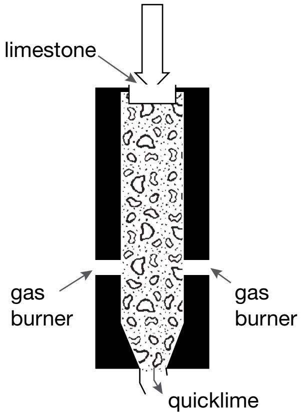
The equation for this reaction is:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

A typical kiln is shown in the figure. The production of lime from such kilns was one of Australia's earliest industries.

a. If limestone containing 83.5% calcium carbonate is used, calculate the mass of quicklime that would be produced from 100 tonnes of this limestone.

b. This process is sometimes called 'lime burning'. Why is this not a correct term to use?



10. What mass of copper is required to react completely with 250 mL of 0.100 M $AgNO_3$?

$$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$

11. Examine figure 17.4, showing measured salinity levels at Morgan in South Australia.
    a. Explain why electrical conductivity can be used to measure the level of salt in the river.
    b. Describe the trends in salt content before the early 1980s and after the early 1980s.
    c. Why is it incorrect to interpret these results as evidence of sodium chloride content?
    d. Why is it more correct to interpret the results in terms of total salt content?
12. Flame emission spectroscopy is similar to atomic absorption spectroscopy, except that 'transmittance' is measured rather than 'absorbance'.

    Sodium levels in mineral waters were analysed by flame emission spectroscopy. Sodium produces an orange light, with one of the wavelengths present being 589 nanometres (589×10^{-9} metres). The detector of the instrument was, therefore, set to measure the intensity of the emitted light at this wavelength.

    Sample standards were run first, followed by two brands of mineral water. The results are shown in the following table.

Results from analysis of sodium levels in mineral water

Sodium (mg L^{-1})	% transmittance
0	0
5	10
10	22
20	44
30	66
40	85
50	100
Sample 1 (Australian)	86
Sample 2 (French)	12

    a. Plot a calibration curve using the sodium standards.
    b. From this curve, deduce the sodium levels in each of the samples analysed.
13. The level of mercury in various waterways can have serious implications for human health if foods such as oysters or flake produced in such waters are consumed. Consequently, various agencies regularly measure mercury levels in such waterways and in food samples produced from them.

    In one analysis, a sample of water was analysed in an instrument and gave a reading of 0.375.

    Without altering the setting of the instrument, standard samples containing known levels of mercury were then analysed, giving the results shown in the following table.

Results from analysis of water sample containing mercury	
Concentration of mercury sample (mg L^{-1})	Reading
2	0.102
4	0.205
6	0.307
8	0.410
10	0.512

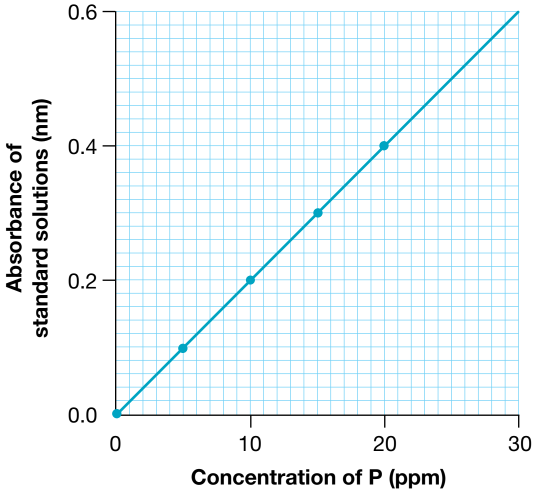
a. Plot a calibration curve using the data in the table.
b. What is the level of mercury in the sample of water tested?
c. What type of instrument do you think was used in this analysis?

14. The level of phosphorus (as phosphate) in a detergent may be determined colorimetrically by reacting it with a molybdate solution to develop an intense blue compound.

In the analysis of one particular brand of detergent, a 1.000 g sample was dissolved in 1000 mL. A carefully measured 20.0 mL sample of this was then treated with molybdate solution and heated to develop the blue colour.

When analysed in an instrumental colorimeter, a reading of 0.260 was obtained.

Four standard solutions containing phosphorus levels of 5, 10, 15 and 20 ppm were similarly treated and their absorbances measured to produce the graph at right.



a. What is the concentration of phosphorus in the diluted solution?
b. What is the percentage by mass of phosphorus in the detergent?
c. What colour do you think the light used for this analysis would be?

15. In each of the following situations, we want to analyse gravimetrically the component written in bold type. From the list in brackets, choose the ions that should be added to achieve this.
a. **Sulfate** (magnesium, aluminium, lead, ammonium)
b. **Chloride** (sodium, barium, calcium, silver)
c. **Carbonate** (potassium, lithium, barium, ammonium)
d. **Calcium** (nitrate, acetate, sulfate, hydroxide)

16. 'Every compound is soluble to a greater or lesser extent.' Discuss the implications of this statement for gravimetric procedures. Would this lead to a systematic or a random error? Explain.

17. The following steps were considered by a student when designing a gravimetric procedure to estimate the level of magnesium carbonate in a sample of Epsom salts. The steps have been written down in random order. Rearrange these steps into an order that would be acceptable, remembering that some steps might be repeated more than once and that some steps might not be required.

A. Filter and collect the precipitate produced.
B. Add a solution of potassium chloride.
C. Accurately weigh out a sample of Epsom salts.
D. Accurately weigh the precipitate.
E. Dissolve the Epsom salts in a volume of water.
F. Dry the precipitate.
G. Measure out an accurate volume of water.
H. Add a solution of barium chloride.

17.6 Exercise 3: Exam practice questions

Question 1 (6 marks)

A sample of limestone is analysed for its calcium carbonate content as follows. A 25.00 g sample is crushed and heated to a high temperature. It is then cooled, weighed and reheated until a constant mass is obtained. The mass remaining at the end of this process is 11.64 g.

a. What gas is evolved during this process? **1 mark**

b. Write the equation for the decomposition of calcium carbonate in the limestone at high temperature. **1 mark**

c. Use the results obtained to calculate the mass of calcium carbonate in the limestone, and hence its percentage purity. **3 marks**

d. State one important assumption in this method. **1 mark**

Question 2 (4 marks)

Phosphorus is an important plant nutrient and is, therefore, present in many fertilisers. It can be determined gravimetrically by carrying out a number of steps, eventually leading to its precipitation as insoluble $Mg_2P_2O_7$. In one such analysis, a 14.298 g sample of fertiliser yielded 4.107 g of $Mg_2P_2O_7$ precipitate.

a. Calculate the mass of phosphorus in the precipitate. **2 marks**

b. Hence, calculate the percentage of phosphorus in the fertiliser. **1 mark**

c. State one important assumption that is made in this method. **1 mark**

Question 3 (7 marks)

As part of a research project into tidal effects in a river estuary, a student takes water samples at various locations and determines the concentration of salt (as sodium chloride) in each. Her method involves collecting 10.0 L of water from each location and boiling it to reduce the volume. Excess silver nitrate solution is then added to precipitate all the chloride ions present as silver chloride. This precipitate is then collected, dried and weighed to constant mass.

Following this procedure, one particular sample produced 4.463 g of precipitate.

a. Write a balanced equation for the reaction that leads to the formation of the precipitate. **1 mark**

b. Calculate the mass of sodium chloride in this particular sample. **3 marks**

c. Hence, calculate the concentration of sodium chloride in this sample, in g L^{-1}. **1 mark**

d. In what way does boiling the sample to reduce its volume affect your calculation? **1 mark**

e. Some important assumptions are made by the student in this procedure. Name one. **1 mark**

17.6 Exercise 4: studyON Topic Test online only

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AREA OF STUDY 2
HOW ARE SUBSTANCES IN WATER MEASURED AND ANALYSED?

18 Analysis for organic compounds in water

18.1 Overview

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18.1.1 Introduction

Many chemicals used in farming are organic in nature, and methods of analysis exist to monitor their levels in water and soil to make certain that such levels are within acceptable limits. If a crop for animal or human consumption is sprayed with organic pesticides, it is also important to test for any residues remaining in the food produced from such crops. Canola is a source of oil used for cooking, and in 'meal' form it is used in the livestock industry. Recently, it has also been used as a feedstock for the manufacture of biodiesel. In 2018, experts estimated that over 3.6 million tonnes of canola were being grown in Australia, making it the third largest agricultural crop in the country.

The analysis of soil and water for pesticide residues is also important for 'organic' farms, which have to be free of pesticide residues and face strict certification guidelines. These often take years to comply with and involve demonstrating that not only is the produce pesticide free, but also the soil and water do not contain any residues from chemicals that might have been applied in the past.

FIGURE 18.1 A canola crop being sprayed to protect it from insect infestation. It is important that such chemicals do not have unwanted effects on the environment and that they stay where they are applied.

In this topic, you will examine the sources of organic contaminants in water, the principles of chromatography and the application of high-performance liquid chromatography (HPLC) to analyse for organic compounds.

18.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- sources of organic contaminants found in water (may include dioxins, insecticides, pesticides, oil spills)
- the application of high performance liquid chromatography (HPLC), including the use of a calibration curve and retention time to determine the concentration of a soluble organic compound in a water sample (excluding details of instrument).

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Digital documents Key science skills (doc-30903)
Key terms glossary — Topic 18 (doc-30933)

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To access key concept summaries and practice exam questions download and print the **studyON: Revision and practice exam question booklet** (doc-30934).

18.2 Organic compounds in the environment

KEY CONCEPT

- Sources of organic contaminants found in water (may include dioxins, insecticides, pesticides, oil spills)

18.2.1 Revision of organic chemistry

From simple molecules such as methane to incredibly complex molecules such as proteins, organic chemistry is all around us. Not surprisingly, organic substances have a wide range of properties. For example, some are very soluble in water while others are virtually insoluble. Many are harmless. Some are critical to life while others are among the most deadly substances known. Many are produced deliberately whereas others are produced as undesired by-products of chemical reactions.

Organic chemistry is the study of carbon compounds. It is generally understood that such compounds contain carbon–hydrogen bonds, thus making hydrogen the second major element that is present. Other elements may also be involved to a lesser degree, such as oxygen, nitrogen, the halogens, sulfur and phosphorus. These can have significant effects on the properties of organic molecules. As expected, the predominant bonding type is covalent molecular.

The ability of carbon to bond with itself to form single, double and triple bonds, as well as its ability to form cyclic structures, means that carbon can form a huge range of compounds. You may wish to review topic 8 to recall the different hydrocarbon groups and their naming rules.

Naming organic compounds

Every organic molecule can be assigned a formal name using IUPAC rules, as explained in topic 8. However, because these names can be extremely long for more complicated molecules, common practice is to refer to such compounds by either abbreviations or more common, shorter names. This is the case for many of the organic compounds used as part of our everyday lives. For example, the explosive TNT is systematically called trinitrotoluene, while the insecticide diazinon is systematically known as O, O-diethyl O-4-methyl-6-(propan-2-yl)pyrimidin-2-yl phosphorothioate!

Solubility in water

Organic substances show a wide range of solubility in water. Some are very soluble while others are virtually insoluble. This is determined by the polarity of their molecules.

Polar molecules have a difference in charge at each end of the molecule, which causes attraction to other polar molecules. If hydrogen is bonded to oxygen, fluorine or nitrogen, a stronger intermolecular hydrogen-bond is formed. Polar molecules dissolve readily in water because water is also polar. Alternatively, molecules containing large regions of carbon–hydrogen bonds are non-polar. These do not dissolve in water but are quite soluble in typical non-polar solvents such as oil and fat deposits. This illustrates the ‘like-dissolves-like’ rule.

18.2.2 What are organic compounds used for?

Modern society uses organic compounds for a large variety of purposes. One of the most well known examples are fuels. Crude oil is extracted in enormous amounts and then refined to make our familiar transport fuels as well as a wide range of other products. Many of these other products are, in turn, used to make further products. This forms the large and diverse petrochemical industry. Other familiar products that rely heavily on organic chemicals in their production include, but are not limited to, plastics and polymers, pharmaceuticals, detergents, solvents, dyes and personal care products.

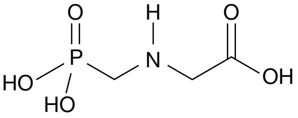
Another very large sector for organic chemicals are **pesticides**. From products that can be used at a personal level, such as insect repellent, to products used over millions of hectares in agriculture, this represents an enormous and dynamic sector in which organic chemistry plays a pivotal role. This group can be subdivided according to the specific pest being targeted. We, therefore, speak of insecticides, herbicides, fungicides, rodenticides and so on. Due to the large areas over which many of these chemicals are applied, there is always a danger that they might end up as contaminants in the soil and in water supplies if they are not applied with care and due diligence.

18.2.3 Sources of organic contaminants in water

Organic substances can enter water supplies in a number of ways. If a substance is water soluble, it can dissolve and be transported in run-off water and storm water. A number of modern herbicides are deliberately designed to be water soluble to enable their uptake through the leaves or roots of plants. This is why many such products include a warning not to spray if rain is expected within 24 hours; rain would wash the herbicide away before it had penetrated the plant. Glyphosate is a good example of this.

Contaminants that are not water soluble enter waterways by **adsorption**. Adsorption is when the contaminant sticks to the surface of a solid, such as a soil particle, by weak intermolecular bonds. The contaminant can occur as atoms, ions or molecules in liquid, gas or in dissolved form. Soil particles with adsorbed contaminants can be washed into rivers and streams and eventually be deposited as silt in these watercourses. Adsorbed contaminants on dust particles may also be spread through the air during activities such as ploughing fields and sowing seed. Prevailing weather conditions can transport adsorbed materials long distances.

FIGURE 18.2 Glyphosate — a common ingredient in many brands of herbicide.

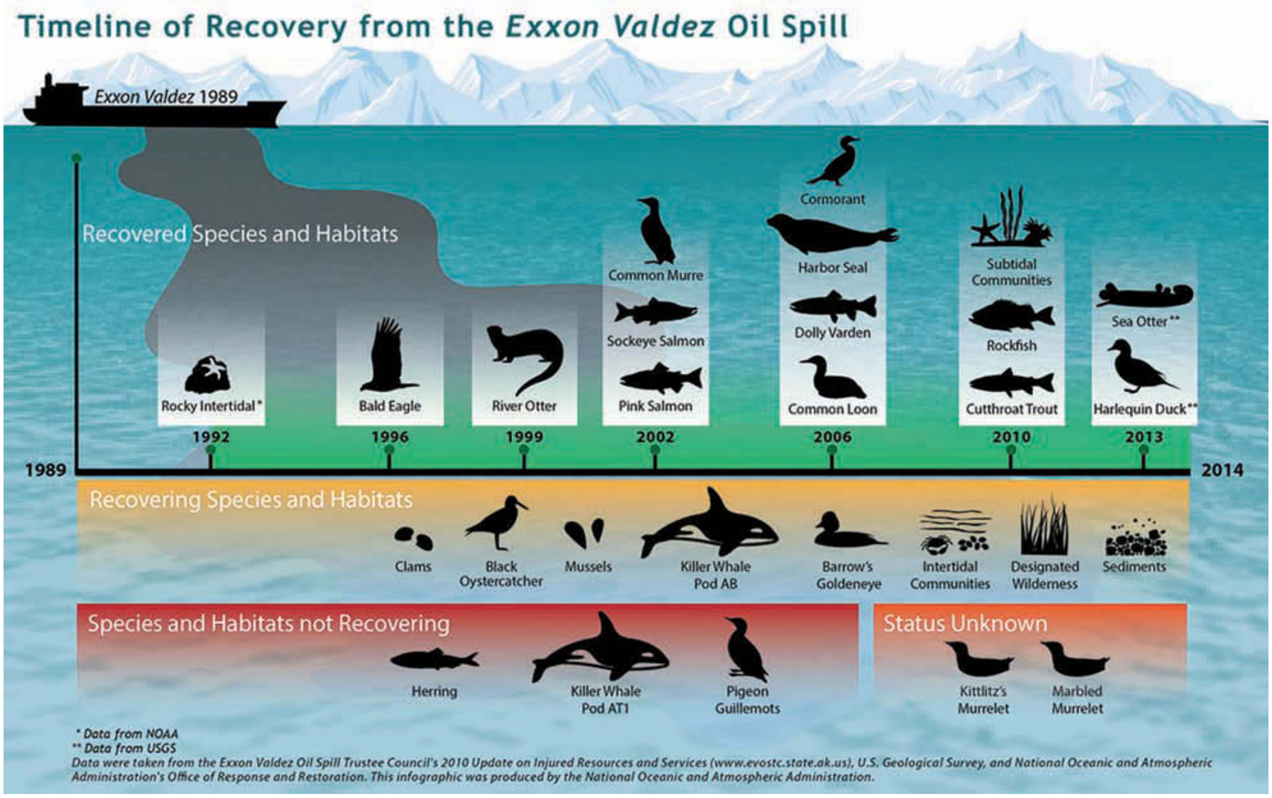


Persistent organic pollutants (POPs) are a group of organic compounds that contaminate the environment for long periods of time. Due to their widespread use before their negative properties were fully realised, they are referred to as 'hand-me-down poisons', and include DDT, dioxins and polychlorinated biphenyls (PCBs). DDT was used extensively in the 1940s to 1960s but is now banned; PCBs have been used in the electrical industry but are now being replaced with less dangerous alternatives. Dioxins are produced as unwanted by-products of combustion and have no purpose (see topic 15). Adsorbed POPs have been found in water and soil on every continent — even Antarctica. These compounds are persistent because they are fat, rather than water, soluble. This means they can concentrate in the fatty tissues of animals, where they accumulate because they cannot be processed by digestion. Concentrations increase higher up the food chain as animals and plants containing POPs are consumed in a process known as **biomagnification**.

Oil spills

Serious causes of oil spills include natural disasters, breakdowns and accidents and deliberate acts. Most major oil spills occur at sea or at the shoreline when transferring oil to tankers, a fact directly attributable to the huge amounts that are transported around the world and to the large amount of offshore drilling today. They can cause environmental disasters because the oil is difficult to remove from an uncontained body of water. On 20 April 2010, the *Deepwater Horizon* oil rig exploded while drilling for oil in the Gulf of Mexico. Over the following weeks, it was estimated that as much as 780 000 m^3 of oil was spilled into the Gulf of Mexico, making it the largest spill in American history. In 1989, the tanker *Exxon Valdez* ran aground in Prince William Sound, Alaska, and spilled over 42 000 m^3 (approximately 42 million litres) of oil into a sensitive marine environment, the effects of which are summarised in figure 18.3. This spill is remembered for the huge impact it had on wildlife in the sensitive area of Alaska in which it occurred. In 1991, oil was deliberately released and set on fire as part of the First Gulf War.

FIGURE 18.3 This timeline shows the recovery of the 28 types of animal, plants and marine habitats that were seriously affected by the 1989 spill.



Timeline of Recovery from the *Exxon Valdez* Oil Spill

* Data from NOAA
** Data from USGS
Data were taken from the Exxon Valdez Oil Spill Trustee Council's 2010 Update on Injured Resources and Services (www.evostc.state.ak.us), U.S. Geological Survey, and National Oceanic and Atmospheric Administration's Office of Response and Restoration. This infographic was produced by the National Oceanic and Atmospheric Administration.

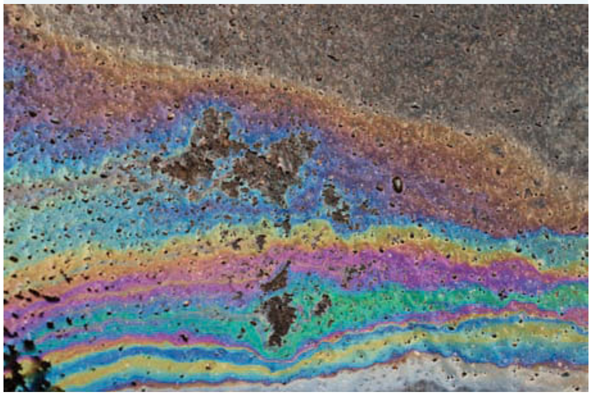
The effects of an oil spill depend on the type of oil that is spilled and where it occurs. Marine spills are difficult to contain and usually affect large areas if left uncontained.

When oil is spilled at sea, because oil is a mixture of mostly non-polar hydrocarbons, it does not dissolve in the water but floats on it instead. Over time, the oil spreads out, due to its low surface tension, to form a thin layer called a 'slick'. Slicks are usually very thin but cover a very large area. Some of the more volatile hydrocarbons in the oil evaporate, and the oil left behind tends to form a more viscous sludge. The action of waves and other physical phenomena then break this up and some clumps may even sink, while other portions of the spill may eventually be washed up on nearby land. Over time, sunlight, bacteria and biological processes begin to break down the oil. On land, it may gradually become covered by natural processes such as silting and sand deposition.

These spills are directly harmful to all forms of marine life. Birds and mammals are affected when their feathers and fur become contaminated by oil and are no longer water repellent. They may drown or suffer hypothermia as a consequence. Ingestion of oil as they try to clean themselves can also be harmful. The effects can be more subtle, however. Affected plants may cause indirect effects further up the food chain. These effects are felt for many years.

FIGURE 18.4 Pelican covered with oil after an oil spill in California, 2015

FIGURE 18.5 Road surfaces become coated with oil, which runs off into drains after rain.



Oil spills occur in many other ways, on land as well as at sea. Often, these represent rather small amounts that happen over a period of time and hence they do not grab the headlines as the previous examples already discussed did. Examples of these include:

- *ground seepage*. Underground oil deposits make their way to the surface, usually through faults in the underlying rock layers. Once on the surface, they can collect in pools, run off into water courses or evaporate into the atmosphere.
- *road run-off*. Oil leaking from vehicles drops onto the road surface. This is then washed away when it rains and is one of the main reasons wet roads are so slippery after a spell of dry weather. This oil can often be seen as a rainbow sheen on the road as it becomes wet.
- *leakage from pipes and storage*. Either accidental or intentional, leaks can occur from oil storage and distribution facilities.
- *dumping of used oil*. It is now illegal to dump used oil (for example, sump oil from motor vehicles). However, oil from a number of old dumping sites can still leak into the environment today.

HOW ARE OIL SPILLS TREATED?

No two oil spills are the same. Weather, type of oil, amount spilled, the natural environment and proximity to human habitation all affect how a spill is treated. A response usually involves one or more of the following procedures:

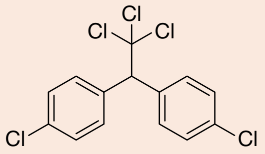
- *Leave it alone.* This may be considered if the spill is relatively small and a large distance from land.
- *Containment.* Floating surface barriers may be deployed to trap floating oil within their perimeter. These can be effective in calm conditions but not when waves are large or when they drift up on shore. Devices such as skimmers or special absorbing materials are often used in conjunction to remove the trapped oil.
- *Application of dispersants.* These surfactants help to disperse the oil into droplets and encourage it to mix with water. The increased surface area that results can also help the natural degradation processes of the oil.
- *Accelerated biodegradation.* Bacteria and other micro-organisms that naturally break down the oil into fatty acids and carbon dioxide can be added and encouraged to proliferate by adding nutrients. This is sometimes used where oil washes up on land.
- *Burning.* Sometimes the oil is deliberately ignited, preferably close to the source of the spill and before more volatile components have evaporated.

FIGURE 18.6 Floating barriers such as these booms are often used to help contain an oil spill.

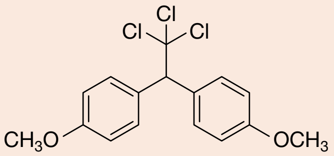
18.2 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

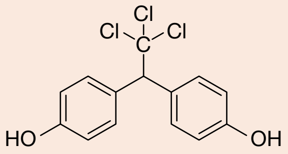
1. Identify three uses of organic compounds.
2. What influences the solubility of an organic substance?
3. (a) Define the term ‘organic chemical’.
    (b) Which two elements are the most common in most organic chemicals?
    (c) What other elements are often present in smaller amounts?
4. A problem with DDT is that it is not biodegradable. Because of this, it builds up in the fatty tissues of organisms and can reach quite high levels in organisms higher up food chains. Scientists have attempted to overcome this problem by producing variants that are still persistent but can be metabolised by animals into water-soluble products. One such insecticide is methoxychlor. The following diagram shows the structures of DDT, methoxychlor and a typical metabolite of methoxychlor.



(a) DDT

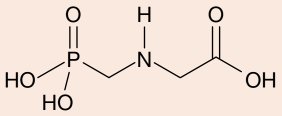


(b) methoxychlor



(c) a typical metabolite of methoxychlor

    (a) Why is it desirable that an insecticide be persistent in the environment, at least to a reasonable degree?
    (b) Explain why the product of methoxychlor metabolism shown in the diagram would be water soluble and, therefore, less likely to build up in fatty tissue.
5. Use the structure of glyphosate shown to write its molecular formula.



6. List the ways that organic chemicals can enter water bodies.
7. Give three reasons an oil spill may occur in the following environments.
    (a) At sea
    (b) On land

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18.3 High-performance liquid chromatography (HPLC)

KEY CONCEPT

- The application of high performance liquid chromatography (HPLC), including the use of a calibration curve and retention time to determine the concentration of a soluble organic compound in a water sample (excluding details of instrument)

18.3.1 Introduction to chromatography

Chromatography is the process where a **mobile phase** (which contains the mixture to be analysed) moves over a **stationary phase** (which contains a material that has a high surface area). As this happens, substances in the mobile phase continually form weak temporary bonds with the stationary surface through adsorption. This is followed by the bonds breaking and allowing the substance to move on. This is called **desorption**. These bonds are similar to the weak bonds that form between molecules.

Because differences exist in how strongly this occurs, substances move over or through the stationary phase at different speeds and are separated from one another as a result.

You may have noticed how certain food dyes, if mopped up with a damp piece of paper towel, spread out into different colours. You may have done simple experiments where a spot of ink is placed onto a piece of filter paper and water allowed to soak up and move the spot, eventually spreading it out into bands of different colours. Figure 18.7 shows such an experiment. The green ink has separated into blue and yellow components. The red ink has one single component, and the black ink has many. The colours that move furthest are those that adsorb and desorb most weakly, while the colours with stronger adsorption and desorption do not move as far.

FIGURE 18.7 Paper chromatography is a simple and inexpensive form of chromatography.

18.3.2 High-performance liquid chromatography (HPLC)

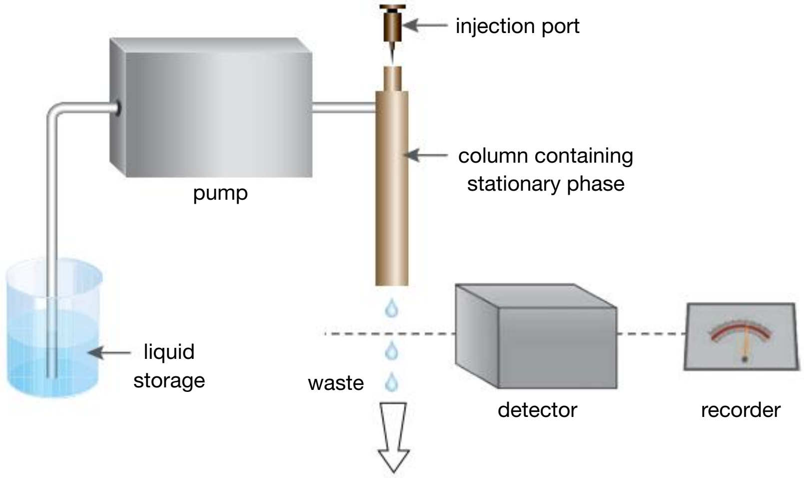
HPLC (originally called high-pressure liquid chromatography) is an adaption of the simple method described in the previous section. Invented in the early 1960s, advances in technology have resulted in this now becoming an extremely sensitive and widely used modern technique. Detection of concentrations in parts per million and parts per billion levels is routine. Advanced instruments can now even detect parts per trillion. It is an excellent method for measuring the concentration of organic compounds in a sample, with applications including research, medicine, pharmaceutical science, forensic analysis, food analysis, drug detection in sport and environmental monitoring.

FIGURE 18.8 In HPLC, the separated components are captured after they have been eluted through the column, which is the steel tube on the left of this photo.

In HPLC, the stationary phase is a narrow diameter tube that is packed tightly with a finely divided powder called an **adsorbent**. This provides the high surface area required for the process and is called a ‘column’.

Columns containing a wide range of adsorbent powders are available and can be chosen to give optimal results for a particular scenario. Columns packed with finely divided alumina or silica are common examples. The sample to be analysed is injected at the top of the column as a liquid, and a further liquid (called the **eluent**) is then pumped through the column at high pressure. This liquid 'flushes' the mixture through the column, allowing adsorption and desorption of components in the mixture to take place. This is the mobile phase. In the same way as paper chromatography, this results in the components of the mixture moving through the column at different rates. Eventually each component reaches the end of the column and exits from it (this is called **elution**). The time taken for this to occur is called the **retention time** (R_t). A detector at this point records retention times in a variety of ways (for example, as a graph on a strip chart (called a **chromatogram**) or on a computer). Figure 18.9 shows the essential components of a HPLC.

FIGURE 18.9 Schematic of high-performance liquid chromatography



Resources

Video eLesson High-performance liquid chromatography (HPLC) (eles-0347)

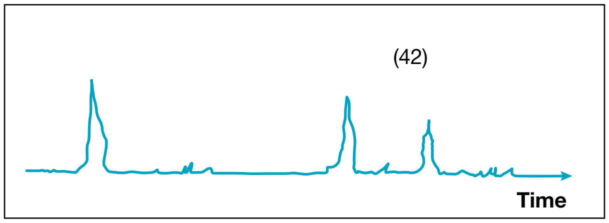
18.3.3 Interpreting the results from HPLC

To understand how HPLC can be used to obtain quantitative measurements, let us consider a mixture containing three different substances: A, B and C. These produce three different signals from the detector as they elute from the column; this is illustrated in figure 18.10. If component C is the one that we are interested in, how can we distinguish its signal from those produced by A and B? The answer is simple. We inject a known sample of C by itself into the column and run it through under identical conditions. By noting retention times, we can then easily identify the correct signal from the mixture. (We assume that substance C takes the same time to move through the column by itself as it would in a mixture containing A and B.)

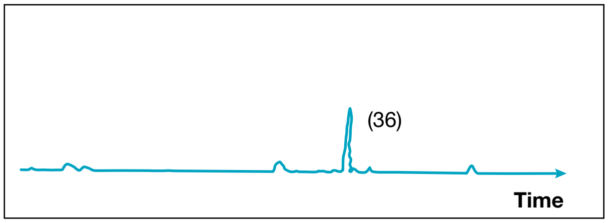
For quantitative analysis, the *size* of the signal is measured in some way. The higher the concentration of the substance causing the signal, the stronger the reading produced. Commonly, this reading is either *peak area* or *peak height*, terms that refer to the peaks produced on the chromatogram. This means that it is possible to produce a calibration curve in the same way as for atomic absorption spectroscopy (AAS; refer to topic 17). A calibration curve is a graph of concentration versus instrumental reading; in HPLC, this reading is either peak height or peak area. A number of standards (containing substance C) at accurately known concentrations are run through the instrument. The readings obtained are then graphed against concentration to produce a calibration curve. The concentration of substance C can then be easily determined from this graph.

FIGURE 18.10 Schematic of the HPLC process (the numbers in brackets represent the peak height value for each tested concentration)

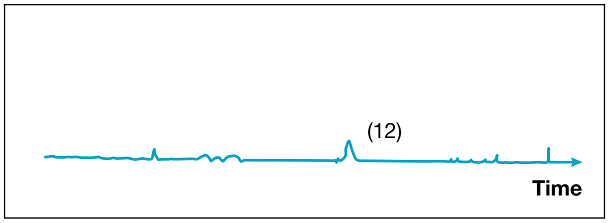
The HPLC process



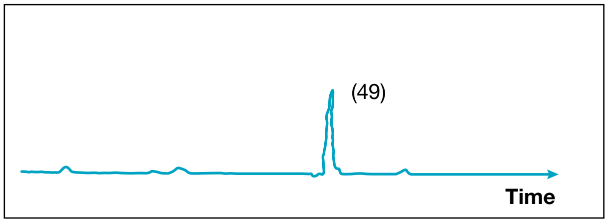
Chromatogram of mixture containing A, B and C
(Which peak is which?)



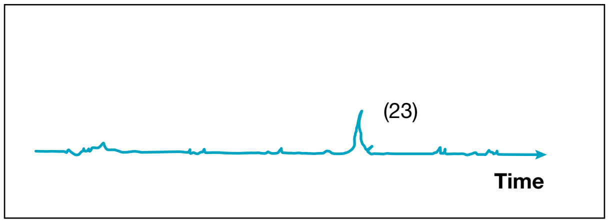
Chromatogram of C at 15 mg L^{-1}



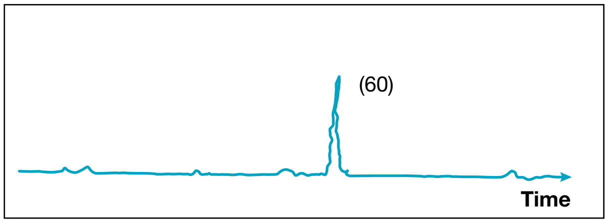
Chromatogram of C at 5 mg L^{-1}
(Peak C is now easily identified.)



Chromatogram of C at 20 mg L^{-1}

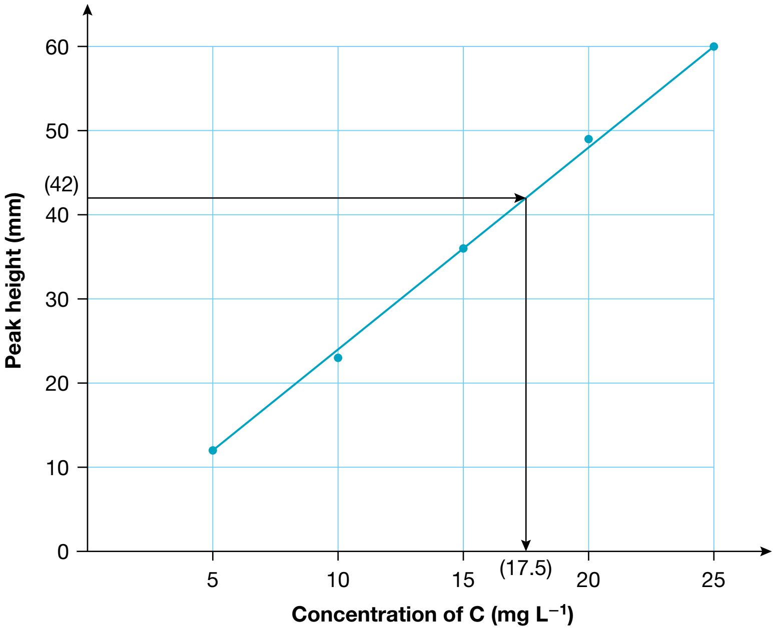


Chromatogram of C at 10 mg L^{-1}



Chromatogram of C at 25 mg L^{-1}

A calibration curve can now be drawn (peak height vs concentration).



Finally, the peak height from the mixture can be read as a concentration.
Concentration of C = 17.5 mg L^{-1}

on Resources

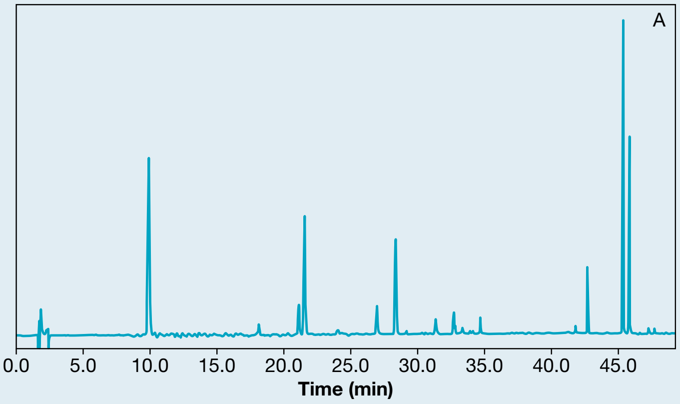
Video eLesson Calibration curves (eles-0348)

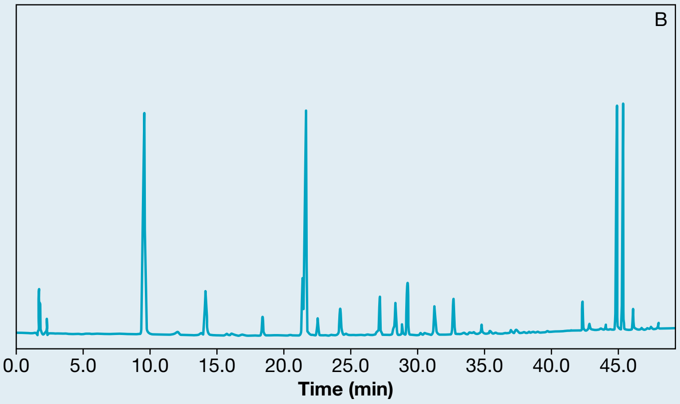
18.3.4 Application of HPLC for concentrations of soluble organic contaminants

HPLC is ideally suited to measuring the levels of dissolved organic substances in water. It has the sensitivity to detect very low levels of non-polar species. For more soluble organic substances, the sample is merely diluted first so as not to 'overload' the instrument, and the result is then adjusted to allow for the dilution. The variety of choice available for both the column and the eluent mean that the instrument can be optimised for a particular analysis.

SAMPLE PROBLEM 1

A sample of lemon oil was suspected of being adulterated. To test this suspicion, HPLC was used to produce chromatograms from a sample of pure oil (A) and from the suspect sample (B). These chromatograms are shown in the following figures.





Has this oil been adulterated?

Teacher-led video: SP1 (tlvd-0606)

THINK	WRITE
Compare the two graphs by counting the number of peaks and their peak area. Discuss the discrepancies.	Although the chromatograms are very similar, some obvious additional peaks appear in B. The most evident is the peak at about 14 minutes. The oil has been adulterated.

PRACTICE PROBLEM 1

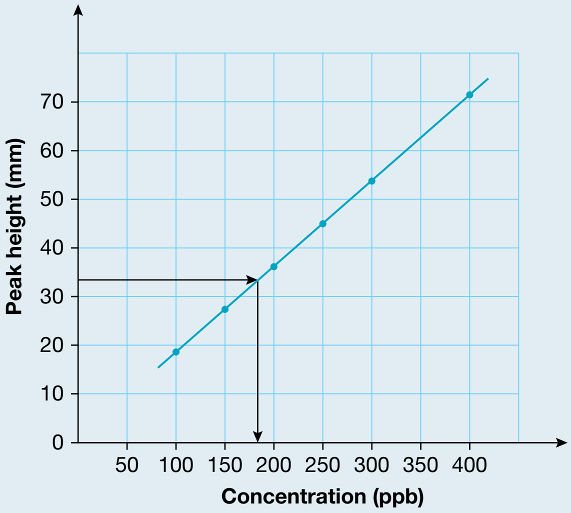
Discuss all the discrepancies you can identify in the lemon oil graph from sample problem 1.

SAMPLE PROBLEM 2

Atrazine is a herbicide that is often applied to sugar cane crops to kill weeds. In order to determine the concentration of atrazine in run-off water after rain, representative samples were collected after rainfall over a period of time. These were then analysed using HPLC. The results of one such sample are tabulated in the following table, together with results from a set of standards.

Atrazine standard concentration (ppb)	Peak HPLC height
100	19
150	28
200	36
250	45
300	54
400	72
Run-off sample	34

a. By examining the table, make a quick estimate of the atrazine level in the run-off sample.
b. Draw the calibration curve from this data and use it to accurately determine the level of atrazine.

THINK	WRITE
a. Compare the peak height of atrazine in the run-off sample with known standards.	The reading suggests that the level should be somewhere between 150 and 200 ppb.
b. Plot the graph for the concentrations given. Draw a horizontal line from the peak height of atrazine in the run-off sample at 34 mm to the calibration curve and then to the *x*-axis to derive the concentration.	 From the calibration curve shown, the level of atrazine is 180 ppb.

PRACTICE PROBLEM 2

If the run-off sample had atrazine concentration of 225 ppm, what would be the expected peak height?

18.3 EXERCISE

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1. Define the following terms as they apply to chromatography.
   (a) Stationary phase
   (b) Mobile phase
   (c) Chromatogram
   (d) Retention time
2. Explain why a set of standards do not need to be evenly spaced to produce a calibration curve.
3. A sample of orange juice was analysed for its ascorbic acid (vitamin C) content. A 100 mL sample was diluted accurately to 1000 mL in a standard flask. After thorough mixing, a 100 mL portion was filtered and prepared for HPLC analysis.

   The results from this analysis, together with those from a set of standards are shown in the table.

Ascorbic acid concentration (mg L−1)	Peak area
10	1690
15	2420
20	3105
25	3797
30	4504
Sample	3380

   (a) Plot the calibration curve for this data.
   (b) From part a, determine the concentration of ascorbic acid in the diluted sample.
   (c) Hence, calculate the concentration of ascorbic acid in the undiluted juice sample.

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18.4 Review

18.4.1 Summary

Organic compounds in the environment

- Organic compounds are used in many situations and in large amounts as part of modern living.
- Due to their widespread use, organic compounds may find their way into water supplies and other bodies of water.
- Organic compounds show a range of solubility in water. Long-chain carbon–hydrogen organic molecules are typically non-polar and hence not water soluble.

- Carbon compounds get into water supplies in a number of ways. Two common methods are by dissolving (sometimes in only small amounts) and by adsorbing (adhering) to solid particles, which are subsequently washed into waterways or transported via wind.
- POPs (persistent organic pollutants) are non-polar, fat-soluble, non-digestible molecules that concentrate in the fatty tissues of animals. They include dioxins, DDT and PCBs.
- Oils spills can occur on land as well as at sea. Spills at sea are more mobile and can affect large areas.
- Oil spills at sea can be treated by a variety of methods, either singly or in combination. These methods include leaving it alone, containment, skimming, absorbing, using dispersants, encouraging biodegradation and burning.

High-performance liquid chromatography (HPLC)

- High-performance liquid chromatography (HPLC) is a technique ideally suited to analysing the concentration of organic substances in water. It is both versatile and sensitive enough to detect the low concentrations that are sometimes present.
- Like all forms of chromatography, HPLC involves a mobile phase moving over a stationary phase. The process of adsorption and desorption that takes place results in components of a mixture moving at different rates over the stationary phase.
- Retention time is the time taken for a substance to move through a chromatography column.
- An important step in using HPLC to determine unknown concentrations is the production of a calibration curve. This involves running a set of standards of accurately known concentration through the instrument. The reading produced for each standard is noted and the results used to draw a graph of concentration against reading. This is the calibration curve.

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18.4.2. Key terms

adsorbent porous substance with a high surface area that adsorbs analytes by intermolecular forces

adsorption process by which substances adhere, or 'stick', to the surface of other particles; commonly involves liquid or gas molecules sticking to the surface of solid particles

biomagnification process by which a compound (such as a pollutant or pesticide) increases its concentration in the tissues of organisms as it travels up the food chain

chromatogram a visible record (commonly a graph) displaying the results of HPLC, showing the detector signal against time

desorption the opposite of adsorption, where molecules are released from the surface of other particles

eluent solvent used to remove analytes from the adsorbent in a HPLC column

elution in HPLC when each component reaches the end of the column and exits from it

mobile phase the mixture to be analysed in the chromatography process

pesticide a chemical substance used to kill organisms that are harmful to cultivated crops or animals

retention time time taken for a substance to travel through the chromatography column; R_t

stationary phase a material with a high surface area, used in the chromatography process

on Resources

Digital document Key terms glossary — Topic 18 (doc-30933)

18.4 Exercises

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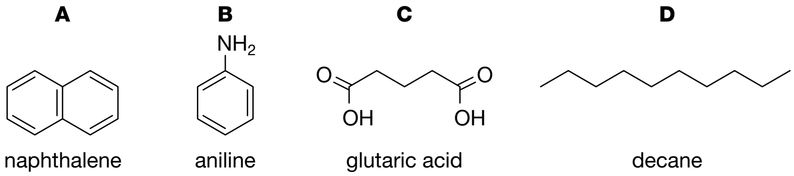
18.4 Exercise 1: Multiple choice questions

1. Which of the following is an organic substance?
   - **A.** H_2O
   - **B.** $C_{12}H_{24}$
   - **C.** CO_2
   - **D.** $CuSO_4.5H_2O$

   Questions 2 and 3 refer to the four compounds shown in the following figure. For ease of reference, these have been labelled A, B, C and D.



2. Which two would be expected to be water soluble?
   - **A.** A and B
   - **B.** B and C
   - **C.** C and D
   - **D.** D and A
3. Which two would be expected to be insoluble in water?
   - **A.** A and B
   - **B.** B and C
   - **C.** C and D
   - **D.** D and A
4. Which of the following statements about dioxins is true?
   - **A.** Dioxins have been used in the past as insecticides.
   - **B.** Dioxins are inorganic chemicals containing heavy metals.
   - **C.** Dioxins are restricted to areas with large chemical industries.
   - **D.** Dioxins are unwanted by-products that have no useful applications.
5. Despite their low solubility in water, many organic contaminants can persist in water bodies for a long time. What is a reason for this?
   - **A.** They adsorb strongly to the surface of soil particles in river beds.
   - **B.** Other liquid impurities in the water dissolve them.
   - **C.** They are broken down by natural microbial action in the soil.
   - **D.** They accumulate in the bodies of aquatic organisms, which then die.
6. What do all forms of chromatography involve?
   - **A.** Transitory phase moving over a mobile phase
   - **B.** Transitory phase moving over a stationary phase
   - **C.** Mobile phase moving over a transitory phase
   - **D.** Mobile phase moving over a stationary phase

7. In the quantitative use of a HPLC, which of the following would *not* be done?
   - A. Measuring the area of the peaks produced
   - B. Drawing a graph of concentration versus time
   - C. Drawing a graph of concentration versus peak area
   - D. Standardising the instrument's readings against known concentrations
8. When preparing a calibration curve from HPLC data, which of the following would *not* be appropriate to use?
   - A. Concentration
   - B. Retention time
   - C. Peak height
   - D. Peak area
9. What would an oil spill at sea *not* be treated by?
   - A. Containment
   - B. Dispersants
   - C. Burning
   - D. Neutralisation
10. Which of the following would produce the most significant land-based oil pollution entering the oceans?
    - A. By adhering to dust particles that are subsequently deposited at sea
    - B. By run-off from urban areas into drains and rivers and subsequently to the sea
    - C. By soaking into the ground and slowly draining to the sea
    - D. From washing off tyres of cars and boat trailers at launching ramps

18.4 Exercise 2: Short answer questions

1. Many organic chemicals are soluble in water while many others are insoluble. What features in their molecular structures account for this difference?
2. An examination of the formula of a complicated organic molecule reveals that it contains carbon, hydrogen, iodine and sulfur. Is this molecule likely to be water soluble? Explain.
3. A student used paper chromatography to produce a chromatogram from black ink, as shown in the figure. The ink was added to the paper near the bottom, and water was used as the mobile phase.
   - a. List the colours shown in order from most strongly adsorbed to least strongly adsorbed.
   - b. Predict what would happen if a longer piece of paper was used.

4. A sample of supposedly pure compound A is suspected of being contaminated with a small amount of compound B. To check this claim, a small amount is dissolved and then analysed using HPLC.
   - a. Draw a possible chromatogram produced by this sample, paying particular attention to the heights of the peaks.
   - b. A small chance exists that only one peak might be obtained. Suggest a possible reason for this.
5. An analyst is designing a procedure to use HPLC to check the level of a water-soluble chemical in a common herbicide. The claim being tested is that the level is 120 g L^{-1}. The configuration of the HPLC requires that the concentration of this chemical be between 1 and 10 g L^{-1} for reliable analysis.

Using the product claim as a guide and assuming that a 100 mL standard flask is to be used, calculate the volume of herbicide that would need to be poured into the flask and then diluted so that a final concentration of 5 g L^{-1} is achieved.

6. The level of glucose, $C_6H_{12}O_6$, in a sample of human blood was analysed by HPLC. After suitable extraction, the sample was analysed, along with a set of known standards. The results are shown in the table
   - a. Draw the calibration curve using the data shown.
   - b. What is the concentration of glucose in the blood sample (in g L^{-1})?

     The normal concentration of glucose in human blood should be 4.4 to 6.1 millimoles per litre or millimolar (mmol L^{-1} or mM).
   - c. Convert your answer from part b into mM.
   - d. Is the glucose within the normal range?

Glucose concentration (g L^{-1})	Peak area
0.4	1650
0.8	3280
1.0	4100
1.2	4920
1.6	6550
Sample	3840

7. 'Snuffem' is a water-soluble insecticide sold for domestic use to kill ants and spiders. Its label claims that its active ingredient is 'taylachlor', at a level of 2.5 mg per 100 g.

   To test this claim, a 10.045 g sample was dissolved in 100 mL of ethanol and filtered to remove the residue. The volume was then reduced to 13.5 mL by gentle heating over a water bath so as not to destroy the 'taylachlor'. Standards containing 10, 15, 20 and 30 mg L^{-1} were prepared. These standards and the condensed sample were then subjected to HPLC and gave the results shown in the table.
   - a. Determine the concentration of 'taylachlor' in the sample as tested.
   - b. Calculate the total mass of 'taylachlor' in the concentrated sample.
   - c. Is the mass of 'taylachlor' in the concentrated sample the same as in the original 100 mL sample? Explain.
   - d. Calculate the determined level of 'taylachlor' in 'Snuffem' (as mg per 100 g). Comment on how this compares to the claim on the label.

Concentration of 'taylachlor' (mg L^{-1})	Peak area
10	200
15	290
20	380
30	560
Concentrated sample	310

8. Crude oil can be classified in a number of ways. For example, *light crude oil* is crude oil that has a high proportion of volatile components. Consequently, it evaporates relatively quickly. On the other hand, *heavy crude oil* contains a high proportion of less volatile substances.

   Suggest possible responses to the following scenarios.
   - a. A small spill of light crude oil close to land
   - b. A small spill of heavy crude oil close to land
   - c. A large spill of light crude oil away from land

18.4 Exercise 3: Exam practice questions

Question 1 (9 marks)

Caffeine is widely used in a number of pharmaceutical products.

HPLC was used to determine the caffeine content in a particular brand of tablet. One tablet, weighing 0.103 g, was crushed and dissolved in exactly 10 mL of ethanol. This was subsequently diluted to 1000 mL in a standard flask. Five caffeine standards were also prepared as shown in the following table.

All six solutions were subjected to HPLC under identical conditions. The results are shown in the table.

Concentration of caffeine (mg L^{-1})	Peak area
5	370
20	1340
40	2620
60	3910
80	5200
Test sample	3400

a. Plot the calibration curve for this data. **5 marks**

b. Determine the concentration of caffeine in the diluted sample as tested. **1 mark**

c. Calculate the concentration of caffeine in the ethanol used for extraction. **2 marks**

d. What was the mass of caffeine in the tablet that was tested? Express your answer in milligrams. **1 mark**

Question 2 (8 marks)

Insecticides can move from where they are applied into other non-intended areas of the environment by a number of mechanisms. Water plays an important role in two of these. Firstly, water can dissolve an insecticide and carry it away in solution as run-off. Secondly, run-off water can carry away soil particles with surfaces that contain adhered insecticide molecules.

The potential of an insecticide to move 'off site' can be predicted by three parameters. These are its *solubility*, the strength of its *adherence* to particles in the soil and *how long* it persists before it is broken down into other products. Adherence is measured by K_{oc} values. The higher the value, the more strongly the insecticide is bound to soil particles. Field half-life (FHL) is an indication of how long the insecticide persists after being applied.

In general, an insecticide with a low FHL is unlikely to move. A chemical with a higher FHL may move in solution, by adsorption to soil particles or both. The following table shows these parameters for three insecticides.

Insecticide	Solubility (mg L^{-1})	K_{oc}	FHL (days)
Diazinon, $C_{12}H_{21}N_2O_3PS$	60	1520	40
Permethrin, $C_{21}H_{20}Cl_2O_3$	0.006	10^6	42
Oxamyl, $C_7H_{13}N_3O_3S$	2.8×10^5	25	4

a. Explain why diazinon is considered to have high potential for both solution and adsorption run-off. **2 marks**

b. Despite its high solubility, oxamyl is considered to have low potential to move. Explain why this is so. **1 mark**

c. Explain what is likely to happen to a field sprayed with permethrin if torrential rain is experienced within the first few weeks after application. **3 marks**

d. If the field in part c experiences gentle misty rain instead, is there likely to be a problem? Explain. **2 marks**

18.4 Exercise 4: studyON Topic Test onlineonly

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19 Analysis for acids and bases in water

19.1 Overview

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19.1.1 Introduction

To determine the concentration of an acid or a base in solution, a technique called titration is often used. One of the pieces of equipment used is a burette, which delivers an accurate amount of solution.

As you will discover, completing a successful titration requires careful cleaning and delicate use of specialised equipment. While professional laboratories now use automated equipment to carry out titrations, learning how to complete a titration by hand is an important skill for all chemistry students. Such is the skill required, there are even titration competitions! Titrations are used in a range of industries, including monitoring blood glucose levels, in urine tests such as for pregnancy, and to determine the levels of salt, sugar, vitamins and fats in foods.

This topic examines the sources of acids and bases in water, their influence on the pH change of the solutions and how stoichiometry can be applied to reactions in solutions. The use of stoichiometry in solutions is known as volumetric analysis, and applications explored in this topic include preparations of standard solutions, calculations in diluted solutions and the use of acid–base titrations to determine the concentration of an acid or a base in a water sample.

FIGURE 19.1 A burette used for titration

19.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- sources of acids and bases found in water (may include dissolved carbon dioxide, mining activity and industrial wastes)
- volume–volume stoichiometry (solutions only) and apply volumetric analysis, including the use of indicators, calculations related to preparation of standard solutions, dilution of solutions and use of acid–base titrations to determine the concentration of an acid or a base in a water sample.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of learning and assessment. Experiments and investigations, supported by a **Practical investigation logbook** and **Teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

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Digital documents Key science skills (doc-30903)

Key terms glossary — Topic 19 (doc-30941)

Practical investigation logbook — Topic 19 (doc-30942)

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19.2 Sources of acids and bases in water

KEY CONCEPT

- Sources of acids and bases found in water (may include dissolved carbon dioxide, mining activity and industrial wastes)

Isn't water neutral? While pure water has a pH of 7, normal surface water can have a pH anywhere between 6.5 and 8. For groundwater (underground water), this can be as low as 6 and still be considered normal, while sea water has a pH of 8.1. As we have already learned, water is an excellent solvent; therefore, virtually all the Earth's water contains dissolved substances. Some of these are dissolved in quantity while others only sparingly. It is this range of solutes that alters the pH of the water.

FIGURE 19.2 The pH scale is the measure of the acidity/basicity of a solution.

19.2.1 Acidity and basicity in water

Water becomes acidic or basic when natural acids or bases, or acids or bases that are generated by human activity dissolve in it.

How is water made acidic?

Water can be made acidic by familiar inorganic acids such as hydrochloric, sulfuric and nitric acids, which can enter the water naturally and as contaminants. Natural sources such as volcanoes and lightning flashes produce small amounts of sulfuric and nitric acids. These acids are also generated as gases in the air as the result of human activity, and are the main cause of acid rain. Even so-called unpolluted rain is slightly acidic due to dissolved carbon dioxide from the atmosphere, making the weak acid, carbonic acid, H_2CO_3.

$$CO_2(g) + H_2O(I) \rightarrow H_2CO_3(aq)$$
$$H_2CO_3(aq) + H_2O(I) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

In addition to the above sources, large numbers of salts can make water acidic. The expanded definition of acids and bases by Brønsted and Lowry (see topic 13) explains how such salts can act as acids, thus affecting the pH of water. Often, such salts contain cations, which are the conjugate acids of weak bases. Ammonium salts containing NH_4^+ ions are an example of this. In other cases, the anion may be **amphiprotic** but may react in water more easily as an acid than a base. For example, hydrogen sulfate ions can react as an acid (donating a proton, H^+) with water as follows:

$$HSO_4^-(aq) + H_2O(l) \rightleftharpoons SO_4^{2-}(aq) + H_3O^+(aq)$$

Since HSO_4^- is a stronger acid than water (refer to table 13.3), it will donate the hydrogen ion to water. In this case, HSO_4^- is more likely to act as an acid than a base, so, salts containing hydrogen sulfate ions tend to make water acidic.

Certain metal cations can also make water acidic due to the behaviour of surrounding water molecules. Aluminium ions and ions of transition metals such as iron, nickel and copper are good examples of these.

Finally, a wide range of organic materials, both natural and contaminants, can act to make water acidic.

How is water made basic?

Natural bases can be sourced from chemicals present in many rocks and minerals that dissolve as the water flows over them. Carbonates are typical examples of this. Bases generated by human activity include pollution from industrial sources.

Many salts contain anions that react as bases in water. Often, such anions are the conjugate bases of weak acids. The carbonate ion, CO_3^{2-}, which is the conjugate base of the weakly acidic hydrogen carbonate ion, HCO_3^-, is a good example of this. Dissolved salts such as sodium carbonate, Na_2CO_3, and calcium carbonate, $CaCO_3$, therefore, tend to make water basic.

Amphiprotic anions may play a role here as well. Hydrogen carbonate ions illustrate this point well. They can react as a base by accepting a proton, H^+, as follows:

$$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$$

Since HCO_3^- is a slightly stronger base than water (refer to table 13.3), it will accept the hydrogen ion from water. Therefore, the presence of HCO_3^- in water will make water more basic.

19.2.2 Alkalinity in water

Alkalinity is a measure of how well a body of water is able to withstand acid pollution.

In the environment, such as in a stream, it is important that changes in pH are not too large and/or happen too quickly. The pH of stream water is the net effect of all the substances that are dissolved in it. However, sometimes the addition of an acidic contaminant does not affect the pH as much as might be expected. This is because some of solutes in the water can act to resist changes in pH. Such substances are called **buffers**. Calcium carbonate, $CaCO_3$, and calcium hydrogen carbonate, $Ca(HCO_3)_2$, are examples of such substances. Calcium carbonate occurs naturally in limestones, and is dissolved by rainwater to produce calcium hydrogen carbonate. If an acid enters water containing these ions, it reacts with them and is neutralised. It is only after all the ions have been removed that any subsequent acid is free to significantly affect the pH. The higher the level of buffers in the water, the more acid pollution the stream is able to withstand.

Similar to acidification of water bodies is the acidication of soils. This significantly affects ecosystems and productivity of agricultural land due to leaching of nutrients and the inability of bacteria to withstand more acidic conditions.

19.2.3 Sources of acids and bases in water

The acids and bases found in water may come from a number of sources, both natural and as a result of human activity. They may be regarded as normal or contaminants depending on the circumstances.

Sources of acids and bases include:

- natural sources such as volcanoes and lightning. These can produce sulfuric and nitric acids. Carbonic acid is formed as water dissolves carbon dioxide from the atmosphere.
- rocks and minerals. Many minerals can act as acids or bases when they dissolve in water. Depending on the geology of an area, surface water and ground water may be naturally acidic or naturally basic because of this.
- decomposition of plant and animal materials. These can produce acids that dissolve in water.
- combustion of fossil fuels, including in the internal combustion engine, which produce sulfuric and nitric acids. Many fuels contain sulfur as an impurity, and the sulfur dioxide produced can ultimately form sulfuric acid. In car engines, the high temperatures produced combine nitrogen and oxygen from the air to form various nitrogen oxides. On exposure to water, these form nitric acid. Acid rain is discussed in topic 13.
- waste water from mines and their tailing sites. When sulfide ores and coal are mined, the sulfur they contain comes into contact with oxygen and water. Oxidation of the sulfur, followed by dissolution in water produces sulfuric acid. An additional problem is that this acidified water may then leach unwanted heavy metals from ores and rocks to cause further contamination.

FIGURE 19.3 Volcanoes are a natural source of nitric and sulfuric acids.

FIGURE 19.4 The Rio Tinto river in Spain has a pH of 2 from a combination of local geology rich in sulfides and over 5000 years of mining.

- manufacturing industries. Industrial processes that use acids include galvanising and electroplating, which use large amounts of both hydrochloric and sulfuric acids. Acids are also used to make dyes and tan leather. Bases such as sodium hydroxide are used to refine petroleum and to make soap. Organic acids are used to make detergents, and organic bases are used to make pharmaceuticals and various herbicides. In every case, careful attention to detail must be followed to ensure that the release of waste water from such activities meets acceptable EPA guidelines.

19.2.4 Why do we analyse for acids and bases?

Water needs to be at, or at least close to, an optimum pH for many of its uses. Knowing how much acid or base in is the water is important when its pH needs to be adjusted. Avoiding one problem could well cause another if the water is 'overadjusted'.

Problems such as metal corrosion in plant equipment and the ability of fish to survive and reproduce are typical problems where water has become too acidic. The deposition of scale in pipes and cloudiness in water are typical problems caused by basic conditions. Owners of backyard swimming pools know that the effectiveness of their chlorine in sanitising the water depends on pH. If it is too high or too low, the chlorine is not as effective. The same applies to chlorine added to domestic drinking supplies. Many plants have optimum pH ranges for the soil in which they grow, which can be altered by the water used to irrigate them. In some cases the opposite applies — where undesirable growth occurs because an ideal pH environment has been inadvertently created. Mosquito larvae live within a pH range of 3.3 to 4.7, and a pH range of 7.5 to 8.4 is optimal for algal growth. Adjusting pH values away from these ranges may solve these issues. In all of these cases, an analysis of the water involved is a key first step.

FIGURE 19.5 Scale inside steam pipes can be avoided by altering the pH of the water.

Another effect of incorrect pH can be **synergy**. This is where pH combines with one or more other factors to have an effect that is greater than the individual effects combined. For example, the presence of metals such as iron, aluminium, lead and mercury in water at the wrong pH can make the metals much more toxic than normally expected.

19.2 EXERCISE

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1. Give an example of acid pollution of water from natural causes.
2. Give an example of acid pollution of water from industry.
3. What are the typical conditions for pipes to clog? How can they be avoided?
4. Define a buffer.
5. What is the pH of natural drinking water?
6. Write the reactions to explain why rainwater is slightly acidic.
7. Define alkalinity.

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19.3 Solution stoichiometry (volume–volume stoichiometry)

KEY CONCEPT

- Volume–volume stoichiometry (solutions only) and application of volumetric analysis, including the use of indicators, calculations related to preparation of standard solutions, dilution of solutions and use of acid-base titrations to determine the concentration of an acid or a base in a water sample

19.3.1 Volumetric analysis

Acids and bases are usually analysed using **volumetric analysis**. As the name suggests, volumetric analysis involves measuring volumes of solutions. If one of the solutions is of accurately known concentration, the concentration of a second solution may be derived using stoichiometry. The more accurately volumes and known concentrations are determined, the more accurate is the derived concentration. Great care is, therefore, taken in how this is done.

Although volumetric analysis is the technique of choice for many acid–base determinations, it can also be applied to other types of reactions that occur in solution. For example, many redox reactions are analysed using this method.

Before we look at the specifics of how volumetric analysis is performed, we first need to examine how stoichiometry can be applied to reactions in solution.

Solution stoichiometry

When some solutions are mixed, chemical reactions may occur. Solution stoichiometry, sometimes also called ’concentration–concentration’ stoichiometry, involves reactions in solution such as precipitation reactions and the neutralisation of an acid and a base to form water and a salt. Solution stoichiometry determines the number of moles of solute in solutions that may react or that may be produced.

Solution stoichiometry uses the formula:

$$n = cV$$

where n = number of moles (mol), c = concentration (mol L^{-1}) and V = volume (L).

TIP: The formula $n = cV$ can be found in table 3 of the VCE Chemistry Data Book.

Solution stoichiometry 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.
2. Calculate the number of moles of the known quantity of substance present, using the formula $n = cV$.
3. From the equation, find the molar ratio of known to unknown quantities in the reaction and use it to calculate the number of moles of the required substance.
4. Calculate the quantity (i. e. concentration) of the required substance.

Solution stoichiometry also uses the same steps as for mass–mass stoichiometry (refer to topic 17). The difference is that the formula $n = cV$ is used for both the given and the required substances.

SAMPLE PROBLEM 1

A solution of cloudy ammonia is analysed for its ammonium hydroxide concentration. A 20.00 mL sample is neutralised with 30.00 mL of 1.10 mol L^{-1} hydrochloric acid. Find the molarity or concentration of the ammonium hydroxide in the cloudy ammonia.

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THINK	WRITE
1. Write the equation, identifying the known and unknown quantities of substance. Convert V units from mL to L.	$NH_4OH(aq) + HCl(aq) \rightarrow NH_4Cl(aq) + H_2O(l)$ $c(NH_4OH) = ?\ mol\ L^{-1}$ $c(HCl) = 1.10\ mol\ L^{-1}$ $V(NH_4OH) = 0.0200\ L$ $V(HCl) = 0.0300\ L$
2. Calculate the number of moles of the known quantity of HCl by recalling the formula $n = cV$.	$n(HCl) = c(HCl) \times V(HCl)$ $= 1.10 \times 0.030\,00$ $= 0.0330\ mol$
3. Find the molar ratio from the equation and use it to calculate the number of moles of the required substance.	$n(NH_4OH) : n(HCl) = 1 : 1$ $n(NH_4OH) = n(HCl)$ So, $n(NH_4OH) = 0.0330\ mol$
4. Calculate the concentration of the required substance.	$c(NH_4OH) = \frac{n}{V}$ $= \frac{0.0330}{0.02000}$ $= 1.65\ mol\ L^{-1}$

PRACTICE PROBLEM 1

A solution of 25.0 mL of 1.00 mol L^{-1} H_2SO_4 was neutralised with 0.250 mol L^{-1} NaOH. What volume of NaOH was required?

19.3.2 Dilution

Dilution of a solution changes its concentration. As the volume increases with dilution, the concentration decreases.

When a solution is diluted by the addition of more solvent (for example, water), the number of moles of solute remains the same. The addition of water to a concentrated solution does not alter the number of moles or the mass of the solute in that solution.

If n_1 represents the number of moles of the initial or concentrated solution and n_2 represents the number of moles of the final or dilute solution, we can say that:

$$n_1 = n_2$$

Using the equation $n = cV$, we can write equations for n_1 and n_2 as follows:

$$n_1 = c_1V_1 \text{ and } n_2 = c_2V_2$$

Since the values of n_1 and n_2 are equal, these equations can be combined.

The relationship between the concentration (*c*) and volume (*V*) of the initial (or concentrated) solution and the final (or dilute) solution is:

$$c_1V_1 = c_2V_2$$

SAMPLE PROBLEM 2

What volume of 10 mol L^{-1} stock hydrochloric acid is needed to prepare 250 mL of a 2.0 mol L^{-1} hydrochloric acid solution?

THINK	WRITE
1. Let c_1 and V_1 represent the quantities of concentrated solution and c_2 and V_2 represent the quantities of dilute solution. Convert volume units from mL to L.	$V_1 = \ ?\,\text{L}$ $\quad V_2 = 0.250\,\text{L}$ $c_1 = 10.0\,\text{mol L}^{-1}$ $\quad c_2 = 2.0\,\text{mol L}^{-1}$
2. Calculate the quantity of 10 mol L^{-1} HCL needed to produce the required dilution.	$$\begin{aligned} c_1V_1 &= c_2V_2 \\ 10.0 \times V_1 &= 2.0 \times 0.250 \\ V_1 &= \frac{2.0 \times 0.250}{10.0} \\ &= 0.050\text{ L} \\ &= 50\text{ mL} \end{aligned}$$

PRACTICE PROBLEM 2

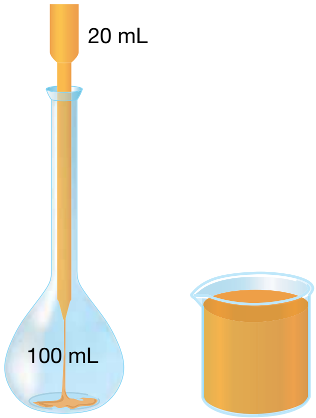
From a stock of nitric acid, a volume of 20 mL was used to prepare 250 mL solution of 0.80 mol L^{-1}. What was the concentration in the stock solution?

The dilution procedure is demonstrated in figure 19.6.

FIGURE 19.6 The dilution of a 0.20 M $K_2Cr_2O_7$ solution to 0.04 M

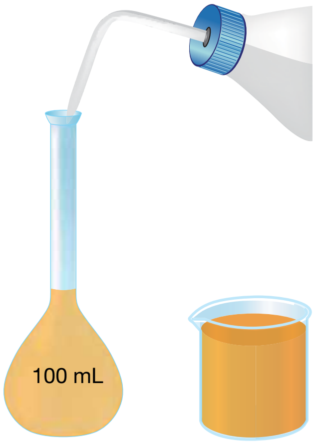
Step 1

Measure 20 mL of 0.20 M $K_2Cr_2O_7$ into a pipette. Slowly add the $K_2Cr_2O_7$ to a calibrated 100 mL flask.



Step 2

Rinse the pipette with water and also pour slowly into the calibrated 100 mL flask.



Step 3

Stopper the flask and mix thoroughly, and then add more H_2O to the 100 mL mark. The original 0.200 M $K_2Cr_2O_7$ has been diluted by a factor of 5, so the concentration decreases by a factor of 5. You now have 100 mL of 0.0400 M $K_2Cr_2O_7$.

19.3 EXERCISE

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1. In an experiment to determine the concentration of a solution of hydrochloric acid, 15.5 mL of the HCl was neutralised by 25.0 mL of a 0.055 mol L^{-1} solution of Na_2CO_3 in a conical flask.
   (a) Write a balanced equation for the reaction.
   (b) Determine the concentration of the hydrochloric acid.

2. What volume of a solution of 0.200 mol L^{-1} potassium hydroxide, KOH, exactly neutralises 20.0 mL of a solution of 0.300 M hydrochloric acid, HCl?
3. What volume of a solution of 0.200 mol L^{-1} potassium hydroxide, KOH, exactly neutralises 20.0 mL of a solution of 0.300 M sulfuric acid, H_2SO_4?
4. What volume of a solution of 0.200 mol L^{-1} potassium hydroxide, KOH, exactly neutralises 20.0 mL of a solution of 0.300 M phosphoric acid, H_3PO_4?
5. The salinity of sea water can be measured by adding silver nitrate solution until all the chloride ions have been precipitated as silver chloride according to the equation:

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

A 20.00 mL sample of sea water required 22.40 mL of 0.500 mol L^{-1} silver nitrate solution to precipitate all the chloride ions from solution. Calculate the molarity (concentration) of the sodium chloride in the sample.

6. Calculate the volume of stock solution needed to prepare 250 mL of a 2.15 M solution of hydrochloric acid (hydrochloric acid stock solution = 12 mol L^{-1}).
7. Calculate the concentration (molarity) of 2.0 L of 2.0 mol L^{-1} HCL solution after dilution with 500 mL of water.
8. If 30 mL of stock solution of initial 2.0 mol L^{-1} was diluted to 0.500 L, what is its new concentration?
9. How much water is required to dilute a solution of 25 mL of 3.00 mol L^{-1} to 0.75 mol L^{-1}?
10. If 170 mL of water was used to prepare a solution of 250 mL of 0.600 mol L^{-1}, what was the original concentration?

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19.4 Acid–base titrations

KEY CONCEPT

- Volume–volume stoichiometry (solutions only) and application of volumetric analysis, including the use of indicators, calculations related to preparation of standard solutions, dilution of solutions and use of acid–base titrations to determine the concentration of an acid or a base in a water sample

Acid–base **titration** is an experimental method that uses specialised equipment to determine the concentration of a solution, based on neutralisation reactions.

Acid–base titration is a type of volumetric analysis where the unknown concentration of a solution is determined by reacting it with a solution of known concentration. The solution with a concentration that is known accurately is called the **standard solution.**

19.4.1 Titration procedure

Successful volumetric analysis relies on accurate measurement of volumes and often, but not always, accurate knowledge of one of the initial concentrations. To this end, a number of very accurate volume-measuring pieces of equipment are used. However, their use is pointless unless the correct operating technique is practised. Correct use of some common volumetric apparatus is outlined in this section.

The standard solution is prepared by dissolving an accurate amount of solute in water using a volumetric flask (see figure 19.7a) that is calibrated to contain the specified volume. The standard solution is poured into a burette, used to deliver definite but variable volumes of liquid. In a titration, the volume of liquid measured by the burette is called a **titre**. The solution of unknown concentration is added to a conical flask (figure 19.7b) using a pipette (figure 19.7c). A pipette is used to deliver a known volume of liquid, which is then called an **aliquot**. Figure 19.8 demonstrates the correct and incorrect handling of equipment during a titration.

FIGURE 19.7 Basic equipment required in a volumetric analysis

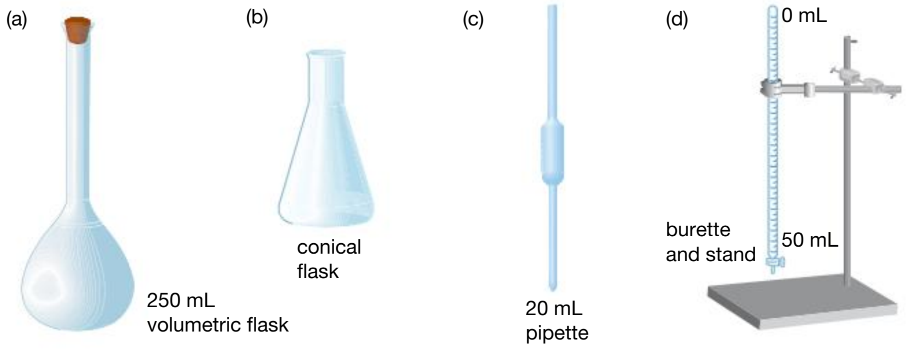
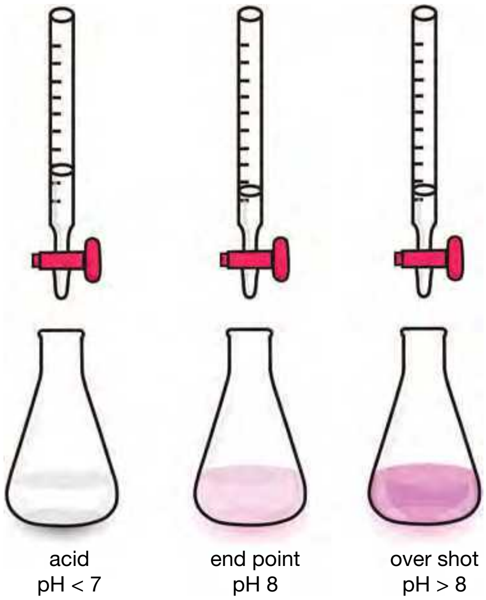


FIGURE 19.8 Correct (one hand on burette, one hand on the conical flask) and incorrect handling of equipment during a titration

A suitable indicator is added to the aliquot in the conical flask. The solution in the burette is slowly added to the aliquot until the indicator changes colour. This process is called titration. The point at which chemically equivalent amounts of acid and base (according to the equation) are present is called the **equivalence point**. The point at which the indicator changes colour is called the **end point** and is usually about one drop after the equivalence point. Titres within 0.05 mL of each other are called **concordant** titres. A minimum of three concordant titres are needed to calculate the **average titre**.

FIGURE 19.9 Acid–base titration using phenolphthalein pH indicator



on Resources

Weblink Acid–base simulation

19.4.2 The pipette

A pipette is used to deliver a fixed volume of solution, called an aliquot. Each pipette has an etched mark above the bulb section. By drawing the liquid up above this mark, and then allowing it to drain down to the mark, the pipette delivers its specific calibrated volume.

Cleanliness is essential if consistent results are to be obtained. Before using the pipette, it should be rinsed with distilled water and then rinsed with a small portion of the solution to be used. This ensures that the solution in use is neither contaminated nor diluted prior to being analysed.

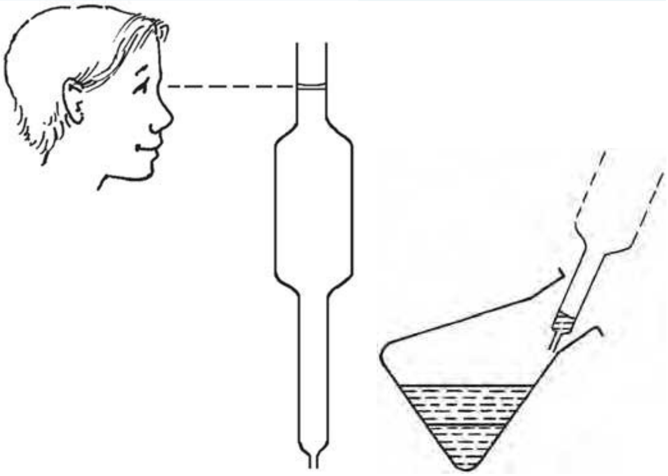
To fill the pipette, a pipette filler should be used. This enables potentially dangerous solutions to be pipetted without any risk to the users.

Once the pipette is filled to the etched mark, the liquid is allowed to drain into the conical flask. This should be done by holding the jet of the pipette against the side of the flask and allowing the liquid to drain. After draining, a small amount of solution will still be present in the pipette. This should not be removed. Calibration of the pipette takes this small amount of remaining solution into account.

Common errors in using a pipette include:

- not keeping the tip of the pipette below the level of solution when filling. This causes the solution to rush up into the pipette filler and contamination may result.

FIGURE 19.10 Correct handling of the pipette



- not reading from the bottom of the meniscus. In line with the usual procedure, pipettes are calibrated to the bottom of the curved liquid surface (meniscus).
- not lining up the etched mark at eye level. This is known as a parallax error.

19.4.3 The burette

Burettes are used to deliver a variable volume, termed a titre, to the conical flask. They are marked with an accurate scale, which is usually calibrated to tenths of a millilitre. Common practice when reading these scales is to interpolate between these marks to obtain a reading to a hundredth of a millilitre (that is, two decimal places). Simply by taking an initial reading, and then a final one, the difference between them corresponds to the volume delivered.

As with pipettes, clean equipment is essential for accurate results and a burette should be prepared in the same way as a pipette. That is, it should first be rinsed with distilled water, and then rinsed with a small portion of the solution it is to contain.

At this stage it is also a good idea to check that:

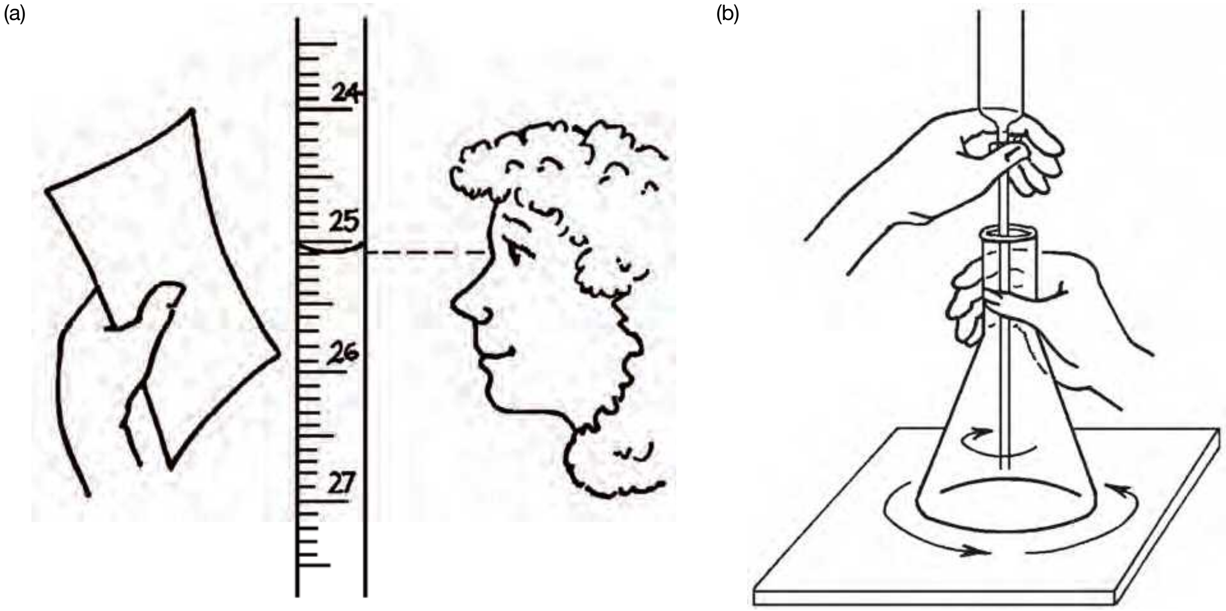
- the tap does not leak
- the tip is not blocked by a build-up of sediment.

If the burette appears to be satisfactory it should then be clamped vertically using a general purpose boss head and clamp, or a special burette clamp. The burette should then be filled with the desired solution. A small funnel should be used to prevent spillage when filling the burette, and then the funnel should be removed. In most cases the solution is added until it is above the zero mark. The tap at the bottom is then opened to drain the solution down until it is on the scale, and also to ensure that the jet of the burette is full. Note that the level of solution does not have to be exactly on the 0.00 mark — all that is required is that it be somewhere on the scale so that a reading can be taken.

For clear solutions, all readings on a burette should be taken from the bottom of the meniscus and should be to two decimal places. When reading the scale, many people find that a piece of white paper held behind the scale is often of assistance. As with any scale reading, the eyes should be level with the level of the liquid to avoid parallax error.

In use, the 'left-hand turn — right-hand swirl' technique should be adopted. This involves wrapping the fingers and thumb of the left hand around the jet and tap and using them to open and close the tap. This leaves the right hand free to swirl the flask as the solutions are mixed. Although a little awkward at first, the technique is quickly mastered and is considered most suitable for this apparatus.

FIGURE 19.11 (a) Correct reading of a burette scale and (b) the 'left-hand turn — right-hand swirl' technique



To help determine colour changes during a titration, a white tile should be placed under the titration flask. When you get near the anticipated end point, the titre should be added dropwise from the burette until the required colour change is observed in the aliquot. To further ensure accuracy, a wash bottle should be kept nearby so that any liquid that splashes onto the sides of the titration flask as it is being swirled can be washed back down into the liquid bulk.

To minimise errors from reading the titre, it is a common practice to repeat the titration several times and until **concordant results** are obtained. At least three concordant titres (volumes with readings within ± 0.10 mL of each other) are required to calculate the average titre.

19.4.4 Volumetric flasks

Volumetric flasks are also called standard flasks and are used when either making up a standard solution or accurate dilution of an existing solution is required.

Preparation and use of volumetric flasks and pipettes is very similar. They are cleaned with distilled water and, once filled to an etched line, contain exactly their designated volume. However, an important difference is that, in filling, a volumetric flask is always filled to the mark from below. The volume of solution should never be allowed to go above the mark.

19.4.5 Indicators

Acid–base indicators change colour over a narrow pH range. They allow us to stop a titration close to the equivalence point if they are chosen carefully.

The most important thing in acid–base titration is knowing when to stop. Acid–base indicators are critical in identifying this point. Certain substances have different colours depending on the pH of their surroundings. A range of indicators exist so that an appropriate choice may be made for a particular titration.

In an acid–base titration, the pH is not necessarily 7 at the equivalence point. If we remember that a conjugate acid and base are always formed in any acid–base reaction, it is possible that these new acids and bases have an effect on the pH. Knowledge of this pH for a particular titration means that an indicator can be chosen that changes colour at (or close to) the correct pH. Table 19.1 shows the details of some common indicators, and this is visually shown in figure 19.12.

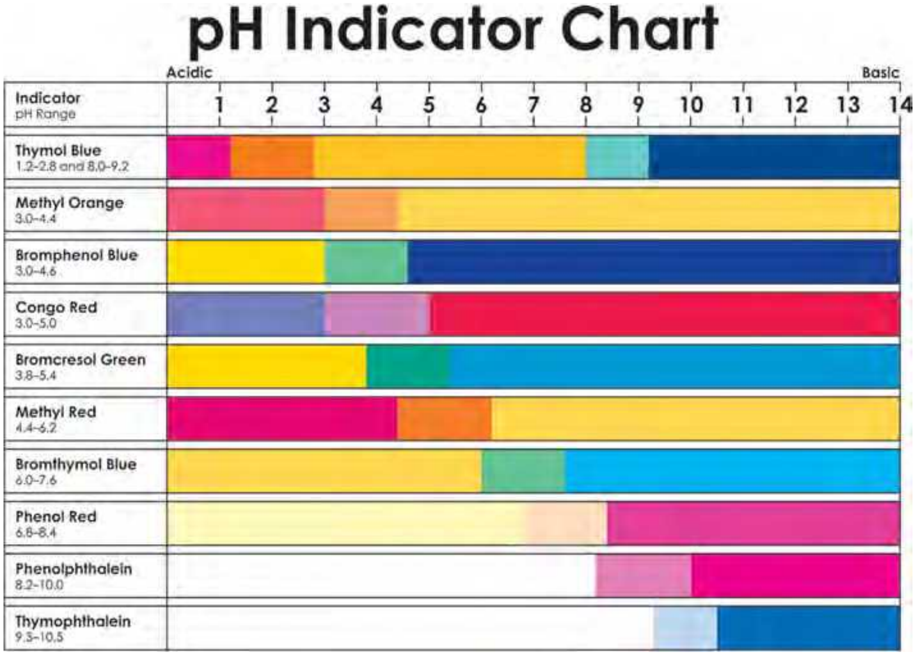
TABLE 19.1 Some common acid–base indicators

Indicator	Colour at lower pH	pH range for colour change	Colour at higher pH
Methyl orange	Red	3.1–4.4	Yellow
Methyl red	Red	4.2–6.3	Yellow
Litmus	Red	5.0–9.0	Blue
Bromothymol blue	Yellow	6.0–7.6	Blue
Phenolphthalein	Colourless	8.3–10.0	Crimson

Looking at table 19.1 it can be seen that, if we know the pH at the equivalence point of a titration is 3.2, a suitable choice of indicator might be methyl orange. On the other hand, phenolphthalein would not be suitable.

It should be realised that other factors may also influence the choice of indicator. How *quickly* the pH changes at the equivalence point is one of these.

FIGURE 19.12 pH chart for some common acid–base indicators



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Sometimes, the end point of a titration may be gauged in other ways that do not involve indicators. Some of these are by:

- using a pH meter and stopping at the correct pH.
- doing a potentiometric titration. This uses two electrodes and measures the changes in voltage produced between them.
- measuring conductivity. If the product of an acid–base reaction is a precipitate, the conductivity in the titration flask falls as ions are removed in the precipitate (remember that the other product of an acid–base reaction is water). As soon as all precipitation has finished, the conductivity increases again as excess ions are now added from the burette.

FIGURE 19.13 Titrations can be performed automatically in many modern laboratories.

19.4.6 Sources of errors

In volumetric analysis, just as in gravimetric analysis, it is important to understand the steps in a procedure and predict the effect on the final calculated result if the steps are not performed correctly. Table 19.2 shows some general areas in which mistakes could be made and what effects these mistakes have.

Other situations can also give rise to mistakes when the properties of the chemicals involved are misunderstood. For example, a common method of standardising hydrochloric acid solution is to react it with a standard sodium carbonate solution. Solid sodium carbonate needs to be anhydrous to qualify as a primary standard. However, if it is left exposed to the atmosphere for too long, it absorbs moisture and begins to rehydrate. Should this then be used to make up a standard solution, not all the mass weighed out would be sodium carbonate, due to the presence of water molecules. The resulting solution of sodium carbonate is, therefore, lower in concentration, and the effect on the calculated concentration of the hydrochloric acid would be an overestimation.

TABLE 19.2 Effect on the calculated result of some possible mistakes during volumetric analysis

	Effect if substance under analysis is in the . . .		
Situation	burette	titration flask	Comments
Rinsing water left in burette	Underestimated	Overestimated	The burette solution is diluted with water, so more is used.
Rinsing water left in pipette	Overestimated	Underestimated	The solution aliquot in the titration flask is diluted.
Indicator chosen changes colour too soon	Overestimated	Underestimated	The choice of indicator can be critical.
Water in titration flask	No effect	No effect	All necessary measurements are made before the chemicals are mixed with this water.

Uncertainties of equipment

Experiments involving glassware and other equipment will always be the source of some uncertainties of the measurements, which lead to errors associated with the equipment.

The glassware and scales used in volumetric analysis typically present the following uncertainties:

- electronic scales with uncertainty of ± 0.0001 g or ± 0.001 g
- 50 mL burettes with ± 0.02 mL for each reading
- 20 mL pipettes with ± 0.03 mL
- 250 mL volumetric flask with ± 0.1 to 0.3 mL (depending on the volume and the grade of the flask)
- 100 mL graduated beaker ± 0.5 mL.

19.4.7 Standard solutions

A standard solution is a solution with an accurately known concentration. Two methods can usually be used to determine a solution’s concentration accurately.

Method 1: by reacting it with another solution with a concentration that is known accurately.

Method 2: by taking a substance called a **primary standard** and dissolving it in a known volume of water. Primary standards are pure substances that satisfy several criteria.

To qualify as a primary standard, a substance must have a number of the following properties:

- It must have a high state of purity.
- It must have an accurately known formula.
- It must be stable. In other words, its composition or formula must not change over time. This can happen, for example, as a result of storage or reaction with the atmosphere.
- It should be cheap and readily available.
- It should have a relatively high molar mass so that weighing errors are minimised.

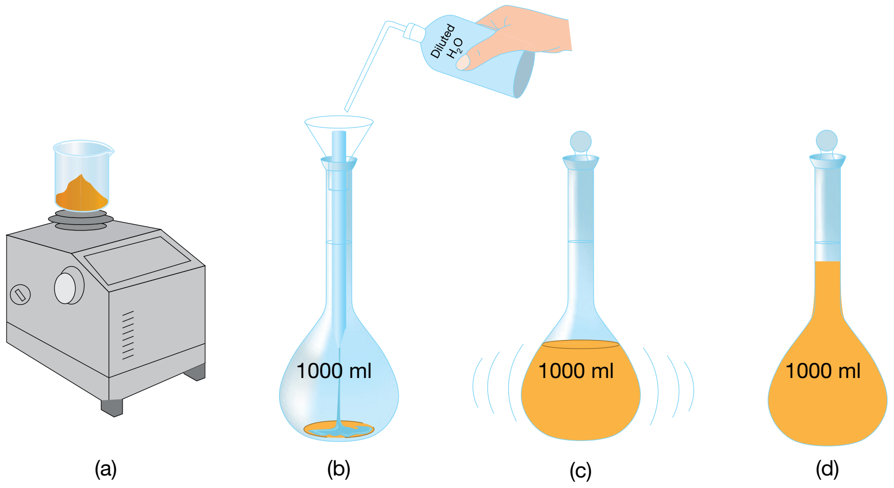
Note that this means that not all substances are suitable for use as primary standards.

A **secondary standard** can be prepared by standardising it (measuring its concentration) against a primary standard.

Preparing a standard solution

1. Weigh out the required mass of solute in a special weighing bottle or small beaker.
2. Transfer this to the volumetric flask using a dry funnel and wash the weighing bottle and funnel a number of times, each time adding the rinsings to the flask.
3. Add distilled water to approximately half the required volume, and then swirl to dissolve the solute.
4. Add further distilled water until the bottom of the meniscus is level with the etched line on the neck of the flask. Fully invert at least 15 times to ensure that the concentration of the solution is uniform.

FIGURE 19.14 Steps in preparing a standard solution



SAMPLE PROBLEM 3

Calculate the concentration of a standard solution of $K_2Cr_2O_7$, which was prepared from 30.0 g of $K_2Cr_2O_7$ in a 250 mL volumetric flask.

THINK	WRITE
1. List all required formulas and convert given units to required units.	$c = \frac{n}{V}$ $n = \frac{m}{M}$ $V = \frac{250\text{ mL}}{1000} = 0.25\text{ L}$
2. Calculate moles of $K_2Cr_2O_7$.	$n(K_2Cr_2O_7) = \frac{m}{M}$ $= \frac{30}{(2 \times 39.1 + 2 \times 52 + 7 \times 16)}$ $= 0.102\text{ mol}$
3. Calculate concentration.	$c = \frac{n}{V}$ $= \frac{0.102}{0.250}$ $= 0.408\text{ mol L}^{-1}$

PRACTICE PROBLEM 3

A standard solution of 0.500 mol L^{-1} Na_2CO_3 is required. What mass of anhydrous Na_2CO_3 should be used to prepare the solution in a 250 mL volumetric flask?

Example of preparation of a secondary standard solution

A standard solution of hydrochloric acid is difficult to prepare because the concentrated form of hydrochloric acid produces fumes and dangers are involved in trying to measure out a very accurate volume. For this reason, hydrochloric acid is usually standardised with a known concentration of a suitable base. This involves a titration and allows accurate concentrations of hydrochloric acid to be determined. A standard solution of sodium carbonate is sometimes used as the base because its preparation is relatively easy.

SAMPLE PROBLEM 4

1.33 g of anhydrous sodium carbonate is placed in a 250.0 mL (0.25 L) volumetric flask and water is added to the required mark.

This standard solution is poured into the burette and titrated against a solution of hydrochloric acid; its accurate concentration is to be determined. It was found that a 20.00 mL aliquot of HCl required a Na_2CO_3 titre of 23.50 mL for complete neutralisation.

Calculate the concentrations of:

a. **the standard solution of Na_2CO_3**

b. **the HCl solution.**

Teacher-led video: SP4 (tlvd-0611)

THINK	**WRITE**
a. To calculate the concentration of the standard solution use $n = \frac{m}{M}$ and $c = \frac{n}{V}$.	$n(Na_2CO_3) = \frac{m}{M}$ $= \frac{1.33}{106.0}$ $= 0.01\,255$ mol $c(Na_2CO_3) = \frac{n}{V}$ $= \frac{0.01\,255}{0.2500}$ $= 0.0502$ mol L^{-1}
b. **1.** Write the equation for the reaction and use it to find the molar ratio between Na_2CO_3 and HCl.	$Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$ Na_2CO_3: HCl 1 : 2
2. Calculate the concentration of the solution.	$n(Na_2CO_3) = cV$ $= 0.0502 \times 0.023\,50$ $= 1.18 \times 10^{-3}$ $n(HCl) = 2 \times 1.18 \times 10$ mol L^{-1} $= 2.36 \times 10^{-3}$ mol $c(HCl) = \frac{n}{V}$ $= \frac{2.36 \times 10^{-3}}{0.020\,00}$ $= 0.118$ mol L^{-1} Therefore, the hydrochloric acid has a molarity of 0.118 mol L^{-1}.

PRACTICE PROBLEM 4

A solution of previously standardised sodium hydroxide was used to accurately determine the concentration of a sulfuric acid solution. An average titre of 18.56 mL was required to neutralise 20.00 mL of 0.0927 mol L^{-1} sodium hydroxide solution.

Calculate the molarity of the sulfuric acid solution.

19.4.8 Titration curves

As a titration is performed, a **titration curve** can be constructed to plot the pH change versus the volume of the titre added during titration. The titration curve shows the pH change over time, as the equivalent point is reached or exceeded. The titration curves will have different shapes depending on the strength of the acids and bases used in titration, as each combination will have a different equivalence point. Therefore, identifying the strength of the acids and bases used in the titration will assist in choosing the correct indicator to be suitable for the expected equivalent point.

Figure 19.15 shows an example of a titration of a strong base (NaOH) titrated with a strong acid (HCl) and figure 19.16 shows an example of a strong acid (HCl) titrated with a strong base (NaOH).

In both cases the equivalent point is around a pH = 7, therefore an appropriate indicator to capture the pH change is required. In this example, any indicator with a colour change between pH 4 and pH 8 will be suitable. From table 19.1 we can see that the best indicators for this example would be either bromothymol blue (colour change between 6.0 and 7.6) or litmus (colour change between 5.0 and 9.0).

FIGURE 19.15 Titration curve of a strong base (NaOH) titrated with a strong acid (HCl)

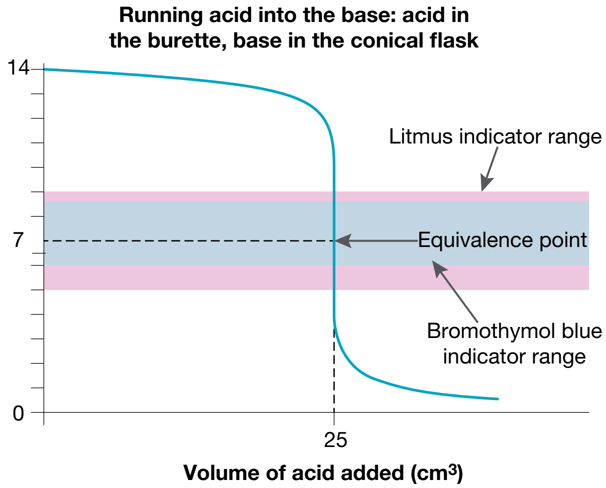
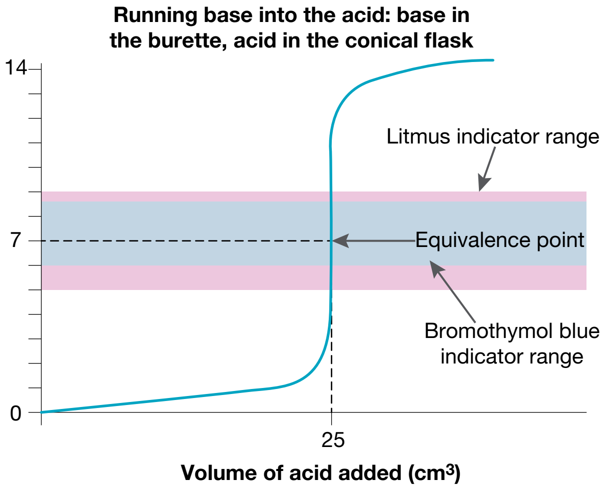


FIGURE 19.16 Titration curves of a strong acid (HCl) titrated with a strong base (NaOH)



19.4.9 Calculations in acid–base titrations

Before starting titration calculations, identify the standard solution information provided. Sometimes calculations for standard solution also need to be employed.

Calculation steps required in a titration:

1. **Write the balanced chemical equation.**
2. **Determine the average titre (concordant volumes only).**
3. **Calculate moles of standard solution from the average titre.**
4. **Use molar ratio from equation to derive the moles of the unknown substance.**
5. **Calculate the concentration of the unknown solution from the calculated moles.**

SAMPLE PROBLEM 5

Calculate the concentration of the NaOH solution if titres of 24.45 mL, 24.78 mL, 24.69 mL and 24.65 mL of 0.300 mol L^{-1} HCl were used to titrate 20.0 mL of NaOH solution.

Teacher-led video: SP5 (tlvd-0612)

THINK	**WRITE**
1. Write the balanced chemical equation.	$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$
2. Determine the average titre (concordant volumes only) for the standard solution and convert volume units from mL to L.	$V = \frac{24.78 + 24.69 + 24.65}{3}$ $= 24.71\text{ mL}$ $= \frac{24.71\text{ mL}}{1000}$ $= 0.0247\text{ L}$
3. Calculate moles of standard solution from the average titre by applying the formula $c = \frac{n}{V}$.	$n = cV$ $n(HCl) = 0.300 \times 0.0247$ $= 0.00741\text{ mol}$
4. Use molar ratio from equation to derive the moles of the unknown substance.	$n(NaOH) : n(HCl) = 1:1$ Thus, $n(NaOH) = 0.00741\text{ mol}$
5. Calculate the concentration of NaOH solution from the calculated moles. **TIP**: Remember to give your answer to the correct number of significant figures.	$c(NaOH) = \frac{n}{V}$ $= \frac{0.00741}{0.0200}$ $= 0.371\text{ mol L}^{-1}$

PRACTICE PROBLEM 5

A standard solution of 0.200 mol L^{-1} K_2CO_3 was used to titrate 23.50 mL of a solution of HCl. The titres used were 20.03 mL, 19. 95 mL, 20.01 mL and 19.98 mL. Calculate the concentration of the aliquot.

19.4.10 Titrations with dilutions

Sometimes when performing titrations with concentrated solutions, we have to use large amount of titres. To minimise the volume used, we can dilute the aliquot to a workable volume.

When dilution is undertaken, we need to remember to account for it in calculations.

SAMPLE PROBLEM 6

A commercial bathroom cleaner contains NaOH as its active ingredient. A 40.0 mL volume of bathroom cleaner was diluted to 250 mL in a volumetric flask.

An aliquot of 20.0 mL of the diluted cleaner was then titrated with 30.0 mL 1.00 mol L^{-1} HCl standard solution. What was the concentration of NaOH in the bathroom cleaner?

Teacher-led video: SP6 (tlvd-0613)

▶

THINK	WRITE
1. Write the balanced equation.	$NaOH(aq) + HCl(aq) \rightarrow NaCl\ (aq) + H_2O(l)$
2. Calculate the moles of standard solution.	$n(HCl) = cV$ $= 1.00 \times 0.0300$ $= 0.0300\ mol$
3. Use molar ratio to calculate moles of diluted aliquot.	$n(HCl) : n(NaOH) = 1:1$ $n(NaOH\ diluted) = 0.0300\ mol$
4. Calculate the concentration of the diluted aliquot.	$c(NaOH\ diluted) = \frac{n}{V}$ $= \frac{0.030}{0.0200}$ $= 1.50\ mol\ L^{-1}$
5. Multiply by the dilution factor to calculate the NaOH concentration in the bathroom cleaner.	Dilution factor: $\frac{250}{40.0} = 6.25$ $c(NaOH\ in\ cleaner) = c(NaOH\ diluted) \times dilution\ factor$ $= 1.50 \times 6.25$ $= 9.38\ mol\ L^{-1}$
Alternatively Use the relationship between the concentration (c) and volume (V) of the initial (or concentrated) solution and the final (or dilute) solution, $c_1V_1 = c_2V_2$.	$c_1\ (concentrated) = ?$ $V_1\ (concentrated) = 40.0\ mL$ $c_2\ (diluted) = 1.50\ mol\ L^{-1}$ $V_2\ (diluted) = 250\ mL$ $c_1V_1 = c_2V_2$ $c_1 = \frac{c_2V_2}{V_1}$ $= 1.50 \times \frac{250}{40}$ $= 9.38\ mol\ L^{-1}$

PRACTICE PROBLEM 6

A commercial bathroom cleaner contains NaOH as active ingredient. A 50.0 mL volume of bathroom cleaner was diluted to 250 mL in a volumetric flask.

An aliquot of 15.0 mL of the diluted cleaner was then titrated with 25.0 mL 0.800 mol L^{-1} HCl standard solution. What was the concentration of NaOH in the bathroom cleaner?

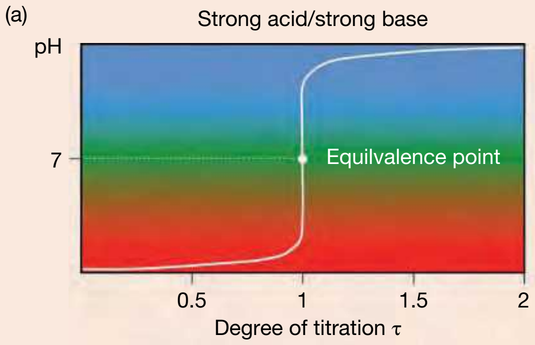
on Resources

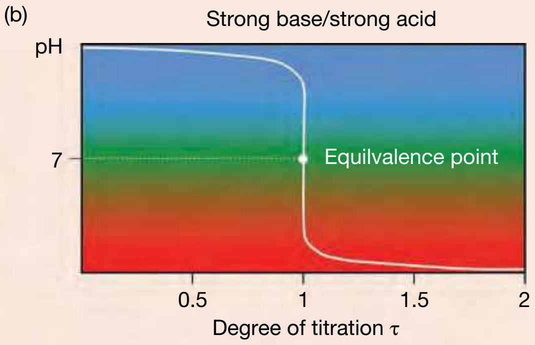
Digital document Experiment 19.1 Carbon dioxide content of fizzy drinks (doc-30856)
Teacher-led video Experiment 19.1 Carbon dioxide content of fizzy drinks (tlvd-0642)

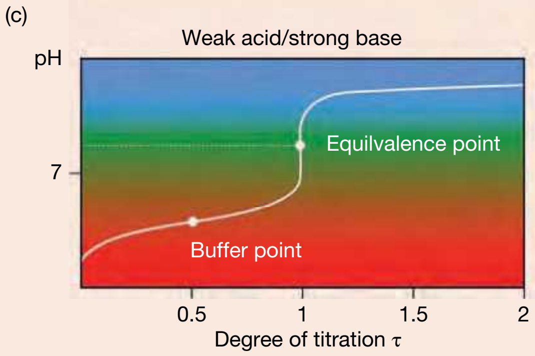
19.4 EXERCISE

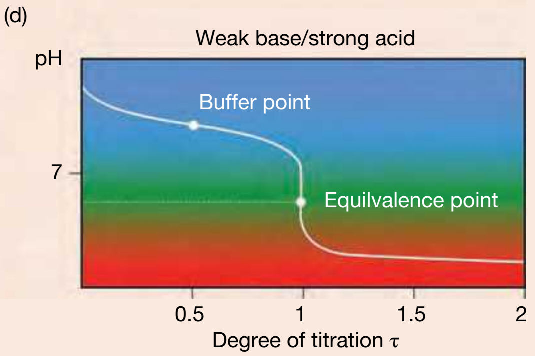
To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Sodium oxalate, $Na_2C_2O_4$, can be used to make a standard solution. Describe how this is done.
2. If 3.183 g of sodium oxalate is dissolved in water in a 250.0 mL volumetric flask, find the concentration of the solution.
3. 1.461 g of dry sodium carbonate, Na_2CO_3, is dissolved in water in a 250 mL volumetric flask.
   (a) Why can this be called a standard solution?
   (b) Calculate the concentration of the solution.
4. 20.00 mL aliquots of the solution from question 3 were titrated with nitric acid, HNO_3, three times. The average concordant titre was found to be 22.17 mL.
   (a) Write the equation for the reaction of sodium carbonate with nitric acid.
   (b) Find the concentration of the nitric acid.
5. Describe the difference between equivalence point and end point.
6. What indicator would be suitable to be used in a titration where equivalence point is at pH 4.1?
7. What indicators would be most suitable for the following titration combinations given that the general titration curves are known?









8. What volume of 0.80 mol L^{-1} HCl will be required to titrate a 20.0 mL solution of 0.60 mol L^{-1} KOH to the equivalence point?
9. A factory is suspected of polluting a stream with hydrochloric acid from its waste water. To gather evidence, a sample of stream water was collected downstream from the discharge point. Back in the laboratory, a 20.00 mL sample of this water was titrated against a solution of sodium carbonate that had been accurately diluted to 0.0010 mol L^{-1}; 17.52 mL was required. Assuming that the only acid in the sample was hydrochloric acid:
   (a) calculate the molarity of hydrochloric acid in the sample
   (b) calculate the concentration of the hydrochloric acid in g L^{-1}, mg L^{-1} and ppm.
10. Ten 1 L water samples were collected from different locations in a lake and then mixed together. One litre of this mixture was then titrated with 0.050 mol L^{-1} sulfuric acid until the pH equalled 4.5. 15 mL of acid was required.
    Calculate the total alkalinity of this water as mg L^{-1} calcium carbonate equivalent.

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19.5 Review

19.5.1 Summary

Sources of acids and bases in water

- Water can have a wide pH range and still be considered normal. This is due to substances that may dissolve in it.
- Water can be made acidic or basic by the substances it comes into contact with which either accept or donate protons, hence changing the pH of the water. These include rocks and minerals, the atmosphere, natural sources, dissolved salts and the results of human activity, including contamination.
- Alkalinity is a measure of the ability of water to withstand sudden pH change caused by acid input.

Solution stoichiometry (volume–volume stoichiometry)

- The principles of stoichiometry can be applied to reactions in solution (concentration–concentration stoichiometry). The formula $n = cV$ is used here.
- When solutions are diluted, the moles are not changed, $n_1 = n_2$, and the formula $c_1V_1 = c_2V_2$ is used to calculate required quantities and results.

Acid–base titrations

- Acid–base titrations are a form of volumetric analysis that use concentration–concentration stoichiometry. They are used to calculate the unknown concentration of an acid or a base.
- Titrations use specialised equipment such as burettes, pipettes and volumetric flasks. Accurate techniques must be used to produce reliable results.
- Acid–base titrations require the use of correctly chosen indicators. These are substances that have different colours at different pH values.
- The equivalence point of a titration is the point at which chemically equivalent amounts of acid and base are present according to the equation.
- The end point of a titration is the point at which the indicator changes colour. The indicator should be chosen so that the end point is close to the equivalence point.
- A standard solution is a solution of accurately known concentration.
- A primary standard is a substance that can be weighed out and dissolved to make a solution, and its concentration is known accurately.

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To access key concept summaries and practice exam questions download and print the **studyON: Revision and practice exam question booklet** (doc-30943).

19.5.2 Key terms

aliquot volume delivered by a pipette
alkalinity a measure of the ability of water to withstand sudden pH change caused by acid input
amphiprotic a substance that can donate or accept protons
average titre titre determined by averaging the results of several titrations
buffers solution that resists a change in pH upon the addition of a small amount of acid or base
concordant individual titres within a certain limit
concordant results results of repeated titrations that are identical (within ± 0.10 mL of each other) and consecutive
end point experimentally determined equivalence point at which the indicator just changes colour or the pH curve becomes vertical
equivalence point point at which two reactants have reacted in their correct mole proportions in a titration
primary standard substance used in volumetric analysis that is of such high purity and stability that it can be used to prepare a solution of accurately known concentration
secondary standard a standard prepared for a specific volumetric analysis that has been standardised against a primary standard
standard solutions solution that has a precisely known concentration
synergy to have an effect that is greater than the individual effects combined
titration process used to determine the concentration of a substance by using a pipette to deliver one substance and a burette to deliver another substance until they have reacted in exactly their mole ratio
titration curve results of a titration showing the pH change versus the volume of the titre added during titration
titre volume of liquid measured by a burette during titration
volumetric analysis measurement of volume of one solution of accurately known concentration to determine the concentration of a second solution using stoichiometry

Resources

Digital document Key terms glossary — Topic 19 (doc-30941)

19.5.3 Practical investigations

Experiment 19.1

Carbon dioxide content of fizzy drinks

Aim: To determine the CO_2 content of different brands of soft drinks

Digital document: doc-30856
Teacher-led video: tlvd-0642

Resources

Digital document Practical investigation logbook (doc-30942)

19.5 Exercises

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

19.5 Exercise 1: Multiple choice questions

1. Which of the following statements about water is *not* true?
   - **A.** Sea water with a pH of 8 would be considered to have been acidified.
   - **B.** Ground water with a pH of 6 would be considered to be normal.
   - **C.** Surface water with a pH of 6 would be considered to have been acidified.
   - **D.** Surface water with a pH of 5 would probably contain a lot of dissolved calcium carbonate.
2. Which of the following acids is least likely to contribute to broadscale acid rain?
   - **A.** H_2CO_3
   - **B.** HCl
   - **C.** H_2SO_4
   - **D.** HNO_3
3. A student dissolved 2.00 g of sodium hydroxide in 1.00 L to produce solution A. The student then diluted 100 mL of solution A with water to produce 1.00 L of solution B. A 10.0 mL sample of solution B was then diluted to 100 mL with water to produce solution C.
   What was the concentration of solution C?
   - **A.** 0.500 mol L^{-1}
   - **B.** 0.0500 mol L^{-1}
   - **C.** 0.005 00 mol L^{-1}
   - **D.** 0.000 500 mol L^{-1}
4. What volume of 11 mol L^{-1} hydrochloric acid is required to produce 2 L of 1 mol L^{-1} acid when diluted with water?
   - **A.** 91 mL
   - **B.** 182 mL
   - **C.** 600 mL
   - **D.** 1838 mL
5. Which of the following is a *not* requirement for a substance to be a primary standard?
   - **A.** It must have a high state of purity.
   - **B.** It must not be a danger to health.
   - **C.** It must have an accurately known formula.
   - **D.** It must be stable.
6. When is the end point reached in an acid–base titration?
   - **A.** When the solution is neutral.
   - **B.** When the indicator changes colour.
   - **C.** When equal volumes of reactants have been mixed.
   - **D.** When reactants have been mixed in the appropriate stoichiometric ratio.
7. The equivalence point in an acid–base titration is the point when
   - **A.** the solution is neutral.
   - **B.** the indicator changes colour.
   - **C.** equal volumes of reactants have been mixed.
   - **D.** reactants have been mixed in the appropriate stoichiometric ratio.
8. What mass of potassium hydroxide, KOH, is needed to produce 200 mL of a 0.25 mol L^{-1} solution?
   - **A.** 0.050 g
   - **B.** 2.8 g
   - **C.** 50 g
   - **D.** 2.8 kg

9. What volume of 0.100 mol L^{-1} H_2SO_4 is required to react completely with 25.0 mL of 0.150 mol L^{-1} NaOH solution?
   - A. 9.38 mL
   - B. 18.8 mL
   - C. 37.5 mL
   - D. 75.0 mL
10. A student obtained the following titres during an analysis: 18.90 mL, 19.02 mL, 18.97 mL, 19.12 mL, 18.98 mL.
    What is the average of the concordant titres?
    - A. 18.95 mL
    - B. 18.99 mL
    - C. 19.01 mL
    - D. 19.07 mL

19.5 Exercise 2: Short answer questions

1. Normal rainfall has a pH of about 5.6. However, the lowest pH at which surface water can be considered normal is 6.5.
   - a. Explain how rainwater becomes acidic.
   - b. What must happen to rainwater after it hits the ground to change its pH?
   - c. Rainfall in some parts of the world has been measured at a pH of 4.3. What is the term given to such rainfall?
   - d. Explain how the situation in part c might come about.
2. Using the internet or other suitable reference, find three naturally occurring rocks or minerals that would make water
   - a. acidic.
   - b. basic.
3. Calculate the volume of stock solution needed to prepare the following.
   - a. 500 mL of a 0.750 mol L^{-1} solution of sulfuric acid (sulfuric acid stock solution = 18 mol L^{-1})
   - b. 200 mL of a 2.5 mol L^{-1} solution of ammonium hydroxide (ammonium hydroxide stock solution = 15 mol L^{-1})
   - c. 350 mL of a 0.150 mol L^{-1} solution of ethanoic acid (ethanoic acid stock solution = 17 mol L^{-1})
4. What volume of water must be added to
   - a. 100 mL of 15 mol L^{-1} H_2SO_4 to dilute it to 5.0 mol L^{-1}?
   - b. 130 mL of 3.50 mol L^{-1} HNO_3 to dilute it to 1.00 mol L^{-1}?
   - c. 170 mL of 2.60 mol L^{-1} HCL to dilute it to 0.250 mol L^{-1}?
5. A student is required to make up 500 mL of 0.100 mol L^{-1} H_2SO_4(aq) by diluting a 0.500 mol L^{-1} solution of the acid. What volume of 0.500 mol L^{-1} H_2SO_4(aq) is required?
6. A 3.50 g mass of $MgCl_2$ is dissolved in 200 mL of water.
   - a. What is the concentration of the solution?
   - b. If 150 mL of water is added to the solution, what is the concentration of the new, diluted solution?
7. Sodium chloride reacts with silver nitrate to give a white precipitate of silver chloride.
   - a. Write an equation for the reaction.
   - b. What volume of 2.0 mol L^{-1} silver nitrate reacts with 120 mL of 1.5 mol L^{-1} sodium chloride?
8. 50 mL of 1.2 mol L^{-1} sulfuric acid was added to 30 mL of sodium hydroxide in order to neutralise it. Find the molarity of the sodium hydroxide.
9. What volume of a 2.30 mol L^{-1} solution of Na_3PO_4 is required to react completely with 560 mL of a 3.25 mol L^{-1} solution of $Ca(ClO_3)_2$, according to the reaction:

$$3Ca\,(ClO_3)_2(aq) + 2Na_3PO_4(aq) \rightarrow Ca_3\,(PO_4)_2(aq) + 6NaClO_3(aq)$$

10. What volume of 0.250 mol L^{-1} HNO_3 reacts completely with 280 mL of 0.200 mol L^{-1} H_2S according to the reaction:

$$2HNO_3(aq) + 3H_2S(aq) \rightarrow 3S(s) + 4H_2O(I) + 2NO(g)$$

11. What volume of 0.460 mol L^{-1} H_2SO_4 is required to neutralise 24.00 mL of 0.620 mol L^{-1} NaOH?
12. A sample of water known to contain sodium hydroxide as the only base was analysed prior to discharge into a river. It was found that a 20.00 mL sample required 8.19 mL of 0.0100 mol L^{-1} hydrochloric acid solution for neutralisation.
    a. Calculate the molar concentration of sodium hydroxide in the water sample.
    b. Express the answer from part a in g L^{-1}.

19.5 Exercise 3: Exam practice questions

Question 1 (6 marks)

Anhydrous sodium carbonate is often used to make standard solutions of sodium carbonate. This solution may then be used in acid–base titrations to calculate the exact concentration of acidic solutions.

a. Calculate the mass of anhydrous sodium carbonate, Na_2CO_3, that is required to make 500.00 mL of 0.0500 mol L^{-1} solution. **2 marks**

b. Exactly 25.00 mL of the solution from part (a) is reacted with hydrochloric acid. The equation for this reaction is:

$$2HCI(aq) + Na_2CO_3(aq) \rightarrow 2NaCI(aq) + CO_2(g) + H_2O(I)$$

It is found that 23.50 mL of the acid is required for exact neutralisation.

Calculate the molarity of the hydrochloric acid solution. **4 marks**

Question 2 (11 marks)

A company operates a batch process that uses sulfuric acid. At the end of each batch, the leftover acid must be diluted before it is released. The amount of water to be added to enable this dilution is carefully calculated following analysis of the leftover sulfuric acid in the batch liquid.

The terms of the discharge licence state that no more than 1.00 g L^{-1} of sulfuric acid should be in the discharge water.

A 25.00 mL sample of the batch liquid was accurately diluted to 250.0 mL in a standard flask. 25.00 mL of this diluted sample was then titrated with 0.100 mol L^{-1} standardised sodium hydroxide solution using a suitable indicator. An average titre of 18.52 mL was obtained.

a. Write the equation for the reaction between sodium hydroxide and sulfuric acid. **1 mark**
b. Calculate the number of moles of sodium hydroxide that were added. **1 mark**
c. From your part b answer, calculate the number of moles of sulfuric acid in the diluted sample. **2 marks**
d. Hence, calculate the molarity of sulfuric acid in the diluted sample. **1 mark**
e. Calculate the molarity of sulfuric acid in the undiluted batch liquid. **2 marks**
f. Calculate the molarity of the sulfuric acid that is allowed in the discharge water. **1 mark**
g. Hence, calculate the number of litres of water that must be added to each litre of batch liquid to meet the terms of the discharge licence. Give your answer to one decimal place. **3 marks**

19.5 Exercise 4: studyON Topic Test online only

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UNIT 2 | AREA OF STUDY 2 REVIEW

AREA OF STUDY 2 How are substances in water measured and analysed?

OUTCOME 2

Measure amounts of dissolved substances in water, and analyse water samples for salts, organic compounds and acids and bases.

PRACTICE EXAMINATION

STRUCTURE OF PRACTICE EXAMINATION

Section	Number of questions	Number of marks
A	20	20
B	8	30
	Total	**50**

Duration: 50 minutes

Information:

- This examination consists of two parts, you must answer all question sections
- Pens, pencils, highlighters, erasers, rulers and a scientific calculator are permitted.
- You may use the VCE Chemistry Data Book for this task.

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Weblink VCE Chemistry Data Book

SECTION A

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

1. One of the ways of measuring salinity makes use of conductivity measurements. These can detect the overall concentration of ions but not the particular ions present. Which of the following solutions would be expected to have the highest conductivity?

A. 0.005 mol L^{-1} $CaSO_4(aq)$
B. 0.005 mol L^{-1} $AlCl_3(aq)$
C. 0.008 mol L^{-1} NaCl(aq)
D. 0.006 mol L^{-1} $MgCl_2(aq)$

2. According to the label on a bottle of mineral water, the concentration of the hydrogen carbonate (or bicarbonate) ion is 248 mg L^{-1}. If a 350 mL bottle of this water was consumed, what would the mass of HCO_3^- ingested in grams be?

A. 0.0868
B. 0.141
C. 1.41
D. 86.8

3. Carbon reacts with highly concentrated nitric acid, producing the gas nitrogen monoxide and other gases. The equation for the reaction is:

$$3C(s) + 4HNO_3(l) \rightarrow 3CO_2(g) + 4NO(g) + 2H_2O(g)$$

If 0.80 mol of nitrogen monoxide, NO, is produced in the reaction

A. 3.0 mol of carbon is consumed and 3.0 mol of carbon dioxide is formed.
B. 0.60 mol of nitric acid is consumed and 0.60 mol of carbon dioxide is formed.
C. 0.60 mol of carbon is consumed and 0.60 mol of carbon dioxide is formed.
D. 0.80 mol of nitric acid is consumed and 0.80 mol of carbon dioxide is formed.

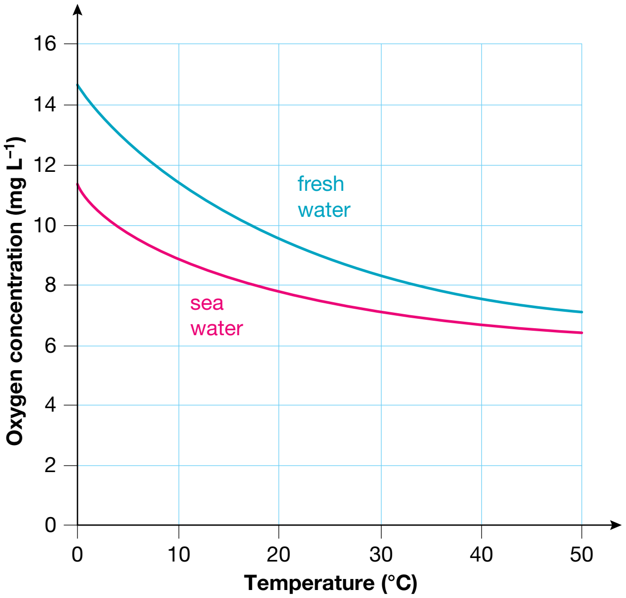
4. The noble gas xenon can be converted to a fluoride compound, xenon tetrafluoride, XeF_4, due to the highly reactive nature of fluorine gas, F_2. It is formed according to the following equation:

$$Xe(g) + 2F_2(g) \rightarrow XeF_4(g)$$

What will happen if a mixture of 100 g of Xe and 200 of F_2 is allowed to react?

A. There will be approximately 100 grams of fluorine in excess.
B. There will be approximately 140 grams of fluorine in excess.
C. There will be approximately 50 grams of xenon in excess.
D. Neither one will be in excess because the masses present are in the same ratio as the equation.

5. The following graph shows the concentration of dissolved O_2 in both fresh and sea water.



According to the graph, at 15 °C, what would the mass of dissolved oxygen in 250 mL of fresh water be closest to?

A. 0.0021 g
B. 0.0026 g
C. 2.1 g
D. 2.6 g

6. Which of the following statements correctly describes a saturated solution of a solid?

I. Adding more solute results in undissolved material at the base of the solution.
II. Doubling the temperature of the solution will allow twice as much solute to be dissolved.
III. Lowering the temperature of the solution causes the formation of crystals.

A. Only statements (I) and (II) are correct.
B. Only statements (II) and (III) are correct.
C. Only statements (I) and (III) are correct.
D. Only statement (III) is correct.

7. A popular alcoholic beverage lists its alcohol content as 35.0% (v/v). The mass of ethanol present in one 'standard drink' is 10.0 g, which corresponds to a volume of 8.90 mL of pure ethanol. What volume of the alcoholic beverage would a person need to consume to reach the alcohol content of a standard drink?

A. 8.9 mL
B. 25 mL
C. 29 mL
D. 35 mL

8. The following table shows the allowable level of heavy metal contaminants in drinking water sources.

Metal	Allowable level (mg L^{-1})
Arsenic	0.010
Chromium	0.050
Cadmium	0.002
Lead	10

A 100 mL sample of drinking water was analysed by atomic absorption spectroscopy and found to contain 0.0010 mg of cadmium and 0.0002 mg of arsenic. What do these results mean?

A. The sample is above the allowable limit for both arsenic and cadmium.
B. The sample is above the allowable limit for arsenic but below the limit for cadmium.
C. The sample is above the allowable limit for cadmium but below the limit for arsenic.
D. The sample is below the allowable limit for both arsenic and cadmium.

9. Which of the following substances is NOT an *organometallic* contaminant?

A. Methyl iodide
B. Tetraethyl lead
C. Tributyl tin
D. Methyl mercury

10. A student weighed out 2.396 g of anhydrous sodium carbonate and made this up to the mark in a 250.0 mL volumetric flask. What was the molarity of the sodium carbonate solution?

A. 0.000 090 42 mol L^{-1}
B. 0.090 42 mol L^{-1}
C. 0.011 55 mol L^{-1}
D. 0.5990 mol L^{-1}

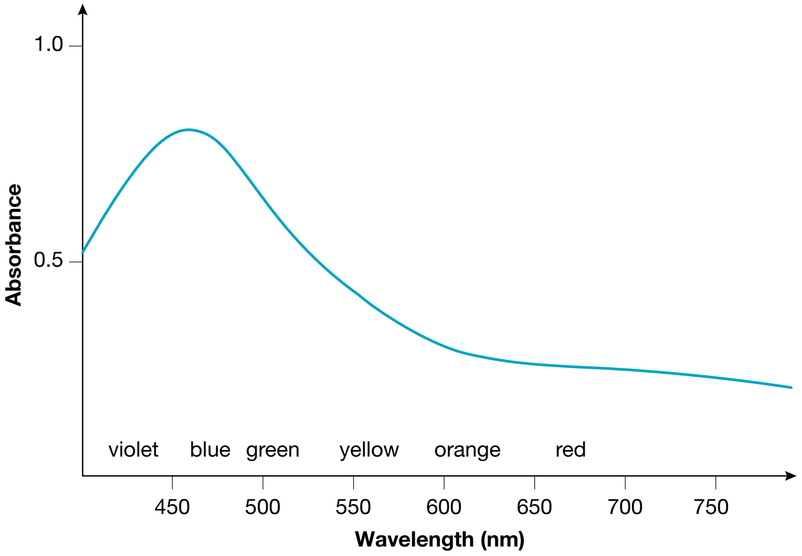
11. Which of the following techniques is NOT an example of quantitative analysis?

A. Using atomic absorption spectroscopy to analyse the level of metal contaminants in streams.
B. Using volumetric analysis to determine the concentration of an acid.
C. Using gravimetric analysis to find the percentage of water in foods.
D. Using chromatography to establish if a contaminant has been added to an oil sample.

12. Which of the following sucrose solutions has the highest amount of solute?

A. 10 mL of a 0.40 mol L^{-1} solution
B. 120 mL of a 0.010 mol L^{-1} solution
C. 40 mL of a 0.12 mol L^{-1} solution
D. 25 L of a 1.00×10^{-4} mol L^{-1} solution

13. The UV-visible absorption spectrum of a transition metal solution is shown in the following figure. The colours corresponding to the wavelengths are also shown.



Which of the following statements is correct?

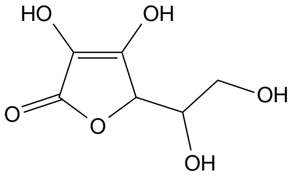
A. The solution is blue and blue light should be used to measure absorbance.
B. The solution is orange and orange light should be used to measure absorbance.
C. The solution is blue and orange light should be used to measure absorbance.
D. The solution is orange and blue light should be used to measure absorbance.

14. A beaker contains 50 mL of 0.15 mol L^{-1} NaCl solution. What volume of water must be added to this solution to change its concentration to 0.050 mol L^{-1}?

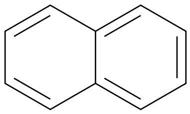
A. 20 mL
B. 50 mL
C. 100 mL
D. 150 mL

15. Which of these organic compounds would be expected to dissolve well in water?

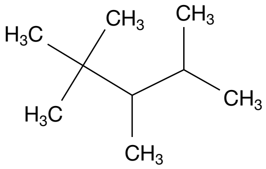
A.



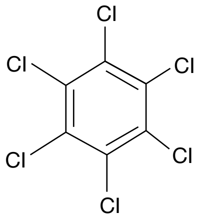
B.



C.



D.



A. Compound A
B. Compound B
C. Compound C
D. Compound D

16. A student needs to prepare a standard solution of anhydrous sodium carbonate for a volumetric analysis. Which set of equipment should be used for this purpose?

A. Electronic balance, measuring cylinder, beaker and spatula

B. Volumetric flask, burette, spatula and pipette

C. Electronic balance, pipette, funnel and spatula

D. Volumetric flask, funnel, spatula and electronic balance

17. Chromatography is a technique used to analyse organic substances. Consider the following statements about chromatography.

I. The more polar a substance, the faster it travels through a chromatograph.

II. The retention time of a component depends on its relative attraction to the mobile phase only.

III. The more attracted a component is to the stationary phase, the faster it will exit the column.

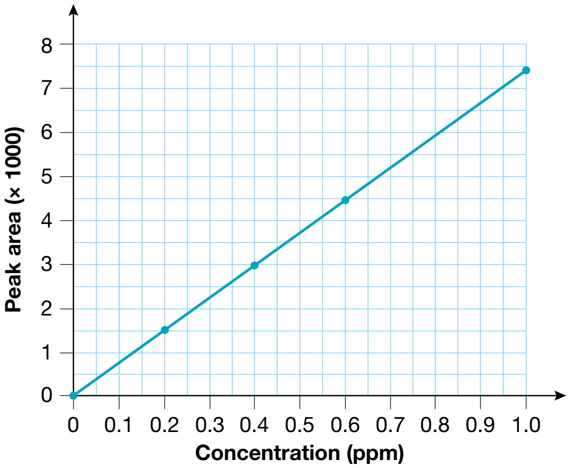
A. Only statement (I) is correct.

B. Only statements (II) and (III) are correct.

C. All three statements are correct.

D. None of the statements is correct.

18. Endrin is an insecticide that is highly toxic to aquatic animals and humans. The levels of Endrin in dam water were tested using high-performance liquid chromatography (HPLC). The following calibration curve was produced for its analysis.



The sample of dam water gave a peak area of 5000 for Endrin.
What mass of Endrin would be present in 500 mL of this water?

A. 0.68 mg

B. 0.34 mg

C. 1.36 mg

D. 5000 mg

19. When comparing rain water collected in an industrial area with rain falling in a national park, what does the rain in the industrial area usually have?

A. Higher pH due to smoke and soot particles

B. Lower pH due to the dissolved oxides of nitrogen and sulfur

C. Higher pH due to increased levels of CO_2

D. Lower pH due to acidic waste released by factories into the waterways

20. A solution of nitrate ions was prepared by mixing 100 mL of 1.0 mol L^{-1} sodium nitrate solution with 100 mL of 1.0 mol L^{-1} aluminium nitrate solution. What was the resulting nitrate ion concentration?

A. 4.0 mol L^{-1}

B. 3.0 mol L^{-1}

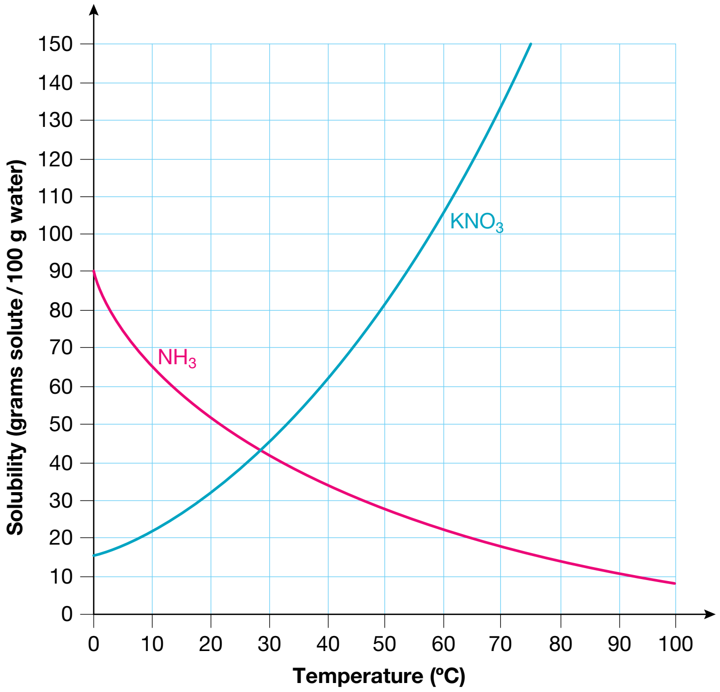
C. 2.0 mol L^{-1}

D. 1.0 mol L^{-1}

SECTION B

Question 1 (3 marks)

1. The solubility curves for ammonia and potassium nitrate are shown in the following figure.



According to these curves:

a. What would be the maximum amount of ammonia that would dissolve in 100 g H_2O at 65 °C? **1 mark**

b. At what minimum temperature would 46 g of KNO_3 dissolve in 50 g of water? **1 mark**

c. A saturated solution of potassium nitrate was prepared in 200 g water at 40 °C. The solution was then cooled to 0 °C. What mass of solute would crystallise? **1 mark**

Question 2 (4 marks)

2. Express the following solutions in the concentration unit given.

a. A 375 mL can of soft drink contains 4.05 g of sucrose. What is the sucrose concentration in g L^{-1}? **1 mark**

b. The National Health and Medical Research Council (NHMRC) of Australia specifies a maximum level of 1.5 milligrams per litre for fluoride in drinking water.

  i. What mass of fluoride, in grams, would be present in a 250 mL glass of this water? **1 mark**

  ii. What would be the molar concentration of fluoride at this level of concentration? **1 mark**

  iii. What would this concentration be in ppm? **1 mark**

Question 3 (4 marks)

3. Iron is produced by reducing iron oxides such as haematite, Fe_2O_3, in a blast furnace. This multi-step process has an overall equation:

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$$

What mass of iron is formed from 1.00 tonne of Fe_2O_3? (*Note:* 1 tonne = 1 Mg.) Express your answer in kg to the correct number of significant figures.

Question 4 (7 marks)

4. Compounds of zinc can cause water to appear turbid (cloudy) in appearance and they also give the water an unpleasant taste. The concentration of Zn^{2+} ions in some turbid water was analysed gravimetrically by precipitating the zinc ions as zinc phosphate using sodium phosphate to form the precipitate.
When excess sodium phosphate solution was added to 250.0 mL of the zinc-contaminated water, 1.844 g of precipitate was formed.

a. Write the *ionic* equation for the formation of the zinc carbonate precipitate from the zinc ions and phosphate ions. Include symbols of state. **1 mark**

b. i. Calculate the amount in mol of zinc phosphate precipitated. **1 mark**

ii. Calculate the molar concentration of Zn^{2+} ions present in the water sample. **2 marks**

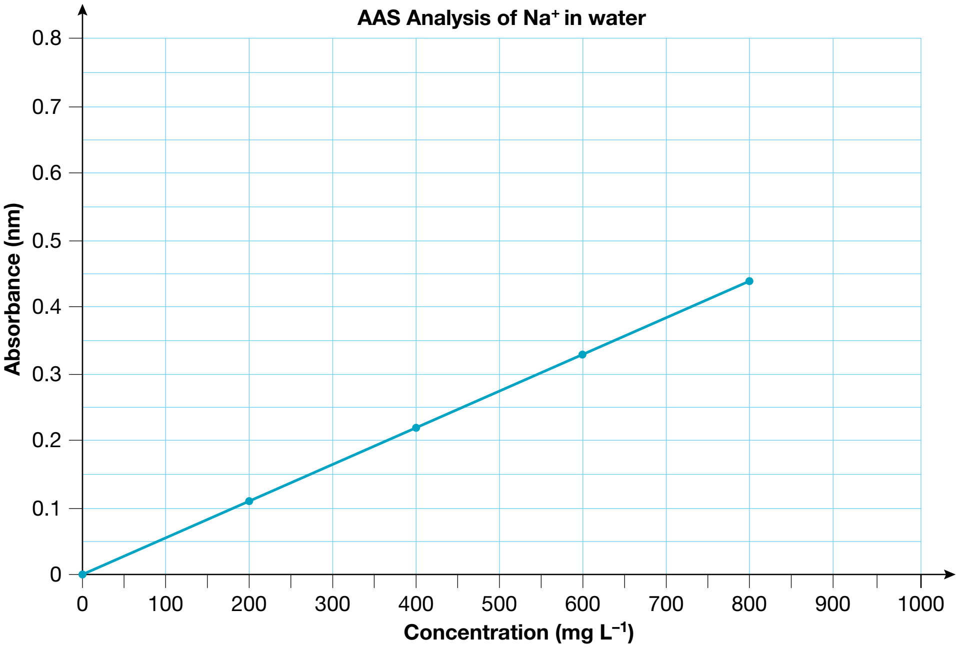
c. What experimental technique should be used to ensure that the precipitate is fully dried before weighing? **1 mark**

d. i. What is one assumption made about the precipitate formed? **1 mark**

ii. What would have been the effect on the calculated Zn^{2+} concentration if 10 mL more of the sodium phosphate solution had been added? Explain. **1 mark**

Question 5 (4 marks)

5. An analysis was made of sodium ions present in the stream water running through a region known to be affected by salinity. Atomic absorption spectroscopy (AAS) was used for the analysis. An initial test of the water in the stream gave a reading that was off scale and so the samples to be tested were diluted to one-tenth of their original concentration before spraying into the AAS. The calibration curve used for the analysis is shown in the following figure.



a. What is the highest concentration (in mg L^{-1}) that can be determined using this calibration graph? **1 mark**

b. The diluted water samples gave an absorbance reading of 0.30.

i. What is the concentration of sodium ions in the stream water measured in mg L^{-1}? **1 mark**

ii. What mass of sodium in kg is present in 3.5 megalitres (ML) of the stream water? **1 mark**

c. What volume of distilled water would have been added to 50.0 mL of the stream water in order to dilute it to one-tenth of its original concentration? **1 mark**

Question 6 (8 marks)

7. The concentration of a barium hydroxide solution was determined by titrating it with a previously standardised solution of hydrochloric acid. The concentration of the acid was 0.0500 M.
In the titration, the hydrochloric acid was placed in a pre-rinsed burette and 20.00 mL aliquots of the barium hydroxide solution were delivered to a conical flask using a pre-rinsed pipette. The titration was repeated four times and the following titres were obtained: 18.64 mL, 18.15 mL, 18.18 mL and 18.10 mL.

a. With what substance should the conical flask have been rinsed before use? **1 mark**

b. Calculate the average of the concordant titres. **1 mark**

c. Write the equation for the reaction between hydrochloric acid and barium hydroxide. **1 mark**

d. Calculate the concentration of the barium hydroxide solution based on the given data. **3 marks**

e. The titration was carried out using bromothymol blue as an indicator. Bromothymol blue is yellow in acidic solutions and blue in basic solutions. What were the colours of the indicator, in order, seen throughout the titration? **1 mark**

f. If the burette had been rinsed with water, how would this have affected the value of the concentration of barium hydroxide calculated in this analysis? **1 mark**

PRACTICE SCHOOL-ASSESSED COURSEWORK

ASSESSMENT TASK — SCIENTIFIC INQUIRY

This scientific inquiry task focuses on a scenario of possible contamination of tap water. The inquiry begins with an open-inquiry investigation brief. You will then apply your knowledge to answer the issues raised by the brief.

- Pens, pencils, highlighters, erasers, rulers and a scientific calculator are permitted.
- You may use the VCE Chemistry Data Book for this task.

Total time: 55 minutes (5 minutes reading, 50 minutes writing)

Total marks: 50 marks

DRINKABLE WATER?

Introduction

In this task, you will measure amounts of dissolved substances in water and analyse water samples for salts, organic compounds and acids and bases.

Task

Congratulations! You have just been appointed chief scientist at Yarra Valley Water. However, you have received several complaints by members of the public that their tap water is cloudy and has a foul odour.

The complainants are threatening to take Yarra Valley Water to court for producing water unfit for human consumption. Your job is to carry out a full analysis of the water to determine whether the water is unfit for human consumption. You have access to a fully stocked research laboratory with all the latest equipment and technology.

The response requires the following to be completed:

1. Selecting five water quality analysis experiments
2. The significance of each of these experiments
3. Explaining the experimentation method employed
4. The experimental limitations.

on Resources

Digital document U2AOS2 School-assessed coursework (doc-31823)

20 Practical investigation

20.1 Overview

Numerous **videos** and **interactivities** are available just where you need them, at the point of learning, in your digital formats, learnON and eBookPLUS at www.jacplus.com.au.

20.1.1 Introduction

Performing practical investigations and writing clear and concise reports are essential skills in chemistry. As part of Unit 2 in VCE Chemistry, you will conduct a practical investigation into an aspect of water quality. Conducting investigations allows conclusions to be drawn, questions to be investigated and answers formulated. In this topic, you will find guidelines for carrying out practical tasks safely and preparing thorough reports and scientific posters using appropriate scientific conventions. You will learn how to develop a question, plan a methodology and present, analyse and evaluate data obtained using appropriate chemical terminology.

FIGURE 20.1 A student conducting a scientific investigation in the classroom

20.1.2 What you will learn

KEY KNOWLEDGE

In this topic, you will investigate:

- the chemical concepts specific to the investigation and their significance, including definitions of key terms, and chemical representations.,
- the characteristics of laboratory techniques of primary qualitative and quantitative data collection relevant to the investigation: sampling protocols; gravimetric analysis, acid–base titrations and/or pH measurement; precision, accuracy, reliability and validity of data; and minimisation of experimental bias.
- ethics of and concerns with research, including identification and application of relevant health and safety guidelines
- methods of organising, analysing and evaluating primary data to identify patterns and relationships, including identification of sources of error and uncertainty, and of limitations of data and methodologies
- observations and experiments that are consistent with, or challenge, current chemical models or theories
- the nature of evidence that supports or refutes a hypothesis, model or theory
- options, strategies or solutions to issues related to water quality
- the key findings of the selected investigation and their relationship to solubility, concentration, acid–base and/or redox concepts
- the conventions of scientific report writing including chemical terminology and representations, symbols, chemical equations, formulas, units of measurement, significant figures and standard abbreviations.

KEY SCIENCE SKILLS

- Develop aims and questions, formulate hypotheses and make predictions
- Plan and undertake investigations
- Comply with safety and ethical guidelines
- Conduct investigations to collect and record data
- Analyse and evaluate data, methods and scientific models

- Draw evidence-based conclusions
- Communicate and explain scientific ideas

Source: VCE Chemistry Study Design (2016–2021) extracts © VCAA; reproduced by permission.

on Resources

Digital documents Key science skills (doc-30903)
Key terms glossary — Topic 20 (doc-30959)

20.2 Key science skills in chemistry

KEY CONCEPTS

- The chemical concepts specific to the investigation and their significance, including definitions of key terms, and chemical representations
- Methods of organising, analysing and evaluating primary data to identify patterns and relationships, including identification of sources of error and uncertainty, and of limitations of data and methodologies
- Observations and experiments that are consistent with, or challenge, current chemical models or theories

KEY SCIENCE SKILLS

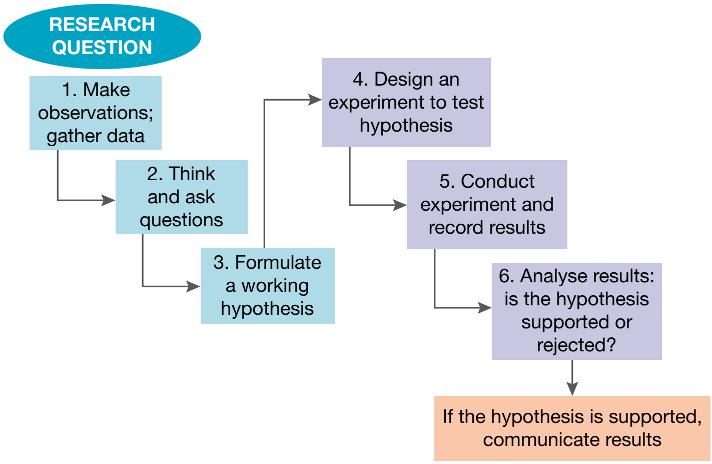
- Develop aims and questions, formulate hypotheses and make predictions
- Plan and undertake investigations
- Comply with safety and ethical guidelines
- Conduct investigations to collect and record data

20.2.1 The scientific method: why do we conduct investigations?

In Unit 2, Outcome 3 you can use your developing experimental skills to investigate a question related to water quality. You will be conducting a practical investigation that uses laboratory or fieldwork to respond to a question. A practical investigation involves considerable planning, and expertise in working scientifically to design a method, record results, analyse findings and communicate conclusions.

The diagram shown in figure 20.2 summarises this process of practical investigations and **scientific method**.

FIGURE 20.2 The scientific method



This investigation will draw upon a number of key skills that you have been developing in Unit 2, Areas of Study 1 and 2, and the specific skills in Unit 2, Area of Study 3. As part of this investigation, you will be required to produce a report. This is usually in the form of a scientific poster, but your teacher may ask for another format such as a practical report, oral presentation or digital presentation. You will also be required to produce a logbook.

In this investigation, you will have the opportunity to show your skill and imagination in experimental design, commitment to a task and your communication ability in explaining your **results**. You will need to develop a question, plan actions to answer this question, undertake an investigation and interpret the data to form a conclusion.

This Unit 2 Outcome 3 task requires between 4 and 6 hours of class time, allowing time to both conduct investigation and communicate findings. Table 20.1 will assist with your planning. The timeline may be different depending on your school and when they conduct the investigation. Your teacher may also have set checkpoints regarding when you are required to submit work and what specific components need to be included.

TABLE 20.1 Investigation planning with sample schedule

Task	Due date
• Introduction of task and expectations from teachers • Brainstorm topics around the idea of water quality • Construct a choice of topics of investigation with your teacher	About two weeks before formal experimentation begins
• Return your list of possible topics for approval by your teacher, who then provides feedback, recommendations and finally approval.	A few days later
• Submit your detailed research proposal for your approved topic.	At the beginning of the week before your experiment begins
• Submit your completed and signed risk assessment (this also must be signed by your teacher).	By the end of the week before your experiment begins
• Your requested equipment is assembled by the teacher and lab technician.	By the end of the week before your experiment begins
Your investigation begins. • Set up equipment • Collect preliminary data • Troubleshoot any issues with equipment • Collect data and measurements, graph your results and evaluate trends • Adjust and refine method	Week 1
• Continue the cycle of measurements and data analysis, leading to a review of progress and further detailed measurements.	Week 2
• Finalise writing the sections of your report and paste them into a poster template (if required). Submit your log book and finished report (and poster if required).	Week 3

20.2.2 Using a log book

As part of your scientific investigation (as well as all practical experiments throughout the year), you are required to keep a logbook.

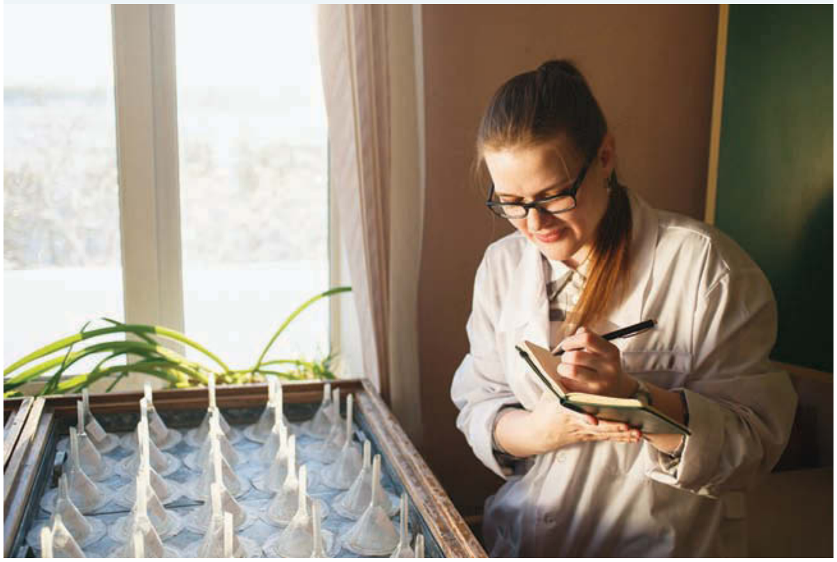
Usually this logbook is a bound exercise book (however, your teacher may request a digital logbook instead). It is vital to show all aspects of your practical investigation within your logbook using the scientific approach.

This logbook will be assessed by your teacher. You must date all work in your logbook to show when it was completed and assist in validating your work.

TABLE 20.2 Components of a logbook

Component	Information included
Chosen question and title	Information about your topic and how you chose it and the question you have selected from this
Introductory material	Background data on your topic, diagrams, notes and tables, information about key terms and past experiments similar to that in which you are conducting
Hypothesis and aim	An outline of what you are investigating and why, and what you believe will happen
Planning a method	Show all equipment you plan to use and a clear method you plan to follow, with detailed steps that could be reproducible by someone else
Experimental results	Results observed and recorded in an appropriate form; tables are particularly useful for this
Discussing and analysing results	A discussion and careful evaluation of results that refers back to your hypothesis and questions; any set discussion questions may help scaffold your thoughts and ideas
Thoughts and questions	Any concerns or questions you have about your investigation and researched answers to these
Conclusion	Connection back to the aim and hypothesis, and a summary of the findings of the investigation

FIGURE 20.3 All observations should be recorded in a logbook



20.2.3 Variables

Independent, dependent and controlled variables

In an experiment, a variable is any factor that the researcher can control or change or measure. Three kinds of variables are commonly recognised (see figure 20.4). For some variables, you will set the value at the start of each experiment; others will be determined by your experiment; and sometimes there may be variables that you calculate using your measurements.

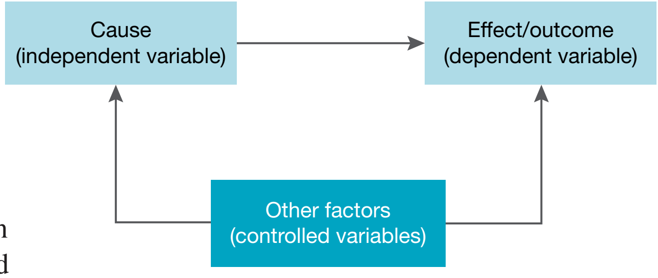
The three main kinds of variables are as follows:

- An **independent variable** is a factor that is *deliberately manipulated by the investigator* and affects the dependent variable. For example, you may be examining the temperature of six different water samples. The independent variable is the different water samples. This may include specific differences in the water samples, such as the source of the sample and the types of substances found in the sample (for example, comparing carbonated water to still water)

  When graphing results, the independent variable is always placed on the horizontal axis.
- A **dependent variable** is the factor that *the investigator measures*. The dependent variable is affected by the independent variable. In the investigation mentioned in the preceding point, the dependent variable would be the temperature.

  The dependent variable is always placed on the vertical axis of a graph.
- **Controlled variables** are all the other factors that the investigator must *maintain at constant values through the course of an experiment*. If these factors are not kept constant, they can confound the experimental results because they can cause changes in the dependent variable. Continuing the same investigation example, controlled variables would include the volume of water tested and the instrument used to record data. Often, controlled variables also include environmental factors such as humidity and air temperature, but these are harder to control.

FIGURE 20.4 The relationships between variables in an experiment

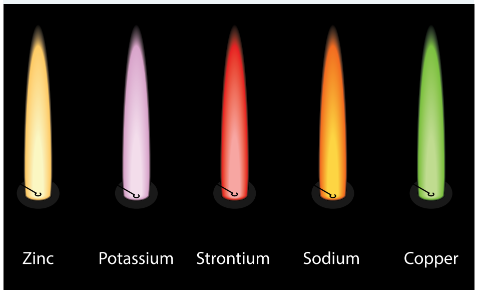


For example, say Allira and Hunter are investigating the colour a flame turns when exposed to different metal ions.

In this investigation, the variables are:

- *Independent variable:* The factor that is being manipulated is the *type of metal ion*.
- *Dependent variable:* The factor that is being measured is the *colour of the flame*.
- *Controlled variables:* The factors to control are ensuring the same amount of metal ion, the same Bunsen burner is used, the temperature of the flame is the same (using the roaring flame for all tests).

FIGURE 20.5 Different metal ions being investigated



We can also apply this to variables associated with water quality, which is the key purpose of this investigation. Unlike the previous example, the changes in independent variable are not always as significant. For example, an investigation could be conducted to examine the pH in different zones in a lake. In this case, your independent variable would be the specific zone the water was obtained from. It is not a different substance or body of water. The dependent variable would be the pH.

Variables can also be considered as numerical (quantitative) or categorical (qualitative). Quantitative includes any value that is numerical; for example, temperature, pH or mass. Qualitative involves variables that are not numerical; for example, colour, the source of the water (whether it was from a lake or an ocean) or the type of salt being examined. See 20.5.2 for further detail on this.

on Resources

Interactivity Variables (int-7731)

20.2.4 Developing aims and questions

Before you start work on your investigation, you need to come up with a topic, create a question to investigate and develop an aim.

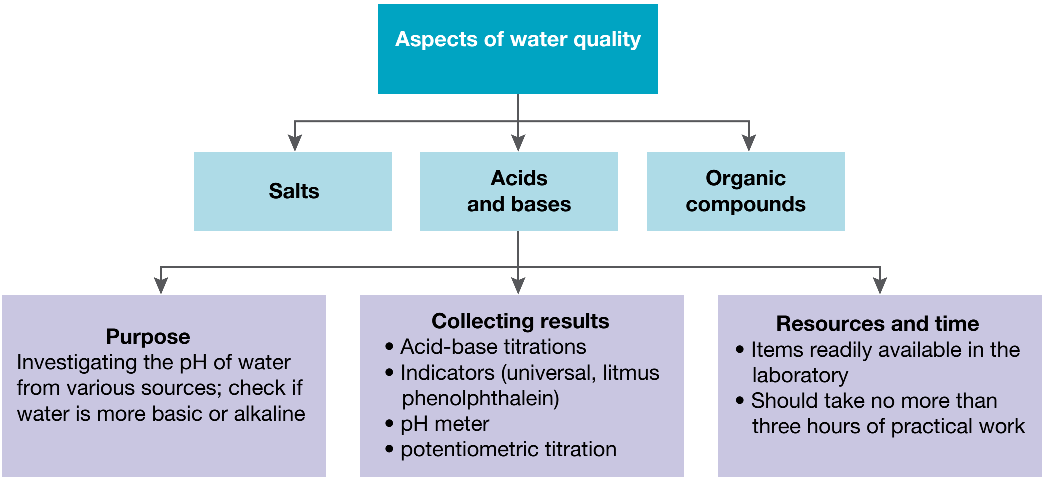
Coming up with a topic

Choosing a topic is not an immediate process — it takes time and careful consideration. It is important you don't just pick a topic that sounds interesting, but instead pick one that is reasonable to complete in the provided time frame and using the available resources.

The topic of your investigation needs to be around water quality, so review the work you completed in Unit 2 Areas of Study 1 and 2 to help brainstorm ideas.

You may wish to create a mind map or a diagram outlining the different aspects of water quality and the different ways that you measure each of these. Research the time this measurement would take and what other research you could conduct. This will help you get your head around the different topics and what requirements there would be.

FIGURE 20.6 Example of the beginnings of a flow chart that may help with selecting a topic



Purpose
Investigating the pH of water from various sources; check if water is more basic or alkaline

Collecting results
- Acid-base titrations
- Indicators (universal, litmus phenolphthalein)
- pH meter
- potentiometric titration

Resources and time
- Items readily available in the laboratory
- Should take no more than three hours of practical work

Some examples of topics may be:

- Salts found in water including minerals, heavy metals and organo-metallic substances
- Sources of acids and bases found in water, such as dissolved CO_2
- Organic compounds found in water, such as dioxins and other contaminants

FIGURE 20.7 Water can vary in pH

Creating a question

Turning the topic into a question focuses your mind on what you want to find out.

The question needs to be:

- one that can be investigated through scientific method
- practicable, given your knowledge, time and the school resources
- asked in a way that indicates what you will do.

As an example of formulating a question from a topic, say your chosen topic is ‘Exploring the amount of sodium in water’. A question you could form from this could be ‘Do samples of water from different taps contain varying levels of sodium?’ Or, ‘How can electrical conductivity be used to determine how sodium concentration differs between various sources of water?’ The first question revolves around examining the variable of sodium concentration in different taps, but doesn’t focus specifically on how it is measured. The second question focuses more on how the sodium concentration is measured, but is less specific with the water source.

You can formulate a question from a topic in many different ways. Just make sure it is something that can be measured, explored and answered in the scope of your practical investigation.

Developing an aim

Often, developing an **aim** of an investigation is done at the same time as formulating a question from your topic. The aim outlines the purpose or the key objective of the investigation. It outlines what you are trying to achieve in order to answer your question and either validate or invalidate your hypothesis.

You can format your aim in two ways:

1. To [determine/investigate/compare] how the **dependent variable** is affected by the **independent variable**
2. To [determine/investigate/compare] how the **independent variable** affects the **dependent variable**

Your aim must:

- be no more than two lines
- be linked to your question
- link the independent and dependent variables.

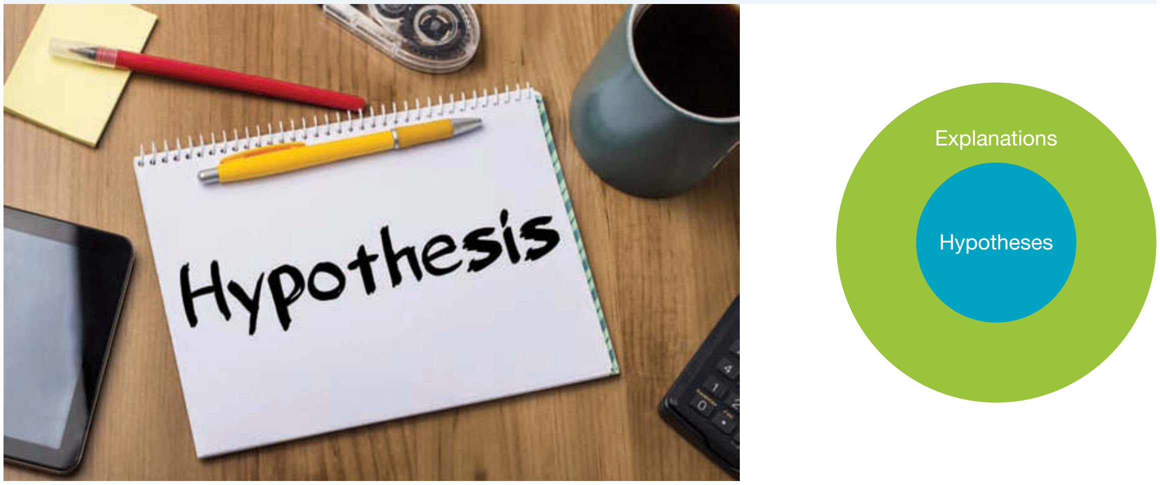
Examples of aims include:

- To determine the solubility of copper chloride in 100g of water
- To investigate the pH of five different brands of soft drink using universal indicator
- To compare the copper concentration between water from indoor taps to the water from outdoor taps
- To observe how changing the temperature of water affects the solubility of copper sulfate
- To determine if blowing bubbles into a bottle of water alters its pH

20.2.5 Formulating hypotheses and making predictions

Formulating a **hypothesis** is an important step in the scientific method.

FIGURE 20.8 A hypothesis is a testable explanation for a concept



A hypothesis is a tentative, testable and falsifiable statement for an observed phenomenon, which predicts the relationship between two variables or predicts the outcome of an investigation.

Testable means that a hypothesis can be easily tested by observations and/or experiments.

Falsifiable means that there has to be a way to prove the hypothesis wrong.

A hypothesis usually predicts the relationship between two kinds of variables: an independent and a dependent variable. Variables were explored further in section 20.2.3.

To write a hypothesis, a good tip is to use the following format:
If [statement involving independent variable], then [prediction involving dependent variable].

Typically, a scientific hypothesis starts with the tentative explanation and includes a prediction by which the hypothesis may be tested.

TABLE 20.3 Example of good working hypotheses

IF	THEN
IF a straw is used to blow bubbles of carbon dioxide into a glass of water	THEN the pH will increase due to the production of carbonic acid
IF the water quality from an old household tap is examined	THEN it will have higher levels of copper due to contamination compared to water from newer pipes

FIGURE 20.9 What would you hypothesise if bubbles were blown through a straw into a glass of water?

FIGURE 20.10 Will water from older pipes contain more copper?

Now consider the following statements and decide if each is an example of a well-formatted hypothesis.

- Statement 1: 'Boiling water will increase its pH.'
  No. This is simply a testable prediction. It does not include a tentative explanation.
- Statement 2: 'IF water is boiled, THEN the recorded pH will increase.'
  No. This does not identify a tentative explanation. The statement is a method followed by a predicted outcome.
- Statement 3: 'IF water with a pH below 7 is boiled and dissolved gases are removed, THEN the pH will increase and become closer to neutral.'
  Yes. This identifies a tentative hypothesis (explanation) and a predicted outcome by which the hypothesis can be tested.

SAMPLE PROBLEM 1

Polly is putting the kettle on. Polly is very curious about science and wants to see how she can change the speed the water boils. She has heard rumours that salt causes water to boil faster. She has four different types of salts in her house: table salt, sea salt, Himalayan pink salt and chicken salt.

Write an appropriate research question, aim and hypothesis for this scenario.

Teacher-led video: SP1 (tlvd-0107)

THINK	WRITE
1. Determine the variables to help write an aim, hypothesis and research question. The factor that Polly is manipulating (the independent variable) is the type of salt. The factor that Polly is measuring (the dependent variable) is the time it takes for the water to boil.	The independent variable is the salt type. The dependent variable is the time it takes for the water to boil.
2. Create a research question based on Polly's problem. Make sure that the question is one that is testable and clearly outlines what is occurring in the investigation.	Does the type of salt added to water affect the time it takes for water to boil?
3. Write an aim that clearly outlines the purpose of the investigation. Be sure to link the independent variable and the dependent variable.	To determine if different types of salts affect the time it takes for water to boil.
4. Write a hypothesis in the 'IF . . . THEN' format. Remember that a hypothesis needs to link the independent variable and dependent variable. Your hypothesis may not be correct, but it must be testable. In this experiment, you may also specify which salt you think will affect the dependent variable the most.	If table salt, sea salt, Himalayan salt or chicken salt is added to water, then the time taken for the water to boil will decrease, with pure table salt causing the largest decrease in time.

PRACTICE PROBLEM 1

Jack wants to know if changing the material of the clothing that Jill is wearing will affect the speed in which she rolls down the hill.

Write an appropriate research question, aim and hypothesis for this scenario.

on Resources

Interactivity Formatting a hypothesis (int-7732)

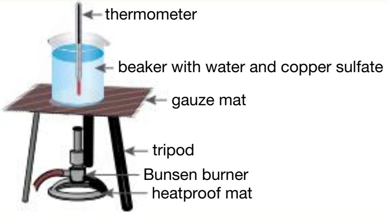
20.2.6 Planning and undertaking experiments

Once your teacher has approved your topic, and you have written a question, aim and hypothesis, the real work begins — planning and conducting the experiments.

Planning experiments

Planning experiments may begin through the use of a practical proposal; an example is given in figure 20.11.

FIGURE 20.11 A practical investigation proposal

Practical investigation proposal	
Name:	Jill
Partner's name: *(optional)*	Jack
Title of your investigation:	Exploring the solubility of copper sulfate in water at various temperatures
Briefly describe its purpose: *(A brief sentence, but needs to be precise)*	To determine if the solubility of copper sulfate in water changes at varying temperatures
Write down three starting questions you want to answer. *(These are to help focus your planning.)*	1. Is there an optimal temperature for copper sulfate solubility in water? 2. Is copper sulfate soluble at all temperatures? 3. What is the solubility of copper sulfate at room temperature?
List independent variables, indicating which are continuous and which are discrete, as well as dependent variables. *(Enables your teacher to see if you have thought of all the obvious variables.)*	Independent: Temperature of the water Dependent: Solubility of copper sulfate, mass of copper sulfate able to dissolve, colour of solution
List the chemistry concepts and relationships that you expect to use in your investigation. *(To give your teacher an indication of the extent of your understanding of the topic)*	Solubility is the extent to which a solute dissolves in a solvent. All sulfates are soluble in water except those formed with silver, lead, calcium, strontium and barium. In water, copper sulfate dissociates as follows: $CuSO_4(s) \rightarrow Cu^{2+}(aq) + SO_4^{2+}(aq)$ The solubility of copper sulfate is known to be 20.5 g per 100 g of water at 20 °C.
List the equipment and measuring instruments that you plan to use. *(For your teacher to see whether you have the right tools for the task.)*	• Copper sulfate solid • Watch glass • Spatula and stirring rod • Scales • Bunsen burner, heat-proof mat, gauze mat and tripod • Thermometer • 200 mL beaker • Deionised water
Sketch your experimental set up. *(This will make your first day of investigating smoother, and your teacher may be able to suggest refinements.)*	
List the steps in your experimental design. *(This is an important stage in your planning and it will enable your teacher to see if there is anything you have forgotten.)*	1. Fill the beaker with 100 mL of water. 2. Place copper sulfate on a watch glass and weigh it. 3. Record the temperature of the room. 4. Slowly add copper sulfate into the water, mixing with a stirring rod. 5. Determine when no more copper sulfate can be dissolved and record the mass. 6. Fill another beaker with 100 mL of water and heat to 40 °C and repeat Steps 4 and 5. 7. Repeat Step 6 at 60 °C, 80 °C and 100 °C.
Any special requests *(E.g. equipment may need to be left set up between classes, or access at lunchtime or after school may be needed.)*	A container to dispose of copper sulfate solution.

At the end of your planning, you should be able to produce three written documents, which should be included in your logbook. These documents are:

- a risk analysis based on the hazards that you have identified (see 20.6.2 for information on risk assessments).
- a detailed list of equipment required, along with quantities (see 20.5.3 for more detail about different equipment used in chemistry)
- an explicit, step-by-step method that takes all of your planning into consideration, including diagrams if relevant.

Conducting investigations

When conducting investigations, it is vital to:

- follow all health and safety protocols
- make sure you know how to use any chosen equipment correctly to minimise errors
- carefully follow your methods and, if any changes are required, note these down in your logbook
- control variables outside your independent variable to keep your results accurate and precise
- clearly record any results obtained, along with the date, including any results that did not go according to plan and any results from both control and experimental groups
- carefully pack up equipment after use. If equipment is required to be set up for a few days, make sure it will be in a location where it cannot be affected by other individuals or environmental factors.
- if time allows, repeat your experiment to improve accuracy and reliability.

20.2 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Using the following topics, create a testable question that could be used for a practical investigation.
   (a) Examining how the pH of water differs at varying temperatures
   (b) Exploring if the amount of salt in water affects its boiling point.
   (c) Calculating the solubility of different compounds in water
2. What is the purpose of a logbook in practical investigations?
3. Describe the difference between a dependent and an independent variable.
4. Why is it important to control variables in an investigation?
5. A student conducted an experiment to measure the effect of changing the amount (measured in moles) on the volume of a sample of gas. This was done in three stages using different amounts of the gas at various temperatures.
   Stage 1: 0.01 mol of gas at 10 °C
   Stage 2: 0.02 mol of gas at 20 °C
   Stage 3: 0.03 mol of gas at 30 °C
   The results from each trial were then analysed to produce an overall conclusion.
   (a) State the independent and the dependent variables in this experiment.
   (b) What were the controlled variables in this experiment?
   (c) In regards to controlling variables, why would the results of this test be difficult to interpret?
6. The following is an excerpt from a report describing the method for an experiment on determining the optimum fertiliser level for lucerne growth.
   Five seeds of *Medicago sativa* (lucerne) were planted in each of 11 identical pots using potting mix from the same packet. These pots were then placed in the same position inside a greenhouse. Each pot was watered daily with 100 mL of fertiliser solution of varying concentrations. The concentrations used in the experiment ranged from 0 to 5 g L^{-1}. After two weeks, the shoots in each pot were cut at soil level and weighed.
   (a) What is the independent variable in this experiment?
   (b) What is the dependent variable?
   (c) State as many controlled variables as you can think of.
   (d) Which pot is the control in this experiment (is not affected by the independent variable)?
7. The following table outlines an investigation topic with some variables identified. Complete the table for the three other topics listed.

Topic	Independent variables		Dependent variables
	Numerical	**Categorical**	
Bouncing basketball	• Drop height • Pressure of the ball	Surface ball lands on, ball type	Rebound height, impact time, energy loss, change in momentum, average force of impact
Conductivity of metal			
Concentration of salt in water			
Boiling point of different soft drinks			

8. Select which of the following is an acceptable hypothesis.
   A. Does cordial have a higher pH than water?
   B. If the amount of hydrogen ions is higher in cordial compared to water, then the cordial is more acidic and will be recorded at a lower pH.
   C. If cordial is more acidic, it will have a lower pH recorded.
9. Testing a scientific question by experiment involves a number of stages. These are shown by the following statements. Use the letters to put these stages into their correct order.
   A. Formulate the hypothesis.
   B. Decide on the question.
   C. Analyse the results.
   D. Communicate the results.
   E. Plan the experiment.
   F. Carry out the experiment.
10. After some preliminary reading, a student has become intrigued by the possibility that hydrocarbons with double bonds (alkenes) have lower boiling points than those with single bonds (alkanes). Therefore, she proposes the question: Do alkanes and alkenes have different boiling points?
    Write a reasonable hypothesis that she could test experimentally based on this question.
11. Which of the following is a characteristic of a good hypothesis?
    A. It must be proven true.
    B. It must be testable by observation or experiment.
    C. It must be based upon experiments done by other scientists.

Fully worked solutions and sample responses are available in your digital formats.

20.3 Concepts specific to an investigation

KEY CONCEPT

- The chemical concepts specific to the investigation and their significance, including definitions of key terms, and chemical representations

KEY SCIENCE SKILL

- Communicate and explain scientific ideas

20.3.1 Concepts specific to investigations

As part of an investigation, it is vital to include key chemistry concepts that are relevant and clearly explain their significance. This shows a clear link to your understanding of an investigation, and allows others to see the connection between theory and practical applications.

Concepts should be researched prior to commencing your investigation, recorded in your logbook (and referenced). This background information also will form part of your introduction in your poster.

Concepts that are relevant to your investigation include:

- explanations of key formulae
- definitions of key terms
- detail about the theories being examined
- information about other practical investigations exploring similar concepts.

An example of this may be in calculating the number of molecules in a set amount of water. In your background information, it would be important to describe previous findings and experiments conducted, identify use and application of the investigation, and describe key formulae relevant to it (for example, $N = n \times N_A$), including the identification of the symbols used.

You should also have clear concepts links to theory in your **discussion** section of the investigation.

20.3.2 Key terms

In practical investigations, defining any key terminology is vital.

This can be done in two ways:

- within a report itself
- as part of an appendix or glossary at the end of the report.

The following excerpt shows an introduction from a scientific report written by a student. This investigation was conducted to explore different types of polymers.

> *Polymers are large molecules made by joining smaller molecules (monomers) together. They form a wide range of substances, both natural and synthetic. Polymers are formed by the addition polymerisation of alkenes, which are hydrocarbons that contain one double bond between carbon atoms.*

This student has clearly defined key terms as part of their introduction within their report itself. What terms have they defined?

Read the following excerpt of an introduction from a scientific report written by a different student, investigating the same practical.

> *In this investigation, different types of **polymers** are being investigated. Polymers that are formed through the addition polymerisation of **alkenes** are specifically being explored in this investigation.*
>
> *Glossary of key terms:*
>
> ***alkenes:** hydrocarbons that contain one double bond between carbon atoms*
>
> ***polymers:** large molecules made by joining smaller molecules (monomers) together*

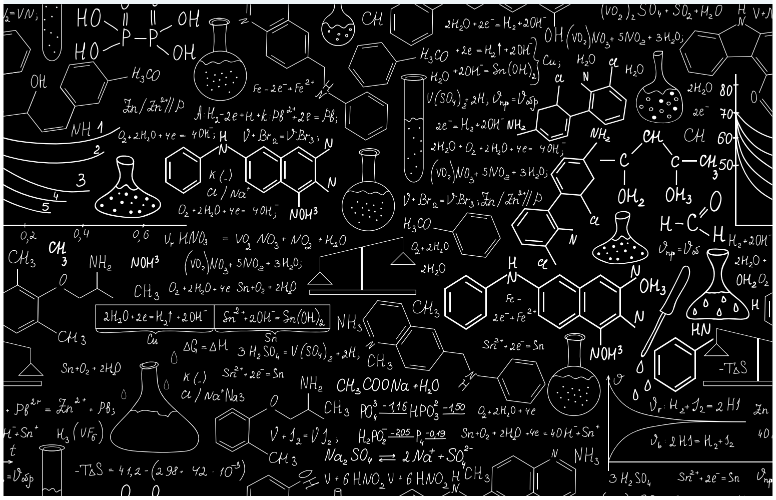
In this situation, the student has not defined the terms in their introduction itself, but has bolded key words that later appeared in their glossary.

Both of these formats are valid methods to define key terms. The method used depends on the format of the report or personal preference. Be sure to check with your teacher which method is most suitable for the practical investigation you are conducting.

20.3.3 Chemical representations

A variety of representations are used in chemistry. This includes the use of models, sketches, graphs, equations, formulae, symbols and diagrams. As well as this, many vital conventions exist regarding the use of numerical data, including significant figures and scientific notation. Perhaps the most common chemical representation is the use of chemical formulas. Care should be taken with capital letters and subscripts and superscripts when representing atoms and ions. For example, CO is carbon monoxide, while Co is the metal cobalt.

FIGURE 20.12 Representations form a vital part of chemistry reporting



Formulas which demonstrate structure

Structures are often drawn in a skeletal form in chemistry, particularly in organic chemistry. One compound can be represented in various ways, as seen in figure 20.13

FIGURE 20.13 Different representations of methane: (a) ball-and-stick model, (b) diagram showing bond angles, (c) structural formula, (d) shape diagram

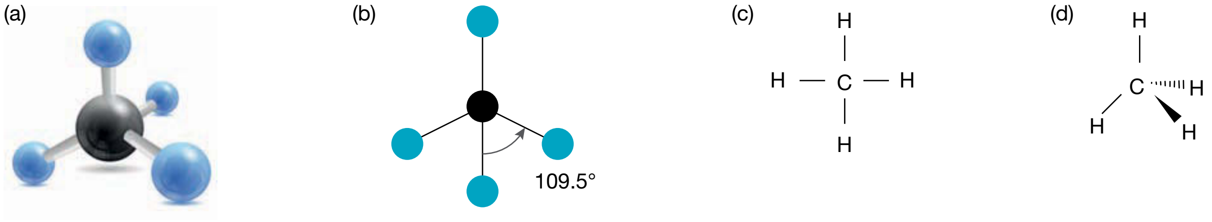
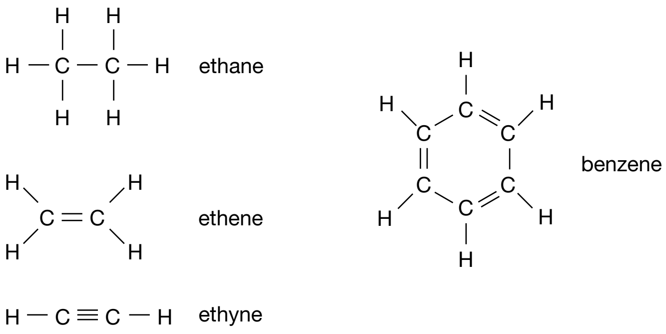
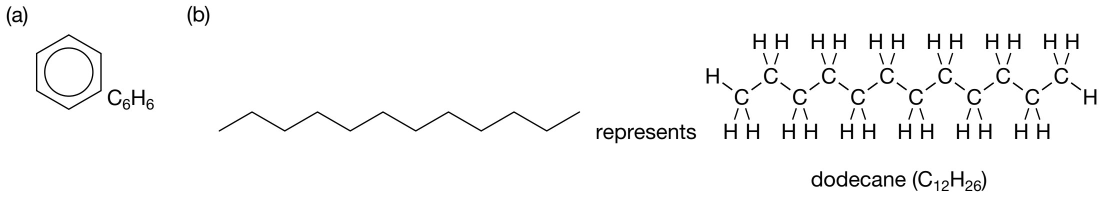


FIGURE 20.14 Some common structural formulas representing various molecules



Lines between atoms represent the number of bonds present. Sometimes we can shorten these representations even further, as shown in the skeletal formulas for the benzene ring and dodecane in figure 20.15. Skeletal formulas do not show specific carbon and hydrogen atoms or the bonds connecting them.

FIGURE 20.15 Skeletal formulas of different molecules: (a) benzene and (b) dodecane.



Scientific notation

Very large and very small quantities can be more conveniently expressed in scientific notation. In scientific notation, a quantity is expressed as a number between 1 and 10 multiplied by a power of 10.

To write in scientific notation, follow the form $N \times 10^a$, where N is a number between 1 and 10 and a is an integer (positive or negative).

When converting numbers into scientific notation, the following steps should be followed:

1. Determine where the decimal point needs to go so that N is between 1 and 10.
2. Count the number of places the decimal point is moved to determine a (the power of 10 or the exponent). If the decimal point was moved to the left, a will be positive; if it was moved to the right, a will be negative.
3. Write the number in scientific notation.

For example, the average distance between the Earth and the moon is 380 000 000 m. This is more conveniently expressed as 3.8×10^8 m, in which the decimal point was moved eight places to the left. The radius of a lead atom in metres is 0.000 000 000 175. This is more conveniently expressed as 1.75×10^{-10} m, in which the decimal point was moved 10 places to the right.

As you can see, very large numbers will have a positive exponent (a), whereas very small numbers will have a negative exponent. For example, 5×10^{-3} can be written out as 0.005, whereas 5×10^3 is written out as 5000.

In chemistry, scientific notation is generally used for numbers less than 0.01 and greater than 1000.

Quantities in scientific notation can be entered into your calculator using the EXP button or ˆ button.

SAMPLE PROBLEM 2

a. The average distance between Earth and the sun is 149 600 000 kilometres. Write this in scientific notation.

b. The mass of a proton is 0.000 000 000 000 000 000 000 001 67 g. Write this in scientific notation.

Teacher-led video: SP2 (tlvd-0103)

THINK	**WRITE**
a. 1. Determine the position of the decimal point for the number to be between 1 and 10 and remove any zeros that are not between non-zero digits. The decimal point would need to go between 1 and 4 to form 1.496.	1.496

2. Determine the exponent (the power of 10) by counting the number of places the decimal point was moved. If the decimal point was moved to the left, the exponent will be positive; if the decimal point was moved to the right, the exponent will be negative. The decimal point was moved eight places to the left so the exponent is 8.	10^8
3. Write the number in scientific notation, remembering to include the units.	1.496×10^8 km
a. 1. Determine the position of the decimal point in order for the number to be between 1 and 10 and remove any zeros that are not between non-zero digits. The decimal point would need to go between 1 and 6 to form 1.67.	1.67
2. Determine the exponent by counting the number of places the decimal point was moved. If the decimal point was moved to the left, the exponent will be positive; if the decimal point was moved to the right, the exponent will be negative. The decimal point was moved 24 spots to the right so the exponent is –24.	10^{-24}
3. Write the number in scientific notation, remembering to include the units.	1.67×10^{-24} g

PRACTICE PROBLEM 2

Express the following quantities in scientific notation.

- a. **The diameter of Saturn's rings, 282 000 km**
- b. **The number of metres that sound travels in one hour, 1 235 000**
- c. **The uncertainty of a highly precise clock, 0.000 000 000 000 000 003 seconds**

20.3 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Explain two ways that key terms can be defined in your report.
2. Why are representations useful in scientific reporting?
3. Express the following quantities in scientific notation.
   (a) A red blood cell, about 0.000 008 m across
   (b) A flea, about 0.0013 m long
   (c) The moon, 384 400 000 metres from the Earth
4. Convert each of the following numbers from scientific notation to the standard form of the number.
   (a) 7.21×10^{-4}
   (b) 5.739×10^5
   (c) 1.552×10^1
   (d) 7.112×10^0
   (e) 4.71×10^2
   (f) 2.577×10^{-6}
   (g) 9.179×10^{-5}

Fully worked solutions and sample responses are available in your digital formats.

20.4 Scientific research methodologies and techniques

KEY CONCEPTS

- The characteristics of laboratory techniques of primary qualitative and quantitative data collection relevant to the investigation: sampling protocols; gravimetric analysis, acid–base titrations and/or pH measurement; precision, accuracy, reliability and validity of data; and minimisation of experimental bias
- Methods of organising, analysing and evaluating primary data to identify patterns and relationships including identification of sources of error and uncertainty, and of limitations of data and methodologies

KEY SCIENCE SKILLS

- Conduct investigations to collect and record data
- Analyse and evaluate data, methods and scientific models

20.4.1 Characteristics of scientific research methodologies

Carefully following **scientific research methodology** when conducting practical investigations is important to ensure that your results are as precise, accurate, reliable and valid as possible. This includes minimising errors and uncertainties in data in order to draw **conclusions** in relation to your question.

Each type of research method has its specific purposes, procedures, advantages and limitations. The researcher's choice depends on which method is most appropriate for the specific topic of research interest and hypothesis being tested.

Different types of scientific inquiry and research methods can be used, as shown in table 20.4.

TABLE 20.4 Types of scientific inquiry

Method	Use
Controlled experiment	Determining the relationship between an independent and dependent variable. All other variables are controlled. An example is testing the melting temperatures of different metals.
Single variable investigation	Exploring a variable change over time and understanding causes of the observations and how other factors affect it. An example is investigating how an element decays over time.
Identification	Investigating if objects or events belong to specific sets or are part of new sets. An example is understanding if a new element is a metal or non-metal.
Designing	Designing a device using scientific knowledge. An example is designing, constructing, testing and evaluating a device that can be used to treat and purify water.
Investigating models	Exploring everyday phenomena and collecting evidence to test the model. An example is devising an investigation to test the law of conservation of mass.

In all scientific inquiries, various processes are followed. This includes:

- formulating a question and hypothesis to be tested
- controlling variables
- completing a log book outlining the introduction, methodology, results, discussion and conclusion of an investigation.
- ensuring that methods are being taken to allow for validity, accuracy, precision and reliability
- ensuring that methods are being used that reduce uncertainties and errors

- collecting data accurately in an appropriate form that best suits the question being investigated
- where possible, having a **control group** (not exposed to an independent variable to compare as a baseline) and numerous **experimental groups** (exposed to the independent variable).

Control groups

Usually, if the identification of controlled variables is complex, or many are present, it might be simpler to include a control group. A scientific control is an experiment or observation that involves all variables except the independent variable. For example, a scientist may be trying to measure the absorbance of light by various concentrations of cobalt chloride, $CoCl_2$, solutions. The independent variable in such an investigation would be the concentration of cobalt chloride, and the dependent variable would be the absorbance reading obtained. However, a number of other variables may affect the result. These include the nature of the solvent, the type of glass that the containers holding the solutions are made from, the distance the light has to travel (especially through the solution) before it is measured and the temperature of the solution.

A convenient way to control all these variables, and maybe even some that you aren't aware of, is to use a control. Everything about the control, from the way it is prepared to how it is manipulated and measured in the experiment, is the same as for the test solutions containing cobalt chloride. The only difference is that no cobalt chloride (the controlled variable) is in the control. This, therefore, allows the scientist to isolate the amount of absorption in each reading that is due to the cobalt chloride alone because it is the only variable left responsible for any differences in absorbance readings.

Another example of control and experimental groups being used may occur in a practical investigation exploring the conductivity of different concentrations of sodium chloride, NaCl, solution. The experimental groups would be the different concentrations of sodium chloride being tested. The control group would be testing the circuit with pure water (0% sodium chloride), to examine the conductivity when no NaCl is present.

20.4.2 Techniques of primary qualitative and quantitative data collection

Data is a set of facts that are collected, observed or generated. Data that you gather may come from primary sources or from secondary sources. Typically, data that you collect is raw data that must later be analysed and interpreted to produce useful information.

Primary sources of data provide direct or firsthand evidence about some phenomenon — as, for example, a research investigation. Your completed logbook will be a primary source of data about pursuing an investigation into your research question.

Secondary sources of data are comments on or summaries and interpretations of primary data. Sources of secondary data include review articles in newspapers and popular science magazines in which other individuals summarise and comment on the research of others.

FIGURE 20.16 Recording data is vital in investigations

Qualitative and quantitative data

Qualitative data (or **categorical data**) are expressed in words. This type of data is descriptive and not numerical and can be easily observed but not measured. Bar graphs or pie graphs are often used to display the frequencies of categorical variables.

The two types of qualitative data are:

- *Ordinal data*, which can be ordered or ranked; for example, ionisation energies (first, second, third) or opinion polls (strongly agree, agree, disagree, strongly disagree).
- *Nominal data*, which cannot be organised in a logical sequence; for example, types of sub-atomic particles (proton, neutron or electron) or the colour of a solution after a metal salt has been dissolved (blue, clear, white, yellow).

Quantitative data (or **numerical data)** can be precisely measured and have values that are expressed in numbers. Line graphs or scatter plots are often used to display the frequencies of numerical variables.

The two types of numerical data are:

- *Continuous data*, which can take any numerical value, such as the release height of a parachute, the temperature of a solution or the mass of a substance.
- *Discrete data*, which can only take on set values that can be counted, such as the number of protons in an atom or the number of electron shells.

Table 20.5 shows examples of how some attributes can be expressed both qualitatively and quantitatively.

TABLE 20.5 Examples of data types

Attribute	Qualitative	Quantitative
Length of pendulum	Long	52 cm
Colour	Green	520 nm
Sound	Loud	85 decibels
Speed	Fast	120 kph
Temperature	Hot	100 °C
Gravitational force	Strong	$9.807\ \text{m s}^{-2}$

When collecting data, it is vital to consider what is most appropriate for your investigation. Normally, the best evidence is primary quantitative data, and for a majority of your investigations this is what should be collected and recorded. However, while preferred, it is sometimes not always possible to collect quantitative data. In this case, qualitative data can be collected instead. Qualitative data can be quite subjective or open to interpretation (one student may say dark, for example, while another may say dark blue and a third may say black), so it is important to make sure that the data you collect and record is clear and as detailed as possible.

In terms of primary data collection, it is important that results are carefully checked to make sure that recorded data is correct. Many people interpret measurements slightly differently, or use the wrong units, so make sure you are double checking data. All your collected data should be recorded in your log book. Ensure you note down all observations (usually in a table); you may find noting down both the qualitative and quantitative data gives the clearest detail for your investigation.

20.4.3 Choosing techniques relevant to an investigation

It is important to select techniques that will best suit your investigation. This includes the use of experiments around water quality and chemical analysis.

When selecting appropriate techniques, it is important to ensure the following:

- the technique can be performed in an appropriate time frame
- the technique is appropriate to your investigation and serves a purpose to answering the question and supporting or rejected your hypothesis
- the data is easily recorded, measured and interpreted, with a particular emphasis on quantitative data
- the technique can be safely performed. This is particular important in a school environment, where health and safety restrictions are closely regulated.

- the equipment used in the technique is available and cost effective. If it is not available in a school, can it be used with permission at other locations?
- the technique allows for the control of other variables. If too many factors cannot be controlled and will impact results, the technique isn't a great choice for an investigation.

Selecting your measuring instruments

You should also consider the most appropriate equipment to use for a particular purpose.

For example, if a liquid volume of 25 mL is required, what would be the most appropriate piece of equipment to measure it? If a high-precision measurement is required (that is, providing a lower uncertainty), a 25.00 mL pipette would probably be used. If, however, this level of precision is not required, a measuring cylinder or even a 100 mL beaker may be more appropriate.

Your school will have a range of measuring instruments, and these will vary in precision and ease of use.

You won't always need to use the most accurate instrument. A simple instrument that allows for quick measurements will be enough more often than not. Sometimes a simple stopwatch is just as good as an electronic timer, for example, or a voltmeter may compare well to a more accurate multimeter.

FIGURE 20.17 Different types of measuring equipment vary in precision

Some instruments that you might consider are as follows, listed based on what they measure.

Mass

- *Beam balance:* Accurate with a large range of values; can be time-consuming to measure several masses.

FIGURE 20.18 Beam balance

- *Spring balance:* Quick to use; covers a large range of masses; not very accurate.

FIGURE 20.19 Spring balance

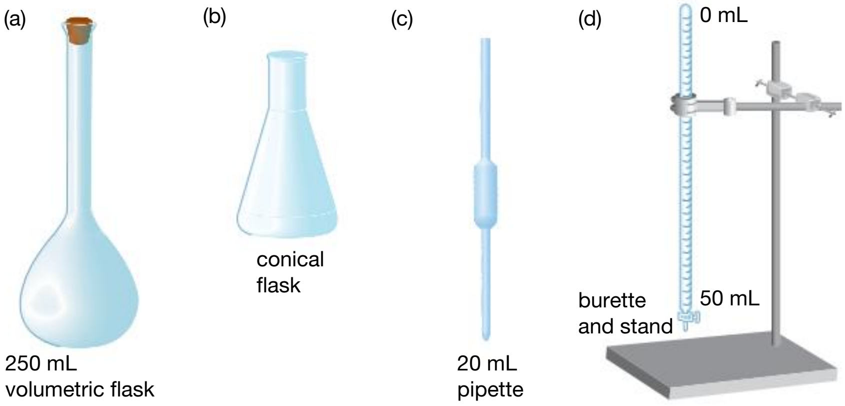
- *Top loading balance:* Very accurate; very good for small masses; simple to use. With equipment set up above the balance, it can be used to measure small variations in attractive and repulsive forces such as magnetic force, electric force and surface tension. If the balance sits on a laboratory jack, force against distance can be easily measured.

FIGURE 20.20 Top loading balance

pH

- *Titrations:* Used to determine the concentration of a solution, particularly through acid–base titrations. Detailed descriptions of titration equipment are provided in topic 19.

FIGURE 20.21 Equipment used in titrations. (a) A volumetric flask is used to prepare a standard solution. (b) A conical flask holds the solution of unknown concentration. (c) A pipette is used to add the unknown solution to the conical flask. (d) The burette holds the standard solution, which is added to the conical flask.



- *pH meter:* Measures pH accurately, but can easily break if not maintained or stored correctly.
- *Litmus paper:* Allows for a quick visual to determine if a solution is acidic or alkaline.
- *Universal indicator:* Allows for an easy visual representation to determine the approximate pH of a solution.

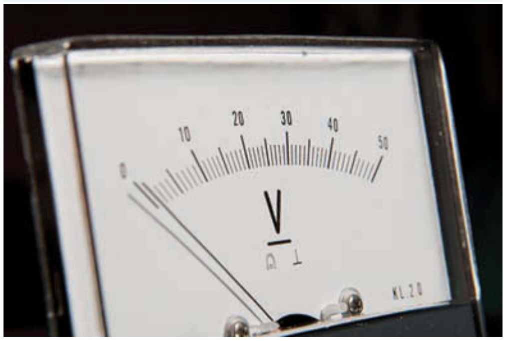
Time

- *Stopwatch:* Simple to use; accurate down to your response time; not reliable for short time intervals.
- *Electronic timer:* Requires some instruction; very accurate; best suited for short time intervals; can be used with electrical contacts and photogates.

Electrical

- *Meters:* Includes voltmeters, ammeters, galvanometers; easy to set up, but care is needed to ensure the meter is wired into the circuit correctly, or the meter can be damaged; large range of values; usually analogue displays.

FIGURE 20.22 Voltmeter



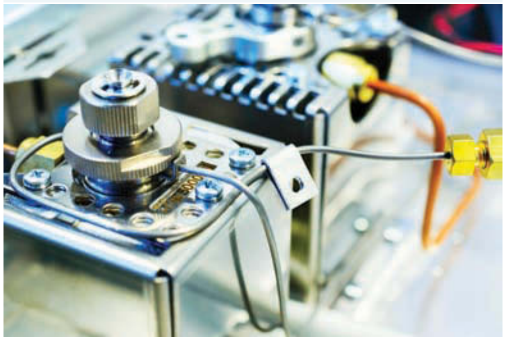
- *Multimeters:* Easy to set up; more tolerant of incorrect use, but can be damaged if incorrectly connected to a high current; large range of values; usually digital displays.

Specialised equipment to measure concentration

You may also have access to specialised equipment you can use in your practical investigation. Although these are unlikely to be available in a general school laboratory, it is important to note that the following equipment is all highly accurate and precise. Discuss with your teacher if any of these are available for use at your school or nearby. Alternatively, you may wish to explore these as a point of discussion when outlining uncertainties and possible errors in the data and results you obtained. Specific uses of this specialised equipment is discussed in topics 17 and 18.

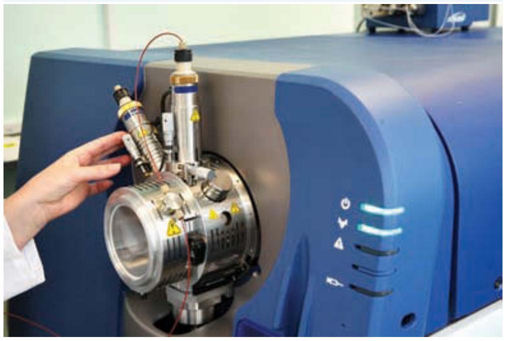
- *Gas chromatograph:* Measures the content of various components in an injected sample.

FIGURE 20.23 Gas chromatograph



- *Mass spectrometer:* An analytical instrument that determines the relative isotopic masses of the different isotopes of an element and abundance.

FIGURE 20.24 Mass spectrometer



- *Instrumental colorimeter:* An instrument that compares the colour in the test sample with the colours produced in samples of known concentration, allowing for the concentration in the sample to be determined based of absorption.

FIGURE 20.25 Colorimeter

- *Atomic absorption spectrometer:* Absorbs light to measure concentrations of metal ions.

FIGURE 20.26 Atomic absorption spectrometer

- *HPLC (high-pressure liquid chromatography):* Used to measure the concentration of organic substances.

FIGURE 20.27 High-pressure liquid chromatography

20.4.4 Precision, accuracy, reliability and validity

For your investigation, it is important to carefully consider the data you obtain and the confidence you can have in the conclusions drawn. Understanding the precision and accuracy of your data is important to ensure your findings are both reliable and valid.

Precision

Precision refers to how close multiple measurements of the same investigation are to each other. Note that results that are precise may not be accurate. It is often difficult to have completely precise results due to random error.

Table 20.6 shows two investigations by different students. Student 1 has more precise results, because the range of their measurements is much smaller when compared to Student 2.

TABLE 20.6 Two investigations by different students measuring the point in which water boils

Investigation by Student 1		Investigation by Student 2	
	Temperature (°C)		Temperature (°C)
Trial 1	98.5	Trial 1	100
Trial 2	98.6	Trial 2	102
Trial 3	99	Trial 3	95
Trial 4	98	Trial 4	99
Trial 5	99.2	Trial 5	106

FIGURE 20.28 Results vary in precision when recording temperature

Accuracy

Accuracy refers to how close an experimental measurement is to a known value. If an archer is accurate, for example, their arrows hit close to the target. Consider an experimental calculation of the boiling point of water, which is known to be 100 °C. A student who obtained a experimental value of 99 °C is more accurate than a student who obtains a value of 105 °C.

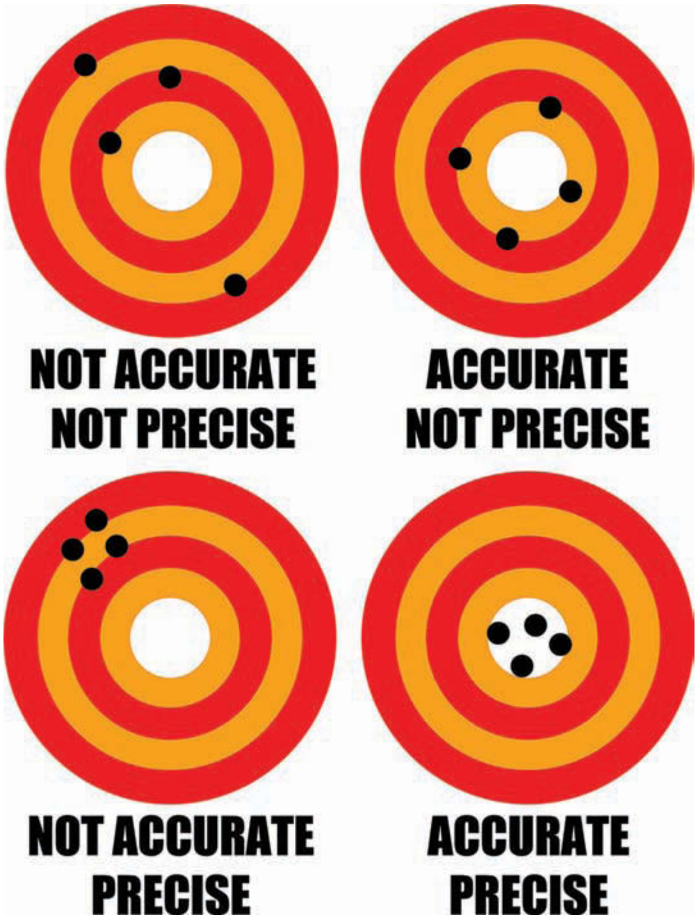
Table 20.7 again shows two investigations by different students. Student 1 has more accurate results, because their results are much closer to the actual boiling point of water. Their measurements are no more than 2 °C from the known value.

TABLE 20.7 Two investigations by different students to determine the boiling point of water

Investigation by Student 1		Investigation by Student 2	
	Temperature (°C)		Temperature (°C)
Trial 1	100	Trial 1	95
Trial 2	102	Trial 2	94.8

It is important to note that a student may have accurate results that are not precise, or precise results that are not accurate. This is seen in table 20.7, in which Student 1 has the most accurate result. However, Student 2 has the most precise results. In order to obtain the best experimental data, we want results that are both accurate and precise as seen in figure 20.29.

FIGURE 20.29 Comparing precision and accuracy



SAMPLE PROBLEM 3

Students conducted an experiment to determine the temperature of a substance as it changed from a solid to a liquid. They repeated the experiment four times and achieved the following results:

Student 1: 56.5 °C, 58.0 °C, 60.0 °C, 55.0 °C
Student 2: 60.5 °C, 61.0 °C, 60.5 °C, 62.0 °C
Student 3: 56.5 °C, 58.5 °C, 57.0 °C, 56.0 °C

Students were then provided with the exact value of the melting temperature of the substance, which was found to be 56.48 °C.

a. **Which student had the least accurate data?**
b. **Which student had the least precise data?**
c. **Was the student with the most precise data also the student with the most accurate data? Explain your answer.**

Teacher-led video: SP3 (tlvd-0218)

THINK	WRITE
a. 1. Review what accuracy means.	Accuracy refers to how close a measurement is to a known value.
2. Explore the data of the three students.	Student 1 had data 1.48 °C lower and 3.52 °C higher than the actual data. Student 2 had data that was up to 5.52 °C higher. Student 3 had data that was 0.48 °C lower and data that was 2.02 °C higher.
3. Determine which student had the least accurate data.	Student 2 had the least accurate data, because their values were the furthest away from the actual value.
b. 1. Review what precision means.	Precision refers how close multiple measurements of the same investigation are to each other.
2. Explore the data of the three students.	Student 1 had a data range of 5.0 °C. Student 2 had a data range of 1.5 °C. Student 3 had a data range of 2.5 °C.
3. Determine which student had the least precise data	Student 1 had the least precise data.
c. 1. Identify the students with the most accurate and most precise data.	Using the results from parts **a** and **b** it can be seen that Student 3 had the most accurate data and Student 2 had the most precise data.
2. Respond to the question and explain your answer	The student who had the most precise data was not the same student who had the most accurate data. Students may have measurements very close together (precise), but these may not be accurate. This may be to errors in their measuring device or their interpretation of the melting point (when the solid is a liquid). Data may also be accurate without being precise: you can be close to the target, but the readings are inconsistent. For reliable and valid results, data should be both accurate and precise.

PRACTICE PROBLEM 3

Students conducted an experiment to determine the temperature of a substance as it changed from a liquid to a gas. They repeated the experiment four times and achieved the following results:

Student 1: 85.4 °C, 92.0 °C, 82.0 °C, 75.5 °C
Student 2: 83.5 °C, 85.0 °C, 85.5 °C, 86.5 °C
Student 3: 85.5 °C, 90.0 °C, 89.5 °C, 81.0 °C

Students were then provided with the exact value of the boiling temperature of the substance, which was found to be 85.4 °C.

a. Which student had the least accurate data?
b. Which student had the least precise data?
c. Was the student with the most precise data also the student with the most accurate data? Explain your answer.

Reliability

Reliability refers to whether another researcher could repeat your investigation by following your method and obtain similar results. As well as this, the more times an experiment is replicated, the more reliable the results are considered to be.

Experiments that are reliable shouldn't just be able to be carried out by the investigator, but also by the third party. If a test is reliable, it usually has both accuracy and precision, because errors are reduced through repetition. In the investigation in which students were measuring the boiling point of water, a student who repeated the experiment 15 times and then took an average would end up with far more reliable results than a student who only repeated the experiment twice.

Validity

The **validity** of data refers to the credibility of the research results from experiments or from observations. Validity factors in both experimental design and implementation. Experiments that are valid usually are using the results from one manipulated variable, where other variables are controlled. Validity is also affected by factors such as experimental bias.

Validity applies more to biology and psychology, where precise measurement is more difficult and the risk of bias or misinterpretation of data is higher. In physics and chemistry, the variables are quantifiable and physically measurable.

If your experimental method clearly relates to the purpose of the investigation and you take care to be precise in your measurements and thorough in your analysis, your results should be valid and meaningful.

Validity can be:

- *Internal:* Internal validity focuses on if the results can be believed and easily interpreted. Data that shows internal validity ensures that the results have not been affected by other factors, and that all variables (other than the independent variable) have been controlled. Did the experiment measure the variable that was being examined in the experimental question and outlined in the hypothesis?
- *External:* This does not often apply in chemistry, but is the idea that the results obtained in a sample are indicative of the results expected for the entire population. External validity can be improved by removing bias and using a large and random sample. This is more prevalent in drug trials to ensure that a varied and large sample size is used.

Measurements that are valid must also be reliable, and therefore must have accuracy and precision.

20.4.5 Minimisation of experimental bias

Bias is an intentional or unintentional influence on a research investigation as a result of systematic errors introduced by the researcher into the sampling or the testing procedures of an experiment. These biases will prejudice the research findings and raise questions about their validity and reliability.

FIGURE 20.30 Selection bias

Numerous types of bias are possible in experiments — some of which apply more to chemistry than others. Types of bias include:

- *Measurement bias:* This bias occurs when experimenters manipulate results in order to get a desirable outcome. Sometimes this can be unintentional (for example, if an experimenter consistently records the boiling point earlier than they should, leading to a lower recorded temperature). Often, however, it is through the deliberate actions of an individual (for example, when measuring the rate of a reaction, an experimenter might deliberately stir one reaction to make the rate appear higher to better suit their hypothesis).
- *Selection bias:* This type of bias can arise when test subjects are not randomly assigned to the experimental and control groups.

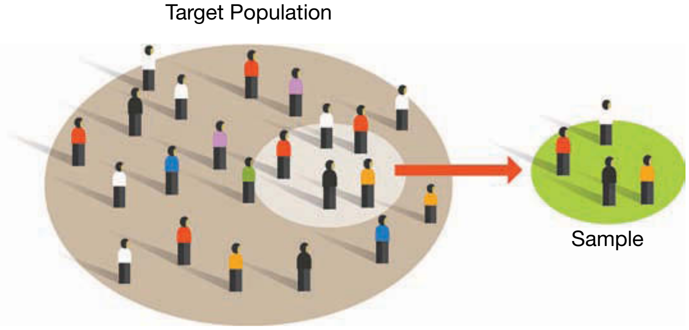
  An example of selection bias is in clinical trials of a new synthetic drug. A doctor may choose family members to receive a drug being tested and have individuals he doesn't know receive a placebo.

  Selection bias can be minimised by randomly and equally allocating subjects to each group.
- *Sampling bias:* This type of bias can arise if the subjects chosen for the study are not representative of the target population. If this occurs, the research results cannot be generalised to that population.

  For example, the average height of students at a school is calculated, but due to time constraints only 50 out of the 600 students are measured. If only year 7 students are measured, this will not be representative of the target population and is an example of sampling bias.

  Sampling bias can be minimised by ensuring that the participants in the study are a reasonable representation of the target population.

FIGURE 20.31 Sampling bias



- *Response bias:* This type of bias arises when only certain members of the target population respond to an invitation to participate in the clinical trial, resulting in an unrepresentative sample of the larger population.

  Similar to sampling bias, response bias can be minimised by ensuring that the subjects responding to be included in the study are a reasonable representation of the target population.

FIGURE 20.32 Response bias

20.4 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Distinguish between validity and reliability.
2. Is it essential that the results of an experiment can be replicated in order for the experiment to be considered reliable? Explain your answer.
3. Give an example of when results would not be considered reliable.
4. Under what circumstances can it be said that the conclusions or findings of research are ‘valid’?
5. List two procedures that could adversely have an impact on the internal validity of an experiment.
6. Explain, with reference to an example, how an experiment can be reliable but not valid, whereas an experiment that is valid must also be reliable.
7. Explain, with examples, the difference between precision and accuracy.
8. Give an example of a strength and a weakness of quantitative and qualitative data.

Fully worked solutions and sample responses are available in your digital formats.

20.5 Ethics, health and safety guidelines

KEY CONCEPT

- Ethics of and concerns with research, including identification and application of relevant health and safety guidelines

KEY SCIENCE SKILL

- Comply with safety and ethical guidelines

20.5.1 Ethics

Ethics are the principles of acceptable and moral conduct. This does not just apply to scientific investigations but to many aspects of life, determining what is ‘right’ and what is ‘wrong’.

Science interacts with ethics in several ways, including:

- the way an experiment is conducted
- confidentiality and morality around research
- conflicts with religious and personal beliefs.

Ethical standards and considerations also apply to any type of research or data collection method involving people (or animals).

FIGURE 20.33 Ethics relate to the idea of moral choice

relation or from a
point of view.
Ethics [eth'iks] n.
moral choices to
value of human
principles that
for what is thou

Ethics are particularly obvious in drug trials, both with animal testing and human trials. It is important that individuals give permission and are made aware of all possible side effects and risks associated with treatments.

While drug trials are not relevant for your practical investigation, ethical considerations still need to be made. For example, you may be conducting surveys on issues relating to opinions on water quality. The confidentiality of individuals who provide responses needs to be considered and permissions gained from respondents.

It is also important to be mindful of individuals in regards to personal beliefs. While drug trials have minimum ethical standards for the use of animals in trials, for some individuals, differing personal beliefs may affect experimentation and interpretation of data. This is an ethical consideration that needs to be evaluated and understood when researching and reporting on these topics.

20.5.2 Health and safety guidelines

Part of the enjoyment of a practical investigation is that the topic may be unconventional or use an innovative method. Such situations, however, can present some risk, so special care needs to be taken to ensure you and others are safe.

General safety rules

Some general safety precautions help to ensure that you and others are not injured in the laboratory. This includes the following:

- Wear protective clothing. This might include laboratory coat, safety glasses and gloves.
- Be aware of the position of safety equipment such as the fire blanket, fire extinguisher, safety shower and eye wash.
- Ask if you are unsure how to operate equipment or how to use apparatus.
- Read labels carefully to confirm contents and concentration of chemicals.
- Clean and return all equipment to the correct places, ensuring lids are placed back on containers when not in use.
- Check for the correct disposal of equipment and chemicals, including damaged equipment (such as broken glassware).
- Read instructions carefully before commencing an experiment.
- Prepare a risk assessment for required chemicals and equipment.
- Do the investigation as outlined in your approved plan. Don't vary your plan without approval from your teacher.
- Don't do experimental work unsupervised unless you have prior approval from your teacher.
- When first setting up electrical experiments, ask your teacher to check the circuit.
- Don't interfere with the equipment set-up of others.

It is important to address health and safety concerns through the use of a **risk assessment**.

A risk assessment is a procedure for identifying hazardous chemicals, what the risks are and how to work safely with them. Table 20.8 lists the usual requirements for a written risk assessment.

This assesses potential hazards with equipment being used and outlines standard handling procedures to ensure the health and safety of individuals and the environment.

Risk assessments should also take into considerations correct disposal of equipment and chemicals to adhere to safety and bioethical guidelines. Many chemicals are harmful to the environment, so correct disposal is paramount.

TABLE 20.8 Requirements for a written risk assessment

Requirement	Information included
Outline of investigation	Title, date and location of task
Summary of method	Brief list of steps indicating how the chemicals and equipment will be used
Equipment/chemicals used	List of materials used in the experiment
Equipment/chemical risk and hazards	List of hazards associated with the equipment
Risk control measures	General precautions taken to limit risks

An example of a risk assessment is shown in figure 20.34.

FIGURE 20.34 (a) An example risk assessment (b) GHS pictograms that can be found through risk assessments to visually show any possible hazards

(a)

ACTIVITY	Cross-linking an addition polymer to make slime

SUMMARY OF EXPERIMENT	
AIM	To investigate how the properties of a linear polymer may be altered by the introduction of weak cross-linking between its chains.
METHOD	1. Pour the polyvinyl alcohol into the beaker and add a few drops of the food dye (optional). 2. Add the borax solution and stir with the paddle pop stick. It will take a few minutes for the slime to appear. 3. Perform tests on the product that will enable you to describe its properties and how these are different to the original polymer.

PROTECTIVE MEASURES				
GLASSES	**GLOVES**	**DUST MASK**	**LAB COAT**	**FUME HOOD**
x	x		x	

SAFETY INFORMATION

REACTANT

Polyvinyl alcohol (solution, 6%)

Hazards	Safety precautions
• Flammable • Irritating to the eyes 	• Wear gloves, safety glasses and a lab coat • Keep away from sources of ignition • Use in well-ventilated areas

FIRST AID	
SWALLOWED	Rinse mouth out with water immediately and repeat until all traces are removed. Seek medical attention.
EYE	Flush out with water. Seek medical attention if pain or irritation persists
SKIN	Wash with soap and water
INHALED	Move into fresh air, give oxygen if required. Seek medical attention if breathing if difficult

REACTANT

Borax (sodium tetraborate, solution, 4%)

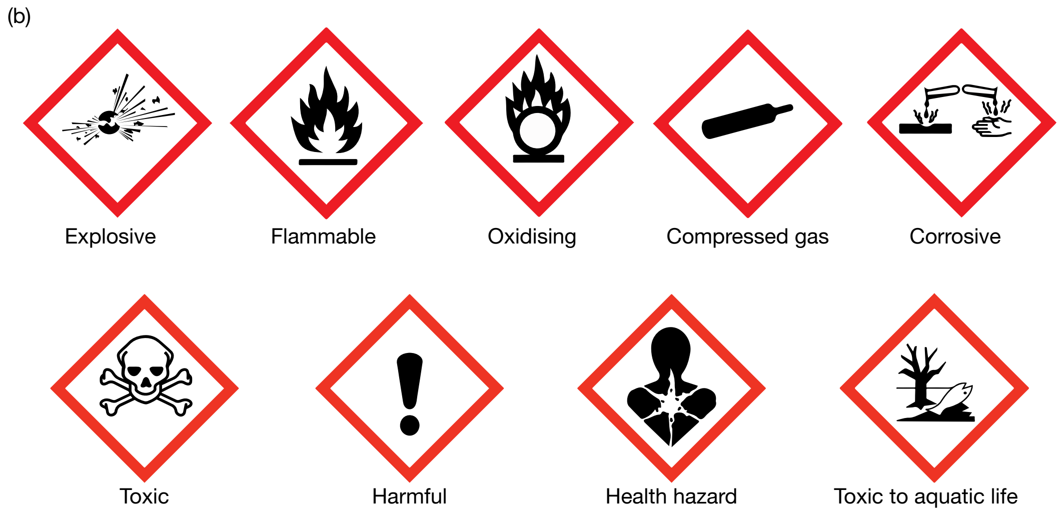
Hazards	Safety precautions
• Not classified as hazardous substance at the concentration used • However, at concentrations above 4.5%, may damage fertility and unborn child 	• Wear gloves, safety glasses and a lab coat • Should not be handled by pregnant women. Those of reproductive age should also avoid the chemical. • Wash hands after use, even if gloves were worn

FIRST AID	
SWALLOWED	Rinse mouth out with water immediately and repeat until all traces are removed. Seek medical attention.
EYE	Flush out with water. Seek medical attention if painor irritation persists
SKIN	Wash with soap and water. Seek medical attention if painor irritation persists
INHALED	Move into fresh air, give oxygen if required. Seek medical attention if breathing if difficult

CONCLUSION

- Wear gloves, glasses and a lab coat for the duration of this experiment
- Make sure that the area is well-ventilated

Signed: ______________________________ **Date:** _______________________



20.5 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Research and identify possible hazards and suggest safety precautions for the following equipment and chemicals.
   (a) 2.0 mol L^{-1} hydrochloric acid
   (b) Burette
   (c) Boiling water
   (d) Thermometer
2. Provide two examples of when ethics may be important in a chemistry investigation.
3. List three purposes of a risk assessment.
4. What other things do you think should be added to a risk assessment?
5. Draw a plan of your laboratory and label the positions of safety equipment.
6. Look around your laboratory and note its safety features and equipment. Then answer the following.
   (a) Does it have any stored pressure fire extinguishers? How are these identified? On what types of fire can these be used and on what types of fire should they not be used?
   (b) Does it have any dry chemical extinguishers? How are these identified? On what types of fire can these be used and on what types of fire should they not be used?
   (c) Where is/are the fire blanket(s) located? Describe a scenario in which a fire blanket would be used and how you would use it.
   (d) Where are the master (emergency) shut-offs for gas and electricity located?
7. The SDS for a chemical to be used in an experiment contains the following risk phrases:
   - Irritating to eyes
   - Skin/flammable/vapours may cause dizziness

   Suggest appropriate methods to reduce these identified risks.

Fully worked solutions and sample responses are available in your digital formats.

20.6 Methods of organising, analysing and evaluating primary data

KEY CONCEPT

- Methods of organising, analysing and evaluating primary data to identify patterns and relationships including identification of sources of error and uncertainty, and of limitations of data and methodologies

KEY SCIENCE SKILLS

- Conduct investigations to collect and record data
- Analyse and evaluate data and methods

20.6.1 Organising primary data

Scientists gather raw data or plain facts from their observations. Data must be recorded at the time in a suitable form — as, for example, text entries, sketches, tables and diagrams in logbooks or in field notebooks. These may be supplemented by audio and video recordings.

FIGURE 20.35 Using logbooks is an easy way to organise data

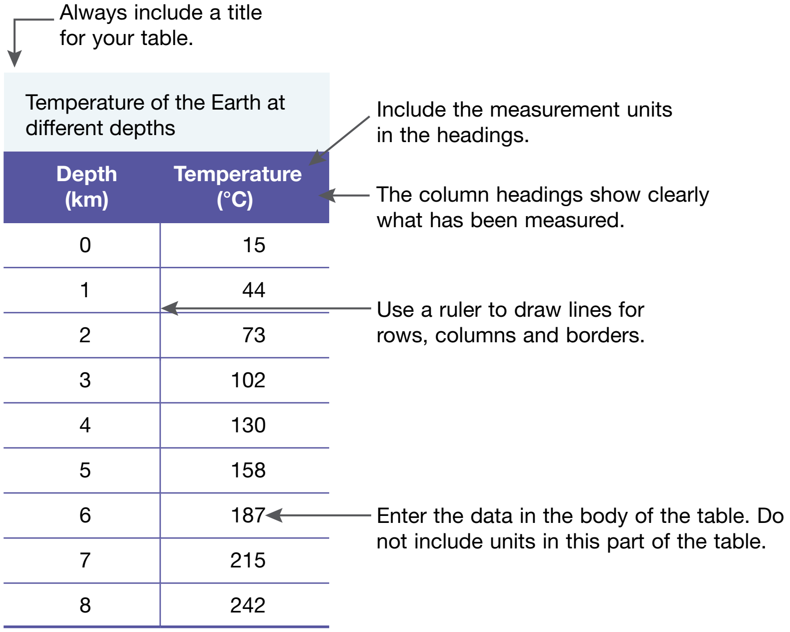
Using a table

Tables should be used when you initially record data. They can help separate and organise your information.

All tables should:

- have a heading
- display the data clearly, with the independent variable in the first column and the dependent variable in later columns
- include units in the column headings and not with every data point
- be designed to be easy to read. If a table becomes too complicated, it is better to break it down into a number of smaller tables.

FIGURE 20.36 Format of a scientific table



Always include a title for your table.

Temperature of the Earth at different depths

Depth (km)	Temperature (°C)
0	15
1	44
2	73
3	102
4	130
5	158
6	187
7	215
8	242

Include the measurement units in the headings.

The column headings show clearly what has been measured.

Use a ruler to draw lines for rows, columns and borders.

Enter the data in the body of the table. Do not include units in this part of the table.

Using a graph

Presenting results as a graph makes it easier to see patterns and trends in your data, allowing more accurate result analysis. While you will usually use a table to record results in your logbook, displaying your data as a graph is preferred on your scientific poster.

When drawing graphs:

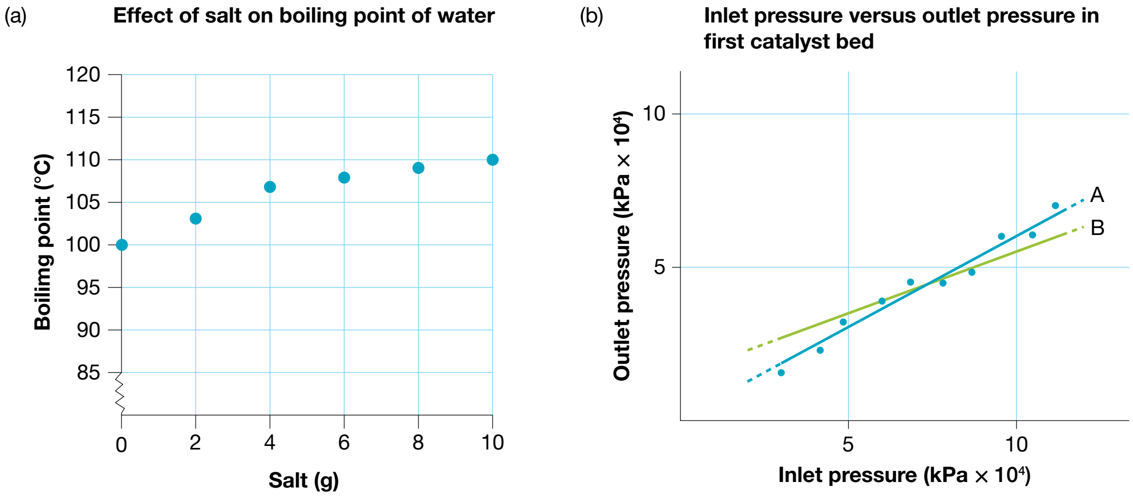
- decide on the type of graph to be used. Different types of information are better suited to different types of graphs. If both the independent and dependent variables are quantitative, a line graph or scatter plot is preferred. Bar graphs are used when one piece of data is qualitative and the other is quantitative. Histograms are used when intervals and frequency are being explored.
- include a title. This should link the dependent and independent variables that are shown in the graph.
- assign axes correctly. In most graph types (excluding pie graphs), the independent variable should be on the horizontal (x) axis, and the dependent variable is on the vertical (y) axis.
- rule axes and label each clearly. Those displaying numerical variables should have a clearly marked scale and units.
- make sure your scale is suitable and the numbers are evenly distributed
- use a line (or curve) of best fit as required. This is a smooth curve or line that passes as close as possible to all the plotted points.
- include the origin, the zero value for the variables, on both axes. You may need to use an axis break symbol if all the values you are plotting are clustered around high values.

The most common graphs you will use in chemistry include scatter plots, bar/column graphs, line graphs and histograms.

Scatterplots

Scatterplots require both sets of data to be numerical. Each dot represents one observation, recorded in regards to the independent and dependent variable. A scatterplot can easily show trends between data sets, and correlations can be identified.

FIGURE 20.37 Examples of (a) a scatterplot graph and (b) a scatterplot graph with a line of best fit



DRAWING A LINE OF BEST FIT

A **line of best fit** can be used to show the general trend of data in a scatterplot graph, and provides a quick summary.

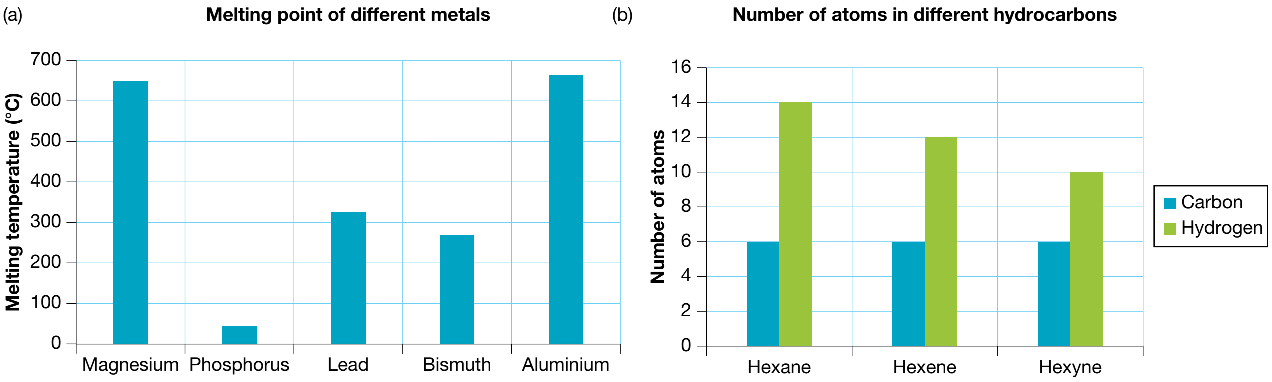
The line of best fit doesn't need to pass through each data point. Although you should try to draw the line through each data point if possible, you may not be able to go through all of them. As a general rule, try to have as many data points above your line as you have below. Don't assume your line must pass through the origin. For example, graph (b) in figure 20.37 shows two possible lines of best fit. Although line A does not pass through any points, it is a better fit than line B. For some data, a curve of best fit may be more appropriate.

Bar graphs

Bar graphs are often used when one piece of data is qualitative and the other is quantitative. The bars are separated from each other. The horizontal or *x*-axis has no scale because it simply shows categories. The vertical or *y*-axis has a scale showing the units of measurements.

Bar graphs can also be used to compare two sets of data by using side-by-side bars, as shown in graph (b) in figure 20.38.

FIGURE 20.38 Examples of (a) a bar graph showing differences in melting points of different metals, and (b) a side-by-side bar chart comparing number of hydrogen and carbon atoms

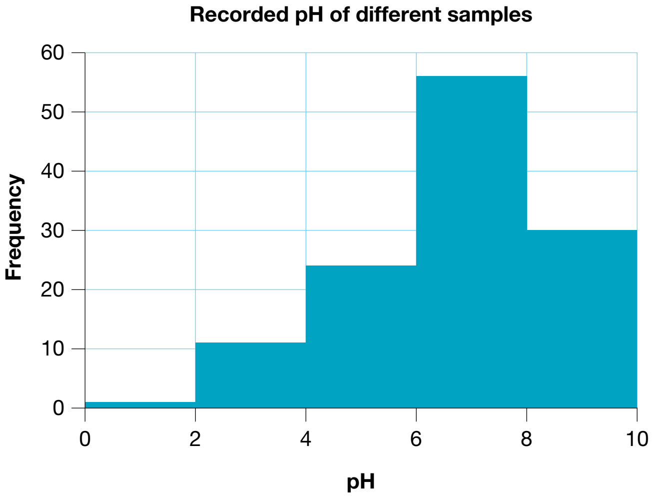


Histograms

A **histogram** is a special kind of bar graph, showing continuous categories. The bars are not separated, unlike in some bar graphs. Histograms are often used when examining frequency.

In figure 20.39, the exact values cannot be determined, because data is displayed in intervals. For example, it can be seen that 30 samples had a pH between 8 and 10. However, we do not know what specific values these are.

FIGURE 20.39 Histogram showing frequency of pH of different samples



Line graphs

In a **line graph**, a series of dots represents the values of a variable, and the dots are joined using a straight line (this is different to a line of best fit, in which the line is straight and does not have to go through each point).

Line graphs are often used to show changes over a continuous period of time, or over space. In particular, line graphs can identify patterns, trends and turning points in a dataset. Line graphs are sometimes curved rather than straight point to point.

FIGURE 20.40 Setting up a line graph

1. Grid

Graphs should always be drawn on grid paper so values are accurately placed. Drawing freehand on lined or plain paper is not accurate enough for most graphs.

2. Title

Tell the reader what the graph is about! The title describes the results of the investigation or the relationship between variables.

3. Setting up and labelling the axes

Graphs represent a relationship between two variables. When choosing which variable to put on each axis, remember that there is usually an independent variable (which the investigator chooses) and a dependent variable. For example, if students wish to find how solubility of sugar changes with temperature, they may choose to measure the solubility (in g/100g) every 20°. The temperature of each measurement was chosen by the students and is the independent variable. The solubility measured is therefore the dependent variable. Usually the independent variable is plotted on the *x*-axis and the dependent variable on the vertical *y*-axis.

After deciding on the variable for each axis, you must clearly label the axes with the variable and its units. The unit is written in brackets after the name of the variable.

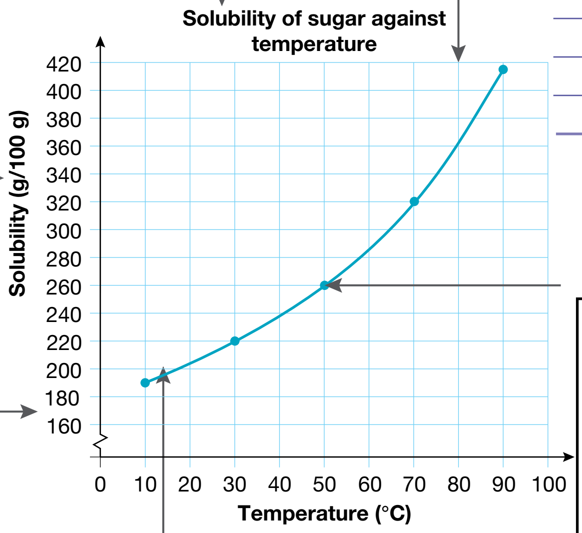
4. Setting up the scales

Each axis should be marked into units that cover the entire range of the measurement. For example, if the solubility ranges from 0 to 415 g/100 g, then 0 and 420 g/100 g could be the lowest and highest values on the vertical scale. If measurements start above zero, as per the solubility measurements on the vertical axis, an axis-break symbol can be used. The distance between the top and bottom values is then broken up into equal divisions and marked. The horizontal axis must also have its own range of values and uniform scale (which does not have to be the same scale as the vertical axis). The most important points about the scales are:

- they must show the entire range of measurements (have an axis-break symbol if required).
- they must be uniform, that is, show equal divisions for equal increases in value.

Data table

Temperature (°C)	Solubility (g/100 g)
10	190
30	220
50	260
70	320
90	415



5. Putting in the values

A point is made for each pair of values (the meeting point of two imaginary lines from each axis). The points should be clearly visible. Include a point for (0, 0) only if you have the data for this point.

6. Drawing the line

A line is then drawn through the points.

A line that follows the general direction of the points is called a 'line of best fit' because it best fits the data. It should be on or as close to as many points as possible. Some points follow the shape of a curve, rather than a straight line. A curved line that either touches all the points or conforms to a best fit can then be used.

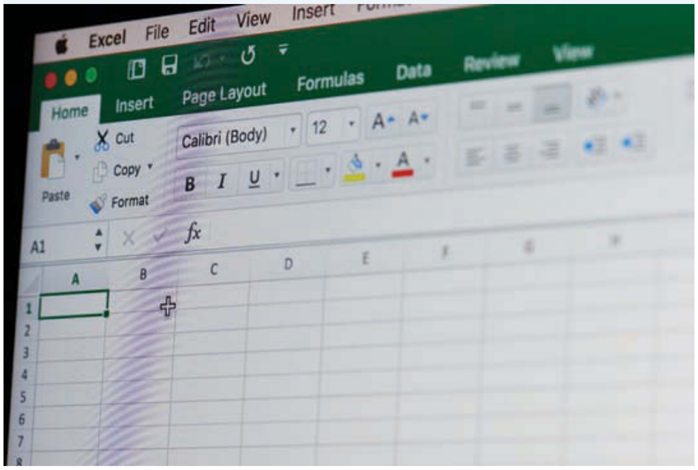
The type of data you graph may lead you to expect either a straight line or a curve. For example, you might expect the increase in temperature of water being boiled to be a straight line because the temperature increases at a steady rate. In this example, the solubility of a sugar with increasing temperature is a curve. Inspecting the data will help you decide whether your line should be straight or smooth and curved.

Using Microsoft Excel

While you may very carefully hand draw your graph, being able to create digital graphs is important, especially for neat presentation on your poster. Microsoft Excel is extremely helpful for this. It can:

- calculate any derived physical quantities, such as speed and acceleration of a parachute or the percentage of energy lost by a bouncing ball. The 'Fill down' command is a time saver.

FIGURE 20.41 Excel is a useful data tool



- be a powerful graphing tool, but must be controlled by the user. You will have to select the graph and choose what aspects of your graph you want to show. For example, what scale on the axis do you want? What do you want to label your axes with? Do you want the data displayed on the graphed points?
- generate a line of best fit. If you right-click on any data point, a window pops up with the option ‘Add Trendline’. This is the Excel command to create a line of best fit. Once selected, you have several choices. If your graph looks like a straight line, choose ‘Linear’.
- create error bars. However, in Excel all error bars are usually the same for each data point, rather than calculated separately.

on Resources

Interactivity Selecting a graph (int-7733)

20.6.2 Analysing primary data

When analysing primary data, it is important to identify and explore trends and patterns. This may include asking questions such as the following:

- Is there a clear positive or negative correlation in data? One variable increasing in response to another increasing shows positive correlation, whereas one variable decreasing in response to another increasing is negative correlation.
- Are there any outliers (unusual data)?
- What results would you expect for specific data that you didn’t observe experimentally?
- Can you calculate the average for your data? Does this differ for the median or the mode?

Analysis of your data often depends on the type of graph selected, because this alters the way that trends and patterns can be seen.

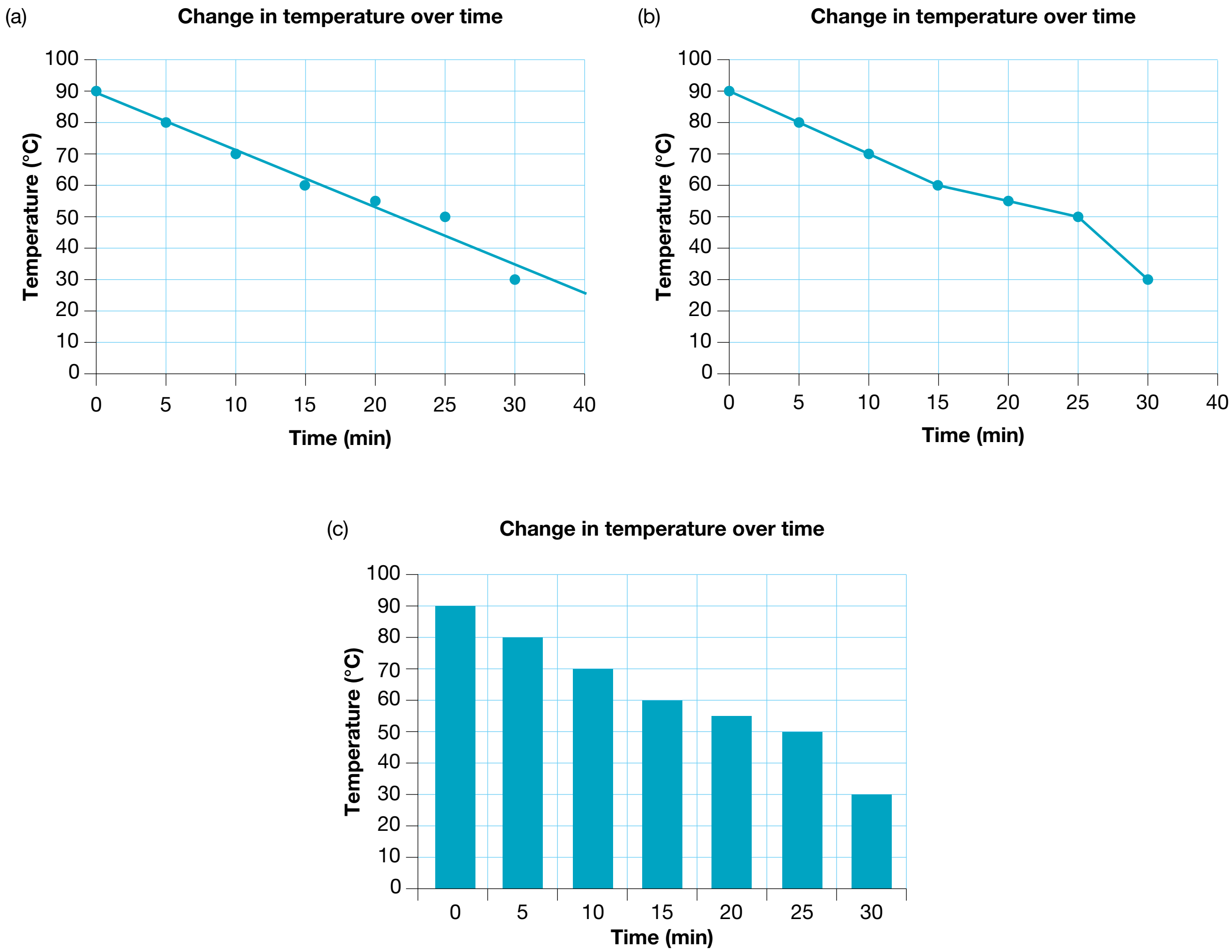
For example, the graphs in figure 20.42 show the same data presented in three ways.

If you were analysing the data from each of the three graphs shown in figure 20.42, the following information may be interpreted:

- Graph (a): a clear downwards trend (negative correlation) can be seen in the data, as the temperature decreases over time. However, at 25 minutes, the temperature can be seen to be slightly higher than expected based on the line of best fit, and the temperature at 30 minutes is slightly lower than expected.
- Graph (b): a clear downwards trend (negative correlation) can be seen in data, as the temperature decreases over time. The rate of temperature drop can be seen to slow after 15 minutes, before the rate increases again between 25 and 30 minutes.
- Graph (c): a clear downwards trend can be seen in the data, as the temperature decreases over time. The temperature can be seen to be lowest at 30 minutes, where it is half the temperature seen at 15 minutes.

The choice of graph is fundamental in any analysis of data. In this case, representation of the data in graphs (a) and (b) is far more powerful than that obtained from graph (c). Regardless, it is important to note down any clear trends and patterns seen in the data, and note down any outliers that can be seen.

FIGURE 20.42 Various graphs showing temperature change over time: (a) a scatterplot graph using a line of best fit, (b) a line graph and (c) a bar graph



Graph analysis can also be used to predict and make assumptions about data that was not gathered experimentally. This can be through interpolation (predicting data points within the data set that were not measured) or extrapolation (predicting data points outside data set based on predicted relationship).

Based on the graphs in figure 20.42 you might estimate:

- the temperature at 35 minutes
- the temperature at 13 minutes
- the time that the temperature was 45 °C.

The estimation you make can vary greatly between all the graphs, so it is important to carefully consider which type you use. Table 20.9 shows the varying results that come from analysing the data in the different graph types in figure 20.42.

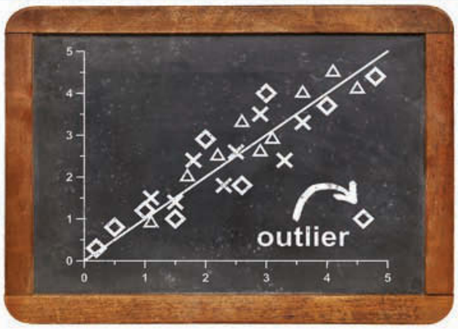
TABLE 20.9 Analysing data from different graph types

	Graph (a)	Graph (b)	Graph (c)
Temperature at 35 minutes	25 °C	15 °C	20 °C
Temperature at 13 minutes	65 °C	62 °C	66 °C
Time the temperature was 45 °C	24 minutes	27 minutes	26 minutes

Outliers

Outliers are results that are a long way from other results and seen as unusual. They should be accounted for and analysed, but are often not included when calculating averages.

FIGURE 20.43 Outliers in data are unusual results.



20.6.3 Evaluating primary data

When you evaluate data, it is important to link back to the initial question. You need to be able to explain and justify your data in relation to your question.
Ask yourself:

- Does the data provide an answer to the question of your investigation?
- Does the data support or refute your hypothesis?
- If any outliers, errors or uncertainty were present in your data, why may these have occurred?
- Does your data link to correctly supported models and theories?
- Could you make further changes to improve your data in future investigations that may reduce errors or limitations?

Two examples of two different students evaluating data are as follows, based on the graphs shown in figure 20.42. Which student do you think has been more clear in their evaluation?

Student 1: It can be seen from the results that temperature decreased.

Student 2: From graph (b), it is clear that temperature decreased over time, dropping from an initial temperature of 90 °C to a final recorded temperature of 30 °C. While there is a clear trend in temperature decreasing over time, the rate of decrease was inconsistent, particularly between 25 and 30 minutes. This may be due to a decrease in external air temperature, causing the rate of heat loss through convection to change.

It is clear that Student 2 provided a more detailed response by describing trends and basing their evaluation on evidence and a link to theory. What else should they add?

FIGURE 20.44 Thermometers can have issues with precision and accuracy.

20.6.4 Sources of uncertainty and error

Understanding sources of **uncertainty** and **error** is vital in scientific reporting, accounting for the fact that the conclusions are being drawn from data that is not perfectly accurate or precise. Uncertainty is acknowledging that no matter how precise an instrument might be, a limit exist to that precision. The uncertainty is a range within which a measurement lies. An error bar is a way of representing that uncertainty graphically.

Errors, on the other hand, are differences between a measurement taken and the true value that is expected. Errors lead to a reduction in the accuracy of the investigation.

Every instrument has a limit to how precisely it measures. The scale or digital display imposes a constraint on how many digits you can record. The scale or display also reveals the tolerance of the measurement.

For example, both errors and uncertainties are associated with using thermometers. Two individuals may look at a thermometer and not get the exact same result, because they might have been standing at different angles or might interpret the reading slightly differently. Or the results may not be identical between two different thermometers. The thermometer also only gives measurements to the accuracy of 1 °C and not decimal values (however, 0.5 °C can usually be seen).

Sources of error

Several sources of error can be identified in an investigation:

- **Systematic errors** are errors that affect the accuracy of a measurement that cannot be improved by repeating an experiment. They are usually due to equipment or system errors. They produce measurements that are consistently too high or too low, as, for example:
  - instrumental errors that arise because an instrument, such as a weighing balance is uncalibrated and incorrectly set to zero
  - environmental errors that arise because of malfunctions, such as a power outage, that affect the conditions under which an experiment is being conducted.
- **Random errors** are chance variations in measurements that affect the precision of measurements and are always present in measurements of continuous data.

  An example of a random error is an error of judgement when reading the smallest division on the scale of a measuring instrument, such as a ruler. Unlike systematic errors, this can usually be resolved with repetition of measurements.

It should be noted that so-called human errors are not errors but are mistakes that result from carelessness. Examples include gross misreading of an instrument, and writing the wrong result in your logbook, such as 40 instead of 4.0.

FIGURE 20.45 Liquids should be measured from the bottom of the dip (meniscus).

Measuring errors

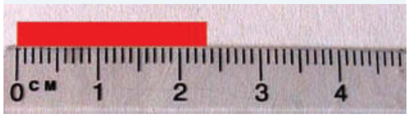
A metre ruler has lines to mark each millimetre, but there is space between these lines. You could measure a length to the nearest millimetre, but because of the space between the lines, if you look carefully, you can measure to a higher precision. You can measure to the nearest 0.5 mm.

The best estimate for the length of the red line shown in figure 20.46 is 2.35 cm. The actual length is closer to 2.35 cm than it is to either 2.30 cm or 2.40 cm. The measurement of 2.35 cm says the actual length is somewhere between 2.325 cm and 2.375 cm.

The way to write this is:

$$\text{The length of the red line} = 2.35 \pm 0.025\,\text{cm}$$

FIGURE 20.46 Errors are also possible when using rulers



The 0.025 represents the tolerance or uncertainty in the measurement (half of the precision that can be recorded).

Reading top-loading balances

Top-loading balances measure to a certain number of significant figures. Say the reading on a digital scale is 8.94 grams. This means the mass is not 8.93 g or 8.95 g. The actual mass is somewhere between 8.935 and 8.945 grams. In this example, the smallest unit of measurement is 0.01 g. Therefore, the tolerance is half of this (0.005 g), because the measurement can be 0.005 g below or above the recorded measurement. The way to write this is:

$$\text{Mass} = 8.94 \pm 0.005\,\text{g}.$$

FIGURE 20.47 Top-loading balance

FIGURE 20.48 What is the tolerance in these measuring instruments?

SAMPLE PROBLEM 4

a. **Record the reading on the scales in figure 20.48 (a), including the tolerance.**

b. **Record the reading on the thermometer in figure 20.48 (b), including the tolerance.**

Teacher-led video: SP4 (tlvd-0104)

THINK	WRITE
a. **1.** Determine the reading on the scale.	128.93
2. Determine the range of the measurement.	The measurement can be between 128.925 and 128.935
3. Determine the tolerance. One way to calculate this tolerance is to halve the smallest unit that can be measured. In this case, 0.01 g is the smallest measurement possible.	$\frac{0.01\text{ g}}{2} = 0.005\text{ g}$
4. Record the reading, including the tolerance.	128.93 ± 0.005 g
b. **1.** Determine the reading on the scale.	47 °C
2. Determine the range of the measurement. In this case, you can tell if the thermometer is a 0.5 increment due to the gaps in the thermometer, so our range is smaller.	The measurement can be between 46.75 and 47.25.
3. Determine the tolerance. The tolerance is half of the smallest unit that can be measured. The smallest unit is 0.5 °C.	$\frac{0.5\text{ °C}}{2} = 0.25\text{ °C}$
4. Record the reading, including the tolerance.	47 ± 0.25 °C

PRACTICE PROBLEM 4

a. Record the reading on scales, including the tolerance, that show a reading of 0.12 grams.

b. Record the reading on scales, including the tolerance, that show a reading of 0.195 grams.

Repeated measurements

Measurements of independent variables are usually precise and careful, so one measurement should be enough. However, measurements of the dependent variables are often prone to some variation.

Variations can be caused by, for example, the human reaction time when using a stopwatch or when judging the rebound height of a basketball. In the case of using a parachute, the unpredictable way the canopy will open each time can cause variations, meaning each reading may be different. So it is sensible to take several readings to obtain an average. You would expect that at least three measurements would be needed, and possibly five, but more than five is generally unnecessary.

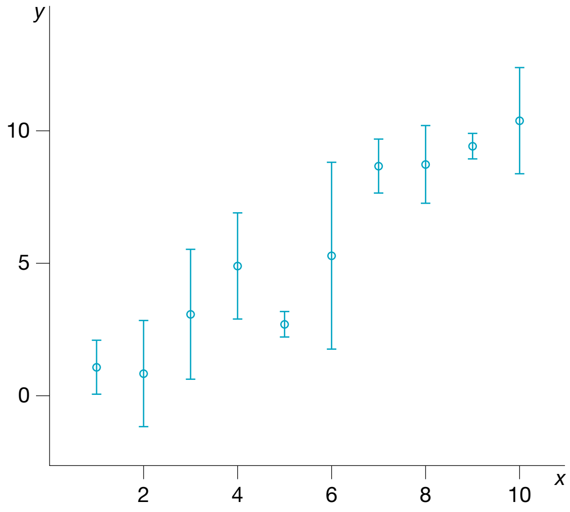
The following example comes from physics, but the graphing techniques equally apply to chemistry. Say your partner dropped a basketball from a height of 80.0 cm, and you judged the rebound height of the ball for five trials as 68 cm, 69.5 cm, 68.5 cm, 68.5 cm and 69.5 cm, the average is 68.8 cm. You would round this to the nearest 0.5 cm because of the difficulty of judging a moving ball, giving an average of 69 cm.

This set of measurements would then be written as 69 ± 1 cm (the 1 cm is the furthest the measurement is from the average).

Graphing error bars

Continuing the basketball rebound example, when you graph your results, the number you will plot is 69 cm. To represent the '±1 cm', you can draw a line through the point, up 1 cm and down 1 cm, with a short line across the top and bottom of the line to make the ends evident. An example of error bars is shown in figure 20.49.

FIGURE 20.49 Error bars



Calculating error bars

Rather than graphing rebound height against drop height, it is more revealing of the physics of the situation to calculate and graph the ratio of the rebound height to drop height against drop height. The ratio is a measure of how much of the original gravitational potential energy is restored.

In this case, the ratio would be $\frac{69}{80.0} = 0.8625$, but how many digits are we entitled to use and how big should the error bar be? The first question is reasonably straightforward. The number of digits in your answer should equal the smallest number of digits in the data you used in the calculation. In this instance, the average height has two digits, so the answer would be written as 0.86. You are not justified in including more digits because you don't know the original data accurately enough.

Working out the size of an error bar takes more effort. If the two pieces of data are 69 ± 1 cm and 80.0 ± 0.3 cm, we can just add the uncertainties to get ± 1.3 cm, but that doesn't make sense when the calculated value is 0.86. Dividing the uncertainties would produce another unusual result.

The method used is to first express the uncertainty for each data value as a percentage. For example:

$$\text{Percentage error of } 69 \pm 1\text{ cm} = \left(\frac{1}{69}\right) \times 100 = 1.4\%$$

$$\text{Percentage error of } 80.0 \pm 0.3\text{ cm} = \left(\frac{0.3}{80}\right) \times 100 = 0.4\%$$

Now add the two percentage errors together.
Total percentage error = 1.4% + 0.4% = 1.8%
Next use this total percentage error to find the error in the calculated answer.
Error = 0.86 × 1.8% = 0.016, which would be rounded to one digit as 0.02.
The full calculated answer would now be 0.86 ± 0.02.

on Resources

🎞 **Video eLesson** Calculating error (eles-2560)

20.6.5 Limitations of data and methodology

The data that is gathered from experimental results will have **limitations**. Limitations arise from several sources that can affect the quality of the data. For example:

- Experiments create artificial situations that do not necessarily represent real-life situations.
- While every effort may be made to identify controlled variables and keep them constant throughout the course of an experiment, it is not always possible to identify and control every one of this type of variable.
- The degree on which results obtained in the laboratory can be generalised to other situations and applied in the real world is limited.

It is important to consider limitations when analysing data and drawing conclusions from data. Limitations should be factors that are out of your control, but need to be discussed in regards to how they might affect your results. As part of your practical investigation, you need to clearly outline limitations and possible suggestions on improvements to the method or data collection to avoid these.

20.6 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

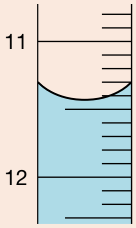
1. Describe the difference between a random and systematic error and provide two examples of each.
2. List two ways that you can minimise uncertainty in an investigation.
3. Identify which graph type would be most appropriate for the following investigations, and justify your choice.
   (a) Comparing the pH of different household liquids
   (b) Showing how pH changes with temperature
   (c) Measuring the temperature inside a car every five minutes for an hour
   (d) Showing the frequencies of different test mark intervals for 400 students

4. Using the data provided in the following table, construct an appropriate graph.

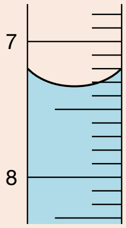
Time (mins)	Volume of ice cube (cm^3)
0	30
5	25
10	21
15	18
20	14
25	10

Once you have constructed the graph, analyse and evaluate the data shown.

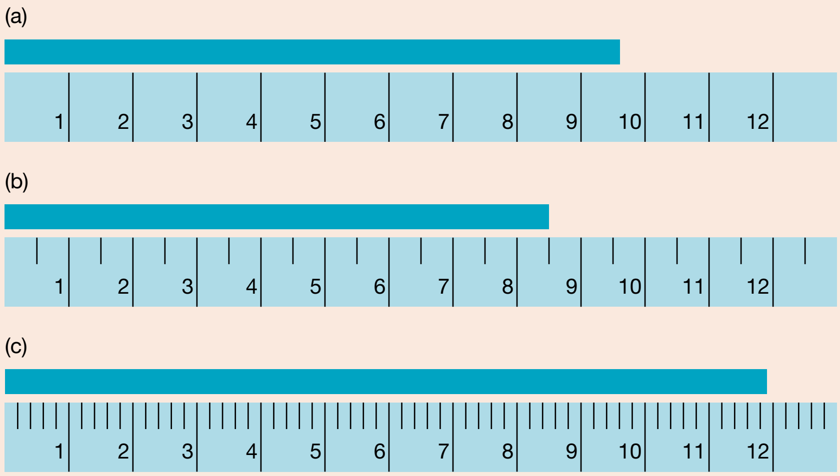
5. Describe how an outlier should be treated in analysing and evaluating data.
6. Record the reading on the burette in the following diagram (remember that the measurements on a burette go downwards, unlike in a measuring cylinder).



7. Explain why a student reading this burette as 8.67 mL is incorrect.



8. Determine the length of each line in the following diagram, showing the tolerance in each case.



9. A student is designing an experiment that involves measuring liquid volumes at various stages. A number of glassware items are available for this purpose, as shown in the following table.

Item	Uncertainty (mL)
volumetric flask (250 mL)	±0.1
beaker (100 mL)	±10
measuring cylinder (100 mL)	±0.1
burette (50 mL)	±0.02
conical flask (250 mL)	±25

Which piece(s) would be most appropriate for the following?

(a) Rinsing a burette with 10 mL of water
(b) Producing 250 mL of a solution of accurately known concentration
(c) Producing 250 mL of a solution of approximately known concentration
(d) Adding water to a conical flask to dissolve a previously weighed tablet, prior to titration

Fully worked solutions and sample responses are available in your digital formats.

20.7 Models and theories to understand observed phenomena

KEY CONCEPT

- Observations and experiments that are consistent with, or challenge, current chemical models or theories

KEY SCIENCE SKILL

- Analyse and evaluate scientific models

20.7.1 Models

As well as data limitations, limitations exist in the models and theories that we use in chemistry. It is important to be able to use models and theories to allow us to understand a variety of observed phenomena. However, we also need to understand limitations in models and theories and the idea that these limitations may change over time.

Models are representations of ideas, phenomena or scientific processes. They can be physical models, mathematical models or conceptual models. Models can provide an explanatory framework that explains observed phenomena, and can help with the understanding of abstract concept. In chemistry, many concepts are hard to visualise, so models help contextualise the idea on a smaller and simpler scale.

While they are very useful, models have their limitations:

- Because of the complexity of the processes, models cannot include all the details of the processes or the things that they represent.
- Models do have some limits in their accuracy and are often simplified and stylised. For example, a ball-and-stick model of methane (see figure 20.50) is useful, but is a very highly simplified and stylised representation that reduces covalent bonds to sticks and atoms to solid balls.
- Models are approximations of the real world; for example, we can model various lattices (as seen in figure 20.51) or bonding, but we often assume ideal conditions around temperature and pressure, which is not realistic about what happens in the real world.

- Models are based off current observations and kodels are constantly being refined as new nowledge at the time. This means that they aren't definite and can change as observations allow for different ideas to come to light.

FIGURE 20.50 Different models of methane, a gas involved the greenhouse effect, have been devised. How do you think these differ from a 'real' methane molecule?

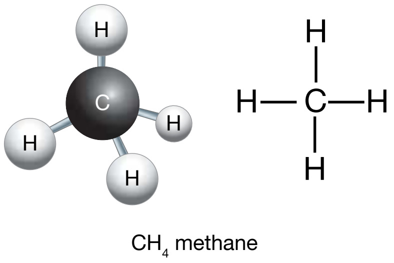
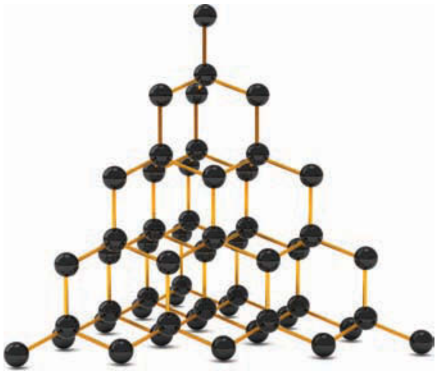


FIGURE 20.51 Model showing the lattice of diamond



Models are constantly being refined as new information becomes available. One model that has continuously changed is that of the atom (see figure 20.52). These changes are described further in table 20.10.

FIGURE 20.52 Various models of the atom, including John Dalton's model and the 'plum pudding' model by J.J. Thomson

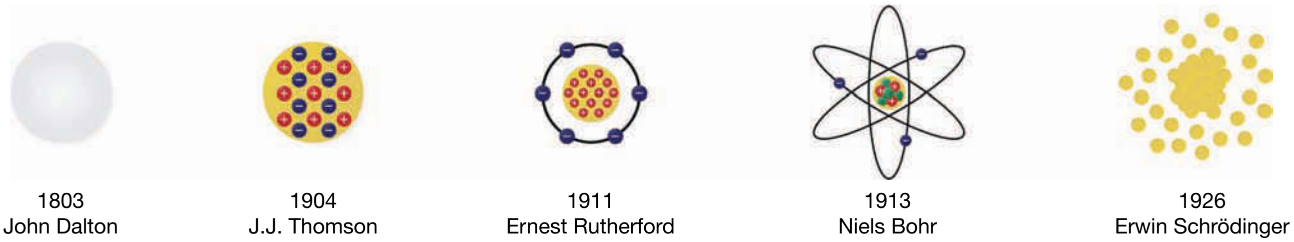


TABLE 20.10 Models are challenged and adjusted

Model	What was the model?	What observations challenged this?
Solid sphere model (John Dalton)	Atoms from different elements are different to each other. Atoms are indivisible and not made of smaller components.	The finding that atoms are made of subatomic particles known as protons, neutrons and electrons.
Plum pudding model (J.J. Thomson)	Negatively charged electrons float around in a large section of positive charge.	The gold foil experiment: when positively charged particles were fired at a very thin layer of gold, only some deflected back, showing the atom is mostly empty space.

Nuclear model (Ernest Rutherford)	The atom is mostly empty space, with a positively charged nucleus in the centre.	Moving electrons: the model didn't explain their orbit around the nucleus.
Planetary model (Niels Bohr)	Electrons move around the nucleus in orbits of fixed size and energy.	Heavier atoms: these would collapse with moving electrons and energy.
Quantum model (Erwin Schrödinger)	Electrons do not move in set orbits, but instead move in waves, or clouds known as orbitals.	Currently supported model

As seen in table 20.10, new observations and experiments often lead to the development and changing of previously supported models. In the future, more discoveries could come to light that change models that we currently believe to be accurate. This shows how important it is to make sure that any research you conduct is current and relevant.

20.7.2 Theories

A **theory** is a well-supported explanation of a phenomenon. It is based on the interpretation of facts that have been obtained through investigations, research and observations.

Limitations can exist with theories. These may include:

- A reliance on theories rather than observations made during practical investigations. Often, individuals manipulate results to match a theory, failing to realise the power of observations. Often, theories are treated as perfect when, in fact, they are able to be disproven as further observations and evidence come to light.
- Theories often rely on a very specific set of conditions to be met. For example, the Law of Conservation of Mass cannot be used in systems that are open systems (not isolated) or systems that involve very large amounts of energy.
- Often, a phenomenon needs to be described using multiple theories. Sometimes, aspects of theories may have some contradictions to each other.

Theories can change overnight or take a very long time to change. Theories that were once popular and well accepted may be discarded when too much evidence builds up against them. They are replaced by a theory which better fits the observations.

Some examples of theories that have been superseded by other theories are:

- miasma theory of disease (the idea that diseases were caused by bad air)
- the theory that the sun and planets orbit around the Earth (geocentrism)
- the flat Earth theory.

In the geocentric model of the Earth, the idea that the planets and sun rotated around the Earth was the prominent theory for a long time, supported by many philosophers and statements in the Bible. The heliocentric theory was proposed quite early in history, but only gained traction when a mathematical model of this system was showed by Nicolaus Copernicus. Individuals like Kepler developed this further with the idea that orbits were elliptical and not spherical. These ideas were further shown and developed by Galileo, through the use of a telescope.

FIGURE 20.53 The flat Earth theory has been replaced with the idea that the Earth is spherical

The flat Earth theory argued that the Earth was a plane or a disc. Hundreds of years ago, this was thought to be accurate. The idea that Earth was a sphere was suggested by Pythagoras, and later evidenced by Aristotle. This theory then replaced the flat Earth theory.

How do refinements of existing theory come about?

Refinements to existing theories come about through a number of ways:

- carefully planned laboratory-based or field-based experiments designed to support or refute a particular hypothesis
- critical reinterpretation of previously accepted facts, producing a new framework
- collection and analysis of new data
- identification and exploration of patterns or anomalies
- new technologies that allow for changes to understanding and more depth of knowledge.

20.7 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Describe three ways in which models can be useful.
2. Describe three limitations of models.
3. What is the difference between a model and a theory?
4. How can theories help us understand phenomena?

Fully worked solutions and sample responses are available in your digital formats.

20.8 Nature of evidence and key findings of investigations

KEY CONCEPTS

- The nature of evidence that supports or refutes a hypothesis, model or theory
- The key findings of the selected investigation and their relationship to solubility, concentration, acid–base and/or redox concepts

KEY SCIENCE SKILLS

- Analyse and evaluate data, methods and scientific models
- Draw evidence-based conclusions
- Communicate and explain scientific ideas

20.8.1 Nature of evidence: supporting or refuting a hypothesis, model or theory

Evaluating a hypothesis, model or theory uses the evidence from the results obtained in an investigation. In your investigation, it is important to understand how to best use evidence in order to correctly support or refute your hypothesis.

Evidence can be strong or weak, and it is vital that the evidence you choose allows the provision of strong and clear conclusions.

Strong evidence should:

- be based on facts derived from studies with high validity and minimal bias
- use data to support conclusions
- clearly link to the aim and hypothesis of an experiment
- be obtained from an investigation that has a reproducible and reliable method.

If the prediction from your hypothesis was validated by your experimental results, you should evaluate your hypothesis as ‘supported’; if your hypothesis was not supported by results, it is termed as ‘rejected’ or ‘not supported’.

Remember: a scientific hypothesis can be rejected (or not supported), but it can *never be proven true*; it can only be *supported*. This is because the nature of evidence is that, as new technologies and information becomes available, evidence can change and be interpreted in different ways. This may than disprove a previously supported hypothesis.

TABLE 20.11 Examples of results that can support or disprove a hypothesis

Observation	Hypothesis	Test of hypothesis	Result	Conclusion
The phone is not charging	The charger is faulty.	Use a different charger	The phone still does not charge.	Hypothesis not supported
The measurement on the scales is higher than expected	The scales were not set to zero before use.	Reset the scales back to zero	The reading is closer to the expected value.	Hypothesis supported
The water did not freeze at 0 °C	There was a contaminant in the water.	Freeze water that is pure (distilled)	The water froze at 0 °C.	Hypothesis supported

20.8.2 The key findings of investigations

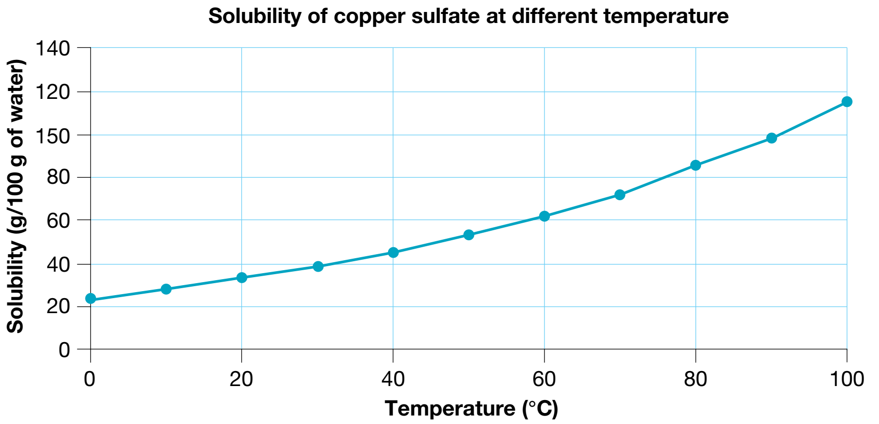
In your investigation, you will need to write a question and a hypothesis, create a clear reproducible methodology, conduct an experiment and collect and analyse data.

Once these steps are complete, it is then important to determine the key findings of your investigation. Your key findings should include:

- information about the data obtained in the practical investigation and any patterns and trends
- relationship of your findings to chemical concepts, including solubility, concentration, acid–base and/or redox concepts
- an answer to the question of your investigation.

For example, your question may have been, ‘How does temperature affect the solubility of copper sulfate?’ Figure 20.54 shows possible data from your experiment.

FIGURE 20.54 Solubility of copper sulfate



You could use the graph shown in figure 20.54 to link this data to the concept of solubility and determine your key findings. For example:

> *Copper sulfate,* $CuSO_4$*, is soluble in water, allowing for the formation of an aqueous solution. The expected solubility at 20 °C is 20.5 g/100 g water.*

In the conducted experiment, it can be seen that at 20 °C, the solubility was recorded at 33 g/100 g water, which is significantly higher than expected. The solution produced was likely to be supersaturated and not saturated as expected.

It is also clear that as temperature was increasing, the solubility of copper sulfate also increased. This relationship can be linked to the particle model, as the increase of kinetic energy from higher temperatures causes particles to spread out and move faster. As well as that, higher temperatures lead to an increase in intermolecular bonds breaking, causing the copper sulfate to more easily be dissolved in 100 grams of water.

From the investigation, it can be observed that the temperature has a significant impact on solubility as observed in the graph shown; however, this relationship is not linear. This supports the hypothesis of this investigation.

20.8 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Why do we only say that we ‘support’ a hypothesis rather than ‘prove’ it?
2. Provide three examples of strong evidence.
3. Describe what key findings you should communicate in a discussion.
4. Why is it important to show the relationship of your results to chemical concepts such as solubility and concentration?

Fully worked solutions and sample responses are available in your digital formats.

20.9 Options and issues related to water quality

KEY CONCEPT

- Options, strategies or solutions to issues related to water quality

KEY SCIENCE SKILLS

- Plan and undertake investigations
- Conduct investigations to collect and record data
- Analyse and evaluate data, methods and scientific models
- Communicate and explain scientific ideas

20.9.1 Monitoring water for contaminants

As examined in Unit 2, water is the most widely used solvent on Earth, vital for survival and life as we know it. As part of Unit 2, Area of Study 3, your investigation is involved in exploring a specific aspect of water quality.

The water that we drink and use is not pure, and contains many contaminants, and some of these can adversely affect our health. Therefore, regularly examining water quality to make sure that it is safe for consumption is vital.

Water may be contaminated by substances such as:

- heavy metals (including lead, chromium and arsenic)
- organic pollutants (including fertilisers, pesticides and pharmaceuticals)
- microorganisms (including bacteria, algae and parasites).

It is important to monitor water for factors, such as:

- salts (refer to topic 17)
- organic molecules (refer to topic 18)
- acids and bases (refer to topic 19).

20.9.2 Key factors when monitoring water quality

Options, strategies and solutions to issues related to water quality were examined in topics 15 through to 19, so should be explored and reviewed to assist with your investigation.

Some key factors to remember when conducting to your practical investigation are as follows:

- Avoid sampling bias when collecting water samples, through making sure that the sample is representative of all water in the area. For example, if you collect water only from the surface, the contaminants may be quite different to those further down. Try to collect multiple samples or make sure that the sample is well mixed (below 100 mm from the surface).
- Make sure that your water can be stored over multiple days of the experiment. If you have to continue to collect new samples, the quality of the water may have changed.
- Use repetition to improve the accuracy and precision of your investigation.
- Make sure you choose the most appropriate technique to sample and analyse your water. Your choice will often depend on the properties of the material under investigation. Consider time and cost effectiveness, safety and the availability of resources in your school. Some examples of possible analysis techniques are shown in table 20.12.

TABLE 20.12 Water analysis techniques

	Techniques	Notes on technique
Analysing salts	Gravimetric analysis	Useful for metal ions and anions
	Instrumental colorimetry	Useful to determine concentration by detecting light
	UV-visible spectrometry	Suitable for coloured solutions or those that absorb in the ultraviolet region
	Atomic absorption spectrometry	Useful to measure concentrations of metal ions
Analysing organic molecules	High performance liquid chromatography	Highly sensitive technique to measure organic molecules*
Analysing acids and bases	Acid–base titrations	Measures concentration based on neutralisation reactions
	Indicators	Quick and easy technique to measure pH, but offers reduced precision
	Potentiometric titration	Uses two electrodes and measures the changes in voltage
	pH meters	Easy for quick measurements

*It is unlikely that your school has access to HPLC. Access to specialised equipment should be determined before choosing this technique.

It is vital to be aware of limitations with the technique or techniques that you select, and take appropriate measures to reduce uncertainties and errors that may occur and which will impact the accuracy and precision of your results.

Finally, as part of your investigation, considering strategies for improving water quality is vital. What are some ways that we can improve water quality? How might we be able to ensure more individuals have access to clean drinking water free of contaminants? These aspects are an important part of Area of Study 3, and looking at them will allow you to receive the best results possible because it shows your understanding of water quality.

20.9 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Why is it important to test water?
2. Provide some examples of the following.
   (a) Heavy metals that may contaminate water
   (b) Microorganisms that may contaminate water
   (c) Organic pollutants that may contaminate water
3. Would it matter what technique you used to detect salts or would all techniques lead to the same results?

Fully worked solutions and sample responses are available in your digital formats.

20.10 Conventions of scientific report writing and scientific poster presentation

KEY CONCEPT

- The conventions of scientific report writing including chemical terminology and representations, symbols, chemical equations, formulas, units of measurement, significant figures and standard abbreviations

KEY SCIENCE SKILL

- Communicate and explain scientific ideas

20.10.1 Conventions of report writing

It is vital in scientific report writing and scientific poster presentation to follow the protocols of scientific written communication (see table 20.13).

TABLE 20.13 Aspects of a written report

Section	Description
Title	A precise and complete description of what you investigated
Abstract	An optional section used to outline the key findings and information
Introduction	A paragraph explaining the relevant chemical and background concepts and relationships, and how they apply to this investigation. It should explore any prior investigations conducted on this topic, and should also include a clear aim and hypothesis.
Methodology	A detailed section that describes your selection of equipment and measuring instruments, and your step-by-step method. This may include diagrams and photos. You should refer to how you controlled variables, achieved the desired accuracy, and overcame, avoided or anticipated difficulties. The methodology explained should be clear enough to enable someone else (at your level) to repeat your experiment. Do not forget to highlight how the relevant ethical and safety concerns have been addressed.
Results	A clear representation of your results, including your data and graphs. If you have too much data, refer to your logbook for the full set. Make sure you present your results in an organised and clear manner, following accepted conventions (such as numbered sequentially). Make sure you include the appropriate units and use the correct number of significant figures. Try to organise your results in such a way that any patterns or relationships start to become obvious, thus making it easier to analyse in the next section. However, the purpose of this section is to present results only; no interpretation of data should be included in this section. Show sample calculations if required.

Discussion	A detailed analysis and evaluation of your results. How does your data support your initial intentions and link to relevant chemistry concepts? What trends and relationships are apparent as a result of your investigation? How much is your analysis limited by uncertainties? Did any outliers appear in your investigation and how were they treated? What limitations and sources of error were present in your investigations and how would you improve this in future repetitions? What would your next steps in the investigation be if you had more time?
Conclusion	This should relate to the aim and must be based entirely on the evidence obtained in the experiment. It should state whether the hypothesis is supported, summarising the meaning of your results in response to your question. No new information should be included.
References and acknowledgements	You should quote the sources of any content that you include that is not your own original work. Unaltered tables, diagrams and graphs are examples that fit this description, as are direct quotes. In your introduction, you may have mentioned previous work that your investigation is based on. This also needs to be acknowledged, along with any sources that inform your discussion of concepts and theory in a more general sense. This section is not counted in your final word count.

It is important to remember that your assessed practical investigation is based on both your scientific report (or poster) and your logbook. The information you will be marked on may be in either your report (or poster) or your logbook, or it may appear in both. It is important that all information is presented.

Some other things that are important to a scientific report (and your poster) are:

- Try to avoid subjective language; where possible, use third-person language.
- Don't just record the data you believe supports your hypothesis — you should also include any errors, uncertainties and outliers.
- If you used any calculations, show your workings.
- Use subheadings throughout your report as shown in table 20.13. Make the report clear to read.
- Provide headings for all graphs, tables and figures and label them sequentially (graph 1, graph 2 and so on).
- Ensure you spell-check your report as the last task before submitting.

20.10.2 Terminology and representations

It is important throughout your report to use clear and concise terminology relevant to the related chemistry concept. The use of key terms was covered earlier in this topic (see section 20.4.2).

Furthermore, chemical representations are required to be accurate and use common conventions. These representations were discussed in section 20.4.3. Other representations used may be appropriate to your specific investigation. These should be used consistently in both your logbook and your poster.

20.10.3 Symbols

Due to the sheer quantity of variables we have to represent, it is important to note that the capital and lowercase versions of a letter usually represent different things. As well as that, sometimes the same symbol is used to represent different variables.

For example, μ can be used to represent:

- mean in statistics
- micro in measurement
- coefficient of friction.

And C can be used to represent:

- carbon (C)
- concentration (c).

It is important to carefully use the correct symbol in your report to minimise confusion. Table 20.14 shows some symbols used in Unit 1 and 2, it is important to know these and other symbols, including those used in equations.

TABLE 20.14 Some common symbols used in chemistry

Symbol	Representation
Δ	Change in variable
$\rightleftharpoons$	Equilibrium arrow
N_A	Avogadro's constant
N	Number of particles
n	Amount in mol
M	Molar mass
m	Mass
c	Concentration
V	Volume
V_m	Molar volume
e^-	Electron
A_r	Relative atomic mass
K_W	Ionisation constant
(g)	Gas (used in chemical formulas)
(aq)	Aqueous (used in chemical formulas)
(l)	Liquid (used in chemical formulas)
(s)	Solid (used in chemical formulas)

Symbols are often used to represent specific variables, and different elements. Many symbols are also specific to drawing the skeletal structure of molecules.

20.10.4 Equations and formulas

When using equations and formulas in your scientific investigation, it is important to:

- define all variables
- provide any figures for constants (for example, Avogadro's number).

Some key formulas in chemistry are included here, and many can be found in your VCE Chemistry Data Book. You should become very familiar with this data book prior to your exam because it contains a huge amount of useful information you do not need to memorise.

Quantifying chemistry

- Relative atomic mass: $A_r = \dfrac{\begin{array}{c}\text{(relative isotopic mass of first isotope} \times \text{abundance)} + \\ \text{(relative isotopic mass of second isotope} \times \text{abundance)} + \ldots\end{array}}{100}$

- $n\text{ (number of moles)} = \dfrac{\text{mass}}{\text{molar mass}} = \dfrac{m}{M}$
- $n\text{ (number of moles)} = \dfrac{\text{number of particles}}{6.02 \times 10^{23}} = \dfrac{N}{N_A}$
- Molecular formula $= n \times$ empirical formula

pH

- $K_w = [H_3O^+][OH^-] = 10^{-14}$
- $\text{pH} = -\log_{10}[H_3O^+]$

Measuring solubility and concentration

- Concentration (c) in grams per litre (g L^{-1}) $= \dfrac{\text{mass of solute in grams}}{\text{volume } (V) \text{ of solution in litres}} = \dfrac{m}{V}$
- Parts per million (ppm) $= \dfrac{\text{mass of solute in micrograms}}{\text{mass of solution in grams}}$
- Concentration %(m/m) $= \dfrac{\text{mass of solute in grams}}{\text{mass of solution in grams}} \times 100\%$
- Concentration %(m/v) $= \dfrac{\text{mass of solute in grams}}{\text{volume of solution in mL}} \times 100\%$
- Concentration %(v/v) $= \dfrac{\text{volume of solute in mL}}{\text{volume of solution in mL}} \times 100\%$
- Molar concentration (c) $= \dfrac{\text{quantity of solute (moles)}}{\text{volume of solution (litres)}} = \dfrac{n}{V}$

20.10.5 Units of measurements

In chemistry, it is vital to use the correct unit of measurement for accurate and clear scientific communication.

Prefixes

Table 20.15 shows the prefixes for units of measurement. These values are from a base unit, which does not have a prefix. Examples of base units include metre, seconds or grams. These base units have the value of 10^0 (or 1). All other prefixes are compared to this base unit using 10^n, in which n varies as shown in table 20.15.

TABLE 20.15 Prefixes for units of measurement, where the base unit of measurement is the metre

Pico (p)	Nano (n)	Micro (μ)	Milli (m)	Centi (c)	Deci (d)	Kilo (k)	Mega (M)	Giga (G)
10^{-12}	10^{-9}	10^{-6}	10^{-3}	10^{-2}	10^{-1}	10^{3}	10^{6}	10^{9}

Understanding the different prefixes allows the correct units to be used in practical investigations, and also allows for easy conversion between different units.

The following formula is used when converting between units:

$$\frac{\text{initial unit } (10^a)}{\text{new unit } (10^b)} \times \text{value}$$

SAMPLE PROBLEM 5

Convert:

a. 12.412 millilitres to microlitres

b. 26153 milligram to decigram

c. 8.7 metres to nanometres.

Teacher-led video: SP5 (tlvd-0105)

THINK	WRITE
a. 1. Determine the conversion between the units. $\frac{\text{millilitres}}{\text{microlitres}}$	$\frac{10^{-3}}{10^{-6}} = 10^3$
2. Multiply this by the value to be converted.	$10^3 \times 12.412\text{ mL} = 12412$
3. Add the new unit.	$12412\ \mu\text{L} = 1.2412 \times 10^4 \mu\text{L}$
b. 1. Determine the conversion between the units. $\frac{\text{milligram}}{\text{decigram}}$	$\frac{10^{-3}}{10^{-1}} = 10^{-2}$
2. Multiply this by the value to be converted.	$10^{-2} \times 26153\text{ mg} = 261.53$
3. Add the new unit	$261.53\text{ dg} = 2.6153 \times 10^2\text{ dg}$
c. 1. Determine the conversion between the units. $\frac{\text{metre}}{\text{nanometre}}$ *Note:* 10^0 is used for metres because it is our standard unit, so is equal to 1.	$\frac{10^0}{10^{-9}} = 10^9$
2. Multiply this by the value to be converted.	$10^9 \times 8.7\text{ m} = 8\,700\,000\,000$
3. Add the new unit (round if required).	$8\,700\,000\,000\text{ nm} = 8.7 \times 10^9\text{ mn}$

PRACTICE PROBLEM 5

Convert:

a. 7823 decigrams into kilograms

b. 213 microlitres into picolitres.

SI units of measurement

SI units or Système Internationale is our metric system of measurements. This system is internationally standardised.

TABLE 20.16 Common SI units used in chemistry

Quantity	Unit	Symbol
Length	Metre	m
Mass	Kilogram	kg
Time	Second	s

Temperature	Kelvin	K
Density	Ampere	A
Amount of substance	Mole	mol
Energy	Electron volt	eV

Derived units

Derived units are units of measurements derived from the SI units. Table 20.17 shows some commonly used derived SI units.

Speed is an example of a quantity that is measured in derived SI units. The SI unit of speed is the metre per second, written as m/s or m s^{-1}.

TABLE 20.17 Common derived SI units used in chemistry

Quantity	Unit	Symbol	Unit in terms of other units
Force	Newton	N	Kg m s^{-2}
Energy and work	Joule	J	N m
Pressure	Pascal	Pa	N m^{-2}
Temperature	Degrees Celsius	°C	+ 273.15 K
Density		*d*	kg m^{-3}
Radiation dose	Sievert	Sv	J kg^{-1}

20.10.6 Significant figures

A degree of uncertainty exists in any physical measurement. The uncertainty can be due to human error or to the limitations of the measuring instrument.

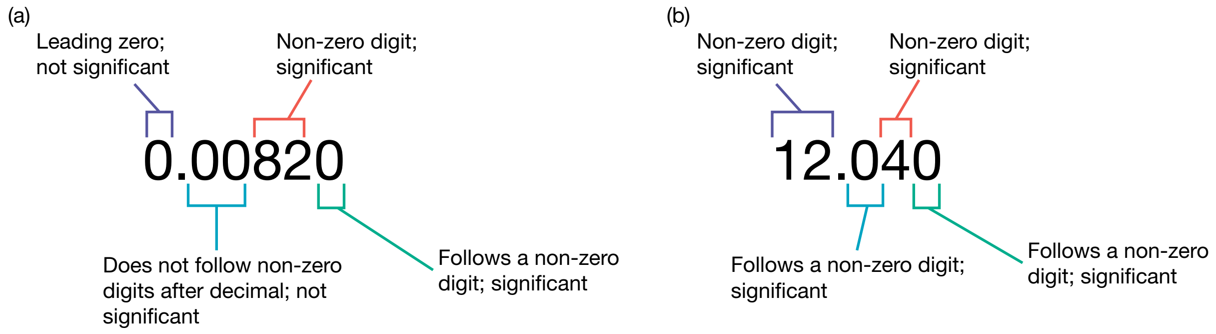
In most physical measurements, the last significant figure shows a small degree of uncertainty. For example, the length of an Olympic competition swimming pool is correctly expressed as 50.00 m. The last zero has a small degree of uncertainty.

Complicated by zeros

Two simple rules can be used to help you decide if zeros are significant.

- Zeros before the decimal point are significant if they are between non-zero digits. For example, all of the zeros in the numbers 4506, 27 034 and 602 007 are significant. The numbers, therefore, have four, five and six significant figures respectively. The zero in the number 0.56 is not significant.
- Zeros after the decimal point are significant if they follow a non-zero digit. For example, in the number 28.00, the two zeros are significant. The number has four significant figures. However, in the number 0.0028, none of the zeros (the one before the decimal point and the two after the decimal point) is significant. They do not follow a non-zero digit and are present only to indicate the position of the decimal point. This number therefore has only two significant figures. The number 0.00280 has three significant figures (the zero at the end follows a non-zero digit, so is significant).

FIGURE 20.55 Examples of significant figures: (a) three significant figures and (b) five significant figures



Working with significant figures

When multiplying or dividing, the answer is written to the least number of significant digits.

For example, if you knew a substance had a mass of 7.6 g and a molar mass of 18.5 g mol^{-1}, the amount in mol would be given by:

$$\begin{aligned}\text{mol} &= \frac{\text{mass}}{\text{molar mass}}\\ &= \frac{7.6}{18.5}\\ &= 0.4108108\ \text{mol}\\ &= 0.41\ \text{mol}\end{aligned}$$

We can be confident of the result up to two significant figures as this is the least number of significant figures on which the calculation was based. Hence, the result should be rounded to two significant figures.

When quantities are added or subtracted, the result should be expressed to the minimum number of decimal places used in the data. For example, if you measured three consecutive volumes of 23.4 mL, 24.63 mL and 20.123 mL, the total volume would be given by:

$$\begin{aligned}23.4 + 24.63 + 20.123 &= 68.153\\ &= 68.2\ \text{mL}\end{aligned}$$

The result should be rounded off to one decimal place because the minimum number of decimal places used in the data is one, in the volume of 23.4 mL.

on Resources

Video eLesson Determining significant figures (eles-2559)

SAMPLE PROBLEM 6

In determining the density of a particular liquid, a student measured the volume of a sample as 8.3 mL. She then weighed the same sample and obtained a mass of 7.2136 g. Calculate the density to the correct level of significant figures.

Teacher-led video: SP6 (tlvd-0106)

THINK	WRITE
1. Determine the number of significant figures provided.	7.2136 = five significant figures 8.3 = two significant figures
2. Determine the least number of significant figures — this is what your answer will be in.	Two significant figures.
3. Calculate the density by recalling the formula: $\text{density} = \frac{\text{mass}}{\text{volume}}$	$\text{Density} = \frac{7.2136}{8.3} = 0.8691$
4. Round down to the appropriate number of significant figures. In this case, 0.8691 needs to be rounded down to two significant figures.	0.87 g mol^{-1}

PRACTICE PROBLEM 6

In determining the density of a particular liquid, a student measured a sample's volume as 21.1 mL. She then weighed the same sample and obtained a mass of 9.762 g. Calculate the density to the correct level of precision.

20.10.7 Standard abbreviations

Often, it is appropriate to abbreviate terms rather than writing them out in full on each appearance, which can make a report bulky and hard to follow.

In chemistry, we often shorten formulas to their formula instead of writing it out in full. This is particularly true with molecules such as water, in which the abbreviation H_2O is universally known.

Some common abbreviations used in Unit 1 and 2 chemistry are:

- IUPAC: International Union of Pure and Applied Chemistry
- pH: potential of hydrogen
- UV: ultraviolet
- AAS: atomic absorption spectroscopy
- HPLC: high performance liquid chromatography
- ppm: parts per million
- ppb: parts per billion.

When using an abbreviation, write the word out in full on the first appearance followed by its abbreviation in brackets; then, in subsequent uses, the abbreviation alone may be used.

An example of this is as follows:

> *Atomic absorption spectroscopy (AAS) was developed in Australia. In AAS, the concentrations of metal ions are detected.*

20.10.8 Acknowledgement of references

A detailed scientific report requires a depth of research for concepts relating to an investigation. This research may involve:

- using other sources for definitions and background material
- finding examples of similar investigations
- researching the obtained results further to link these results to scientific understanding.

If you use any material that is the work of another person, you *must* acknowledge its source. Do not claim it as your own work. Acknowledgments come in two formats: a short version when it occurs in the body of your poster, and a longer version when it occurs in the *Reference and Acknowledgments* section at the end of your poster or scientific investigation.

You can make such acknowledgments in many ways, and various institutions and publications often use different systems. Details of these systems can be found online and can be quite complicated. You should check with your teacher as to how your referencing is expected to be done.

Acknowledging sources within your report: in-text referencing

An in-text reference is a shortened or abbreviated form of a reference and should be used in the body of your report in the location in which the sourced information is referred to. This is used for not only direct quotes, but also tables, images and any information that has been paraphrased.

You can do this numerous ways and it depends on what style you are using so, again, check with your teacher.

Author–date system

The author–date system lists the last name of the author and the year of publication. This style of in-text referencing is more commonly used, particularly in the APA and Harvard styles of referencing. As well as the shortened in-text reference, a full reference is included in the reference list.

The in-text reference appears directly after the end of the information being used. This may mean the reference appears in the middle of a sentence. For example:

> *. . . over the past 10 years, the number of eligible children has increased (Kringle, 2008) and a need has, therefore, developed for sleighs to travel faster to meet the required delivery schedule. More efficient fuels are required for this purpose.*

Items such as tables, diagrams and graphs that are being inserted without being substantially altered can often be acknowledged by stating the details directly underneath them. The reference at the end of table 20.18 is an example.

Note the following for the author–date system:

- If the online article is undated, put (n. d.) in place of the date.
- If no author is listed, use the title in place of the author's name.
- If the source has up to three authors, list them all.
- If the source has more than three authors, only use the name of the first author and follow it by the phrase 'et al.' (meaning 'and others').
- If you quote directly from an author or want to cite a specific idea or piece of information from the source, also include the page number of the quote in your in-text reference (see the example in table 20.18).

TABLE 20.18 Energy content of commonly available sleigh fuels

Fuel	Energy content (kJ L^{-1})
Rudolphene	45
Polar plus	29
Super sleigh	53

(Claus, 2016, p. 45)

Footnotes system

This style of in-text referencing is usually used in the Chicago and Vancouver style of referencing.

In this style, the full citation is included at the bottom of the page the referenced material appears on, with a superscript number showing the point in the text where the material has been used. For example:

> *. . . over the past 10 years, the number of eligible children has increased[5] and a need has, therefore, developed for sleighs to travel faster to meet the required delivery schedule. More efficient fuels are required for this purpose.*

In the footnotes section of the relevant page (depending on the format of the report), the work is then referenced alongside the corresponding number, and then again included in the reference list.

[5] Kringle, K. (2008). *Journal of Polar Transport*, Vol. 34, p. 356.

Acknowledging sources at the end of your report: Reference list

At the end of a scientific report or poster, a reference list is included.

In your reference list, references should be listed alphabetically. If the footnotes version of in-text referencing was used, references should be listed in order of footnote number.

The following provides examples of the use of the Harvard style in creating references.

Book

Author surname(s), initial(s) (Year published). *Title*. Edition (if applicable). Place of publication: publisher.

Example:

Langley, P & Jones, M (2013). *The search for Richard III: The king's grave.* London: John Murray.

Note: do not use 'et al.' in your reference list. This is only appropriate in your in-text referencing.

Journal

Author surname(s), initial(s) (Year published). *Title of article*. Title of Journal. Volume number. Page numbers.

Example:

Boiero, D & Bagaini, C (2018). *S-wave splitting intensity analysis and inversion.* Geographical Prospecting. Vol. 67(2). 362–378.

If available, you may also include the DOI (digital object identifier) after the page numbers. A DOI is a permanent identifier for a journal article and is often used in place of a URL.

TV programs

Title of program (date), (*TV program*) Channel identification.

Example:

Gene editing made simple (2016), (*Catalyst*), ABC Sydney.

Websites

When using websites, you should be using those written for an academic target audience in your report.

Author surname(s), initial(s) (Year published)c. *Title of page*. Name of website, date and website of retrieval

Example:

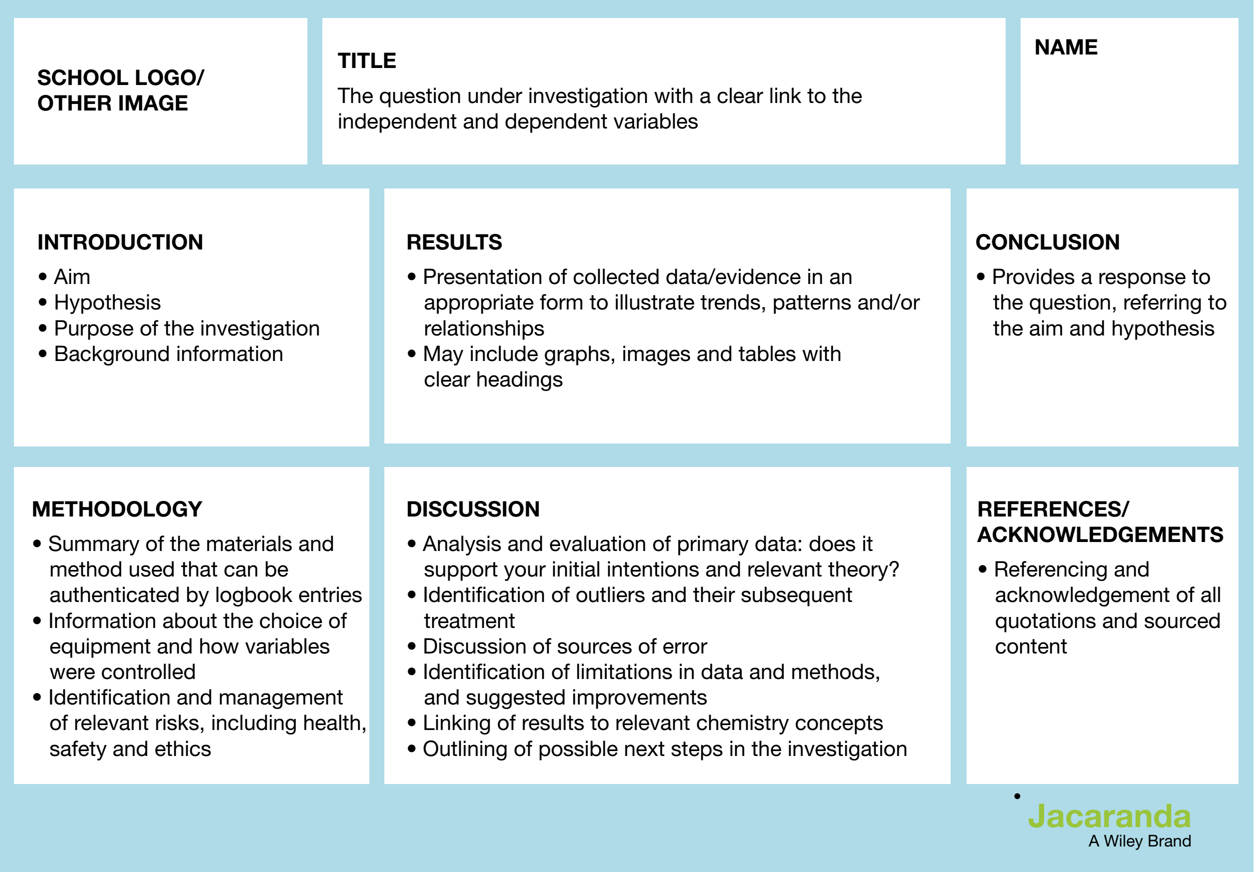
Fein, B (2011). *The Patriot Act is at war with the constitution*. The American view: God, family, republic, accessed 8 November 2011, <http://archive.theamericanview.com/index.php?id=1770>

20.10.9 Presenting a scientific poster

Your logbook will form a key part of your assessment. You will also be required to show a scientific report. This may be in the form of a poster, oral communication or digital communication as selected by your teacher.

A sample format of the poster is shown in figure 20.56.

FIGURE 20.56 Example format of a scientific poster



SCHOOL LOGO/ OTHER IMAGE

TITLE

The question under investigation with a clear link to the independent and dependent variables

NAME

INTRODUCTION

- Aim
- Hypothesis
- Purpose of the investigation
- Background information

RESULTS

- Presentation of collected data/evidence in an appropriate form to illustrate trends, patterns and/or relationships
- May include graphs, images and tables with clear headings

CONCLUSION

- Provides a response to the question, referring to the aim and hypothesis

METHODOLOGY

- Summary of the materials and method used that can be authenticated by logbook entries
- Information about the choice of equipment and how variables were controlled
- Identification and management of relevant risks, including health, safety and ethics

DISCUSSION

- Analysis and evaluation of primary data: does it support your initial intentions and relevant theory?
- Identification of outliers and their subsequent treatment
- Discussion of sources of error
- Identification of limitations in data and methods, and suggested improvements
- Linking of results to relevant chemistry concepts
- Outlining of possible next steps in the investigation

REFERENCES/ ACKNOWLEDGEMENTS

- Referencing and acknowledgement of all quotations and sourced content

20.10.10 Practical investigation checklist

☐ Your name, the title and the aim/hypothesis are listed.
☐ An introduction describes the purpose and outlines the investigation in a logical and concise manner. Key terms are defined and variables are stated clearly. Relevant theory is addressed.
☐ The method is outlined clearly in step form, including a consideration of ethics, health and safety. A risk assessment is provided.
☐ Your logbook contains dates, headings and complete records.
☐ Any abbreviations are explained.
☐ Results are presented in an organised way, in a table if possible. All relevant measurements are recorded with appropriate accuracy and units.
☐ Observations are clear and concise, as are all diagrams, graphs and tables used.
☐ Any calculations are shown.
☐ A concise summary and interpretation of key findings is included, outlining trends and any unexpected results with connection to theory.
☐ The experimental design is evaluated and possible improvements are included.
☐ There are suggestions for future investigations.
☐ The conclusion concisely summarises how your results support or contradict your original hypothesis.
☐ All sources are acknowledged and references correctly cited.
☐ The use of key terms, symbols and equations is appropriate.

A poster or other form of scientific report should address the sections outlined in table 20.13 without going into too much detail. For example, you would display only a subset of the data to convey your findings and accuracy. Similarly, not all your graphs need appear.

20.10 EXERCISE

To answer questions online and to receive **immediate feedback** and **sample responses** for every question, go to your learnON title at www.jacplus.com.au.

1. Describe the following aspects of a scientific report.
   (a) Introduction
   (b) Discussion
   (c) Conclusion
2. Convert the following units to the SI base unit shown.
   (a) 1984 mL (to L)
   (b) 1.231 km (to m)
   (c) 53153 μs (to s)
3. Convert the following to the unit shown.
   (a) 0.123 kg (to mg)
   (b) 167283 μmol (to mmol)
4. How many significant figures are in the following?
   (a) 1549.12
   (b) 0.0404
   (c) 0.01
   (d) 142.4000
5. Calculate the following and express in the appropriate number of significant figures.
   (a) 4.25 + 9
   (b) 0.04 + 3.7
   (c) 5.640 + 70.435
   (d) 0.80 − 0.3
   (e) 840 − 627.03
   (f) 12.01 + 6.7
6. Solve the following problems using the correct number of significant figures.
   (a) Density = 27.8 grams/1.2 mL =
   (b) Mass = 23.45 grams + 5.332 grams =
   (c) Moles = 1.221M × 2.6 L =
   (d) $m(\text{Cu}) = 0.45 \text{ mol} \times 63.5 \text{ g mol}^{-1} =$
   (e) The total time of a reaction to occur from the following three split times: 14.23 seconds, 10.1 seconds, 11.29 seconds.
7. Write a reference in Harvard style for this textbook. Show both the in-text referencing and the full reference for the reference list.

Fully worked solutions and sample responses are available in your digital formats.

20.11 Review

20.11.1 Summary

Key science skills in chemistry

- Choose a topic that allows for the development of a question and observations to be made.
- Variables are factors that an investigator can control, change or measure.
- An independent variable is manipulated by the investigator; for example, the type of metal examined.
- A dependent variable is measured by the investigator and is influenced by the independent variable; for example, the melting temperature of a metal.

- A controlled variable is one that is kept the same in an investigation, and an experiment usually has numerous controlled variables; for example, the device used to record temperature, environmental factors such as air temperature and humidity, and the mass of the metal being investigated.
- It is important to control variables in order to examine one independent variable at a time to better understand the relationship to the dependent variable.
- An aim is a one to two sentence outline of the purpose the investigation, linking the dependent and independent variables.
- A hypothesis is a tentative, testable and falsifiable statement for an observed phenomenon; it acts as a prediction for the investigation.
- Part of the scientific method or process involves designing an experiment, conducting the experiment and analysing the results.

Concepts specific to investigation

- Key terms can be defined within a scientific report or in a glossary.
- Key terms that are selected should be those that are vital to understanding the investigation.
- Various representations are used in chemistry, including symbols, models, equations and formulas.
- Understanding various representations helps make communication easier and clearer.

Scientific research methodologies and techniques

- Primary sources of data are direct or firsthand evidence, often gained through practical investigations.
- Data can be quantitative (numerical) or qualitative.
- It is vital to carefully choose the best techniques and equipment for an investigation. These should be precise, accurate and easy to conduct within a school environment.
- In an investigation, it is vital to consider the accuracy, precision, reliability and validity.
- Accuracy is how close a measurement is to a known value (like hitting the target on a bullseye).
- Precision is how close multiple measures are to each other (like hitting the same spot on a target three times in a row).
- Reliability shows that if an experiment is replicated (by the original researcher or another researcher), the results should be similar.
- Validity is how much credibility can be given to the results; for example, do the results measure what they intended to and back up the claims made?
- The best-designed experiments are accurate, precise, reliable and valid, while minimising errors and uncertainties.

Ethics, health and safety guidelines

- In an investigation, care must be taken to follow ethical, health and safety guidelines.
- Ethics are involved in moral conduct, particularly when humans and animals are involved as test subjects; this does not tend to be as big an issue in chemistry as in biology and psychology.
- Health and safety is often evaluated using a risk assessment, where different hazards are analysed and safety requirements are made.

Methods of organising, analysing and evaluating primary data

- Data can be organised in a variety of ways, including through the use of tables and graphs.
- Scatterplots and line graphs are used when both variables are quantitative.
- Bar graphs are used when one piece of data is qualitative and the other is quantitative.
- Histograms are used when intervals and frequency are being explored.
- It is important to choose an appropriate type of organising data in order to analyse trends, patterns and relationships.
- Graphs can be used to predict data outside of the investigated set.
- When evaluating data, it is important to link this to scientific concepts, models and theories, and link data back to your aim and hypothesis, describing trends and patterns, and identifying any outliers or limitations.

- Both errors and uncertainty can affect the results gained. Errors are a measurement of the difference between a measured value and the true value (the accuracy), whereas uncertainty is limits to precision and the range of values that can be measured.

Models and theories to understand observed phenomena

- Models represent ideas, phenomena or scientific processes. They can be physical models, mathematical models or conceptual models.
- Limitations of models include an oversimplification of concepts, an inability to be completely accurate, and being specific to a set of conditions, which may or may not reflect the real world.
- Theories are well-supported explanations of phenomena.
- Similar to models, theories are often oversimplifications of concepts, specific to a set of conditions and subject to change as new information becomes available.

Nature of evidence and key findings of investigations

- Evidence can be used to support or reject a hypothesis.
- A hypothesis is not ever stated to be confirmed or proven true because new evidence may come to light that later leads to its rejection.
- When reporting key findings, it is important to provide information about the data obtained in the practical investigation and highlight any patterns and trends, show the relationship of your findings to chemistry concepts (including solubility, concentration, acid–base and redox concepts) and provide an answer to the question of your investigation.

Options and issues related to water quality

- It is important to be able to monitor water quality because it is vital for life.
- Contaminants found in water may include heavy metals, organic pollutants and microorganisms.
- Water also needs to be monitored for factors such as salts, organic molecules and acids and bases.
- Selecting an appropriate technique to analyse water is important in order to gain accurate results in your practical investigation.

Conventions of scientific report writing and scientific poster presentation

- Both your logbook and scientific poster are forms of scientific reporting and combine for your mark for Unit 2 AOS 3.
- Components required in a scientific report include a title, date, name, introduction, aim, hypothesis, methodology (method and materials), results, discussion, conclusions and references.
- A variety of key symbols, formulas, equations, terminology, representations and abbreviations are used in chemistry and are required for effective communication.
- All sourced information must be correctly and clearly referenced using a consistent style, such as Harvard or APA.
- Your poster will not have every piece of information that your logbook will, but will show the important and key points to allow for a clear understanding of your investigation.

on Resources

Digital document Topic Summary — Topic 20 (doc-30960)

20.11.2 Key terms

accuracy how close an experimental measurement is to a known value
aim statement outlining the purpose of an investigation, linking the dependent and independent variables
bar graph graph in which data is represented by a series of bars; usually used when one variable is quantitative and the other is qualitative
bias the intentional or unintentional influence on a research investigation

categorical data data that has labels or names rather than a range of numerical quantities; also known as qualitative data

conclusion section at the end of a scientific report that relates back to the question, sums up key findings and states whether the hypothesis was supported or rejected

control group group that is not affected by the independent variables, and is used as a baseline for comparison

controlled variable variable that is kept constant across different experimental groups

dependent variable variable that is influenced by the independent variable; the variable that is measured

discussion detailed area of a scientific report in which results are discussed, analysed and evaluated, relationships to concepts are made, errors, limitations and uncertainties are assessed and suggestions for future improvements are outlined

error difference between a measurement taken and the true value that is expected; errors lead to a reduction in the accuracy of the investigation

ethics acceptable and moral conduct determining what is ‘right’ and ‘wrong’

experimental group test group that is exposed to the independent variable

histogram graph in which data is sorted in intervals and frequency is examined; used when both pieces are data are quantitative

hypothesis tentative, testable and falsifiable statement for an observed phenomenon; acts as a prediction for the investigation

independent variable variable that is changed or manipulated by an investigator

limitations factors that have affected the interpretation and/or collection of findings in a practical investigation

line graph graph in which points of data are joined by a connecting line; used when both pieces of data are quantitative (numerical)

line of best fit trend line added to a scatterplot to best express the data shown; these are straight lines, and are not required to pass through all points

model representation of ideas, phenomena or scientific processes; can be physical models, mathematical models or conceptual models

numerical data data that involves numbers and can be measured or counted; also known as quantitative data

outlier result that is a long way from other results and seen as unusual

precision how close multiple measurements of the same investigation are to each other

primary source direct or firsthand evidence about some phenomenon

qualitative data categorical data that examines the quality of something (e.g. colour or gender) rather than numerical values

quantitative data numerical data that examines the quantity of something (e.g. length or time)

random error chance variation in measurements

reliability whether another researcher could repeat an investigation by following the outlined method and obtain similar results

results section in a scientific report in which all data obtained is recorded, usually in the form of tables and graphs

risk assessment document that examines the different hazards in an investigation and suggests safety precautions

scatterplot graph in which two quantitative variables are plotted as a series of dots

scientific method procedure that must be followed in scientific investigations that consists of questioning, researching, predicting, observing, experimenting and analysing; sometimes referred to as scientific process

scientific research methodology the principles of research based on the scientific method

secondary source comments on or summaries and interpretations of primary data

systematic errors errors that affect the accuracy of a measurement that cannot be improved by repeating an experiment; usually due to equipment or system errors

theory well-supported explanation of a phenomena, based on facts that have been obtained through investigations, research and observations

uncertainty limit to the precision of equipment; a range within which a measurement lies

validity shows how much results measure what the investigation was intending to, and how well they show the claims the investigation makes

on Resources

Digital document Key terms glossary — Topic 20 (doc-30959)

20.11 Exercises

To answer questions online and to receive **immediate feedback** and **sample responses** for every question go to your learnON title at www.jacplus.com.au

20.11 Exercise 1: Multiple choice questions

1. Belinda conducted an experiment to test the combustion of different types of fuel. In this experiment, what is the volume of fuel being tested?
   - **A.** The dependent variable
   - **B.** The independent variable
   - **C.** The controlled variable
   - **D.** The control group
2. Students were conducting an experiment to investigate the temperature of different rooms in the school. Each measurement was taken three times using thermometer and recorded by the student. A highly accurate digital thermometer was also used.

   The following table shows the results recorded by each student.

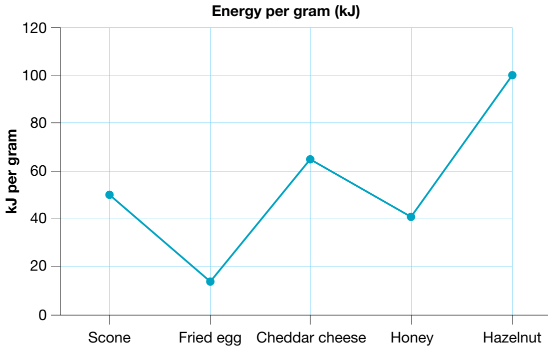
	Recorded temperature using thermometer (°C)			
Student	**Measurement 1**	**Measurement 2**	**Measurement 3**	**Digital thermometer reading (°C)**
Mel	26	25	25.25	25.5
Josette	24	32.5	20	24.8
Fiona	28.5	28.25	28.5	24.7
Judy	29.9	31.2	32	31.2

   Which of the following statements is most correct about the results recorded?
   - **A.** Mel’s data is accurate but not precise.
   - **B.** Josette’s data is precise but not accurate.
   - **C.** Fiona’s data is the most accurate but not the most precise.
   - **D.** Judy’s data is more accurate than Fiona’s.
3. Paul was investigating the energy content from different types of food through combustion. The results of his investigation are shown in the following table.

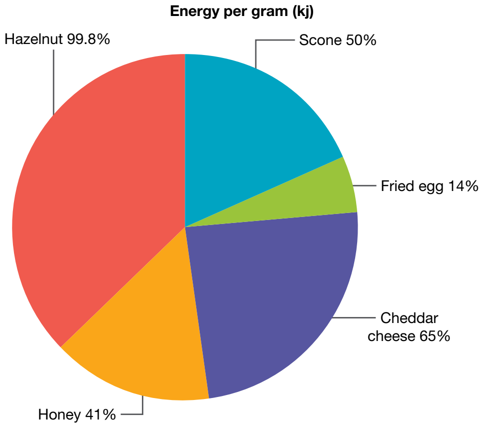
Type of food	Energy per gram (kJ)
Scone	50
Fried egg	14
Cheddar cheese	65
Honey	41
Hazelnut	99.8

What is the most appropriate graph to use to show and analyse trends in this data?

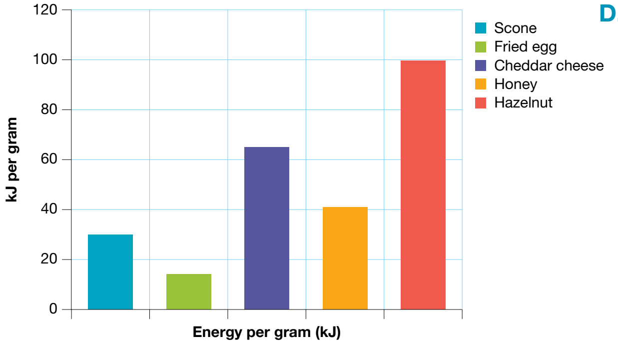
A.



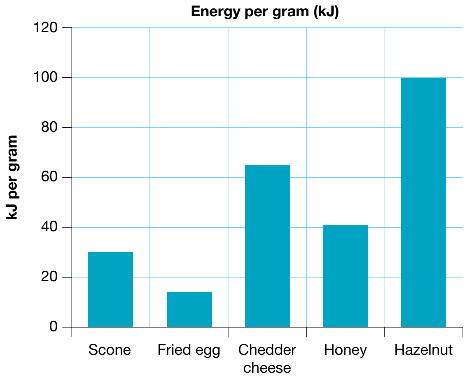
B.



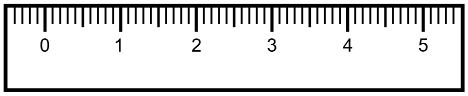
C.



D.



4. What are qualitative results?
   - **A.** Results that are expected to be observed during an experiment
   - **B.** Results that are measured in an experiment
   - **C.** Results that are observed during an experiment
   - **D.** None of the above
5. What would the measure of uncertainty be in the ruler shown?
   - **A.** 0.5 cm
   - **B.** 0.1 cm
   - **C.** 0.05 cm
   - **D.** 0.025 cm

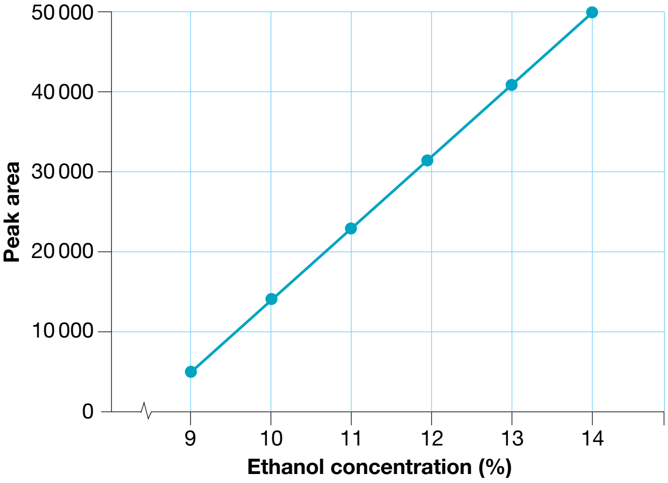


6. Systematic errors can
   - **A.** affect the accuracy of a reading.
   - **B.** be caused by human fault.
   - **C.** be improved by repetition an experiment.
   - **D.** be caused by deliberate changing of results due to bias.
7. What is the difference between an aim and a hypothesis?
   - **A.** The aim of an experiment is a statement, but the hypothesis is a question.
   - **B.** The aim includes the dependent and independent variables, but a hypothesis doesn't.
   - **C.** The aim explains the expected data, but the hypothesis explains how you will get the data.
   - **D.** The aim outlines the purpose of the investigation, but the hypothesis is a testable prediction.
8. Stacey and Bridget both swam 200 metres. On a very close finish, it was found that Stacey was 999 milliseconds faster than Bridget.
   Which of the following is correct about her result?
   - **A.** It is equal to 9.99 seconds.
   - **B.** It contains two significant figures.
   - **C.** It can be written as 9.99×10^2 milliseconds.
   - **D.** It can be shortened to 999 Ms.

9. What are theories?
   - A. Concepts that were once accurate, but have now been rejected
   - B. Diagrammatic representations of abstract concepts
   - C. Ideas that are unable to be rejected due to the amount of evidence they have
   - D. Well-supported ideas where evidence has been gained from investigations, research and observations
10. On a risk assessment, a chemical has an image of a flame and an exclamation mark. What does this mean the chemical is?
   - A. An oxidiser and toxic
   - B. Flammable and harmful
   - C. Flammable and toxic
   - D. Harmful and explosive

20.11 Exercise 2: Short answer questions

1. An investigation was conducted to observe the differences in pH of five different swimming pools.
   - a. Identify two pieces of quantitative data that you could record and measure.
   - b. What instruments, if any, would you need to make these quantitative observations?
   - c. Identify two pieces of qualitative data that you could observe.
   - d. What instruments, if any, would you need to make these qualitative observations?
2. Identify a key difference between the members of the following pairs.
   - a. Independent and dependent variables
   - b. Qualitative and quantitative data
   - c. Control group and experimental group
   - d. Primary and secondary sources of data
   - e. Uncertainty and error
3. In an investigation conducted in class, Jazz recorded a concentration of 0.0540 mol L^{-1}.
   - a. How many significant figures does this recorded concentration have?
   - b. Write this concentration in scientific notation.
4. Chris used gas chromatography (GC) to measure the ethanol content of some alcoholic beverages. GC works on a similar principle to high-performance liquid chromatography (HPLC). In both instruments, it is necessary to produce a calibration curve in addition to obtaining readings for the test samples. One such calibration curve is as shown.



   - a. Sample A produced a reading of 36 000 from the GC. Estimate the level of ethanol in this sample. What is the name of the process used to obtain this answer?
   - b. Sample B produced a reading of 50 500. Estimate the level of ethanol in this sample and comment on your answer. What is the name of the process used to obtain this estimate?
   - c. Sample C produced a reading of 95 000. Is it possible to estimate the ethanol level in this sample? Explain why or why not.
   - d. Write a conclusion for this investigation.

5. Ammie conducted an experiment to explore how the concentration of cordial affects the freezing rate. The results she recorded are shown in the following table.

Concentration of cordial (%)	Time to freeze (mins)
0	25
10	32
20	45
30	53
40	60
50	65

a. Using graph paper, plot the data using the most appropriate graph type, including a scale and labels.
b. Describe the trends and patterns in this graph
c. What conclusions would you make from this investigation?

20.11 Exercise 3: Exam practice questions

Question 1 (10 marks)

Vicki decided to investigate how solubility is affected by the temperature of the solution. She had a container of sodium sulfate powder. She put on some safety gloves and glasses and measured the temperature of the water in a test tube using a thermometer. She added sodium sulfate until it could no longer dissolve, and recorded how much sodium sulfate she added to reach this point. She then got a new test tube with water and heated it to 50 °C and repeated the process. She found that more sodium sulfate dissolved at 50 °C compared to that at room temperature

a. Comment on the safety precautions taken. Do you think they were sufficient? **2 marks**
b. Identify the dependent variable in this investigation. **1 mark**
c. List two variables that need to be controlled in this investigation. **1 mark**
d. Describe what measuring equipment should be used in this investigation and identify **one** factor that could affect the accuracy of this piece of equipment. **2 marks**
e. Write a clear experimental method for Vicki that shows how she can investigate how solubility is affected by temperature. You may need to add to her method to make sure that more data is being collected and to ensure it is reproducible. **4 marks**

Question 2 (10 marks)

An investigation was being conducted to examine if different brands of bottled water have different pH.

a. Write a suitable aim for this investigation. **1 mark**
b. Identify the hypothesis for this investigation. **1 mark**
c. Describe **one** piece of qualitative data and **one** piece of quantitative data that may be collected in this investigation. **2 marks**
d. Explain **two** factors that may lead to differences in results between different students. **2 marks**
e. Outline an appropriate method for this investigation. **4 marks**

Question 3 (8 marks)

A student conducted an experiment examining the relationship between temperature and the pH of water. The results from this investigation are shown in the table.

Temperature (°C)	pH
0	7.7
10	7.5
20	7.3
30	7.0
40	6.7
50	6.4

a. On graph paper:
  - Plot the graph showing the data provided, ensuring that pH is shown on the vertical axis.
  - Label the graph and axes as appropriate.
  - Draw a line showing the relationship between the points. **4 marks**

b. Describe the trends seen in your data. **2 marks**

c. Based on the graph, explain if you could accurately determine the pH of the solution at 90 °C. **2 marks**

Question 4 (7 marks)

Two students, Alan and Will, conducted an investigation to confirm that the boiling point of water is actually 100 °C, regardless of the volume of water being tested. They examined five different volumes and recorded the temperature of boiling.

Their results are shown in the table.

	Alan	Will
Volume (mL)	Recorded boiling point (°C)	Recorded boiling point (°C)
20	100.5	97
40	100	97.5
60	98.5	95.5
80	101.5	96
100	102.5	96.5

a. Describe **one** error that may have occurred that reduced the accuracy of the data for both Alan and Will, and identify how this error may be avoided. **2 marks**

b. Calculate the average temperature recorded by Will, using correct significant figures based on the data provided. **2 marks**

c. Describe, with reference to the known value of the boiling point of water, why the results obtained by Alan are more accurate than those obtained by Will. **2 marks**

d. It was found that the density of water at the start of the investigation was 0.987 g mL^{-1}. Convert this into grams per litre. **1 mark**

Question 5 (12 marks)

A student conducted an investigation to determine the rate in which the temperature of water decreases when in different types of insulating material.

The student's report is as follows.

> *Introduction: In this experiment, three materials were examined: aluminium foil, paper and wool. This experiment was conducted over three days, with aluminium foil explored on the first day, paper on the second day and wool on the third day. All materials were wrapped around a polystyrene cup in which boiling water was placed.*

Aim: To determine how different materials affect the temperature decrease of water.

Hypothesis: If aluminium foil, paper and wool are used as insulation material, then aluminium foil will work best.

Method:

1. *Collect a polystyrene cup and wrap a piece of aluminium foil around it.*
2. *Place 100 mL of boiling water in the cup and record the temperature.*
3. *Record the temperature every five minutes.*
4. *Repeat Steps 1 to 3 with paper and then with wool.*

Results:

	Aluminium	Paper	Wool
Initial	100.0 °C	98.5 °C	95.0 °C
5 minutes	95.5 °C	85.0 °C	80.0 °C
10 minutes	92.0 °C	78 °C	71.5 °C
15 minutes	85.0 °C	62.0 °C	60.0 °C
20 minutes	80.5 °C	50.0 °C	48.5 °C
25 minutes	75.0 °C	40.0 °C	39.5 °C

a. Describe an issue with the hypothesis written by the student. How would you adjust this to make it testable? **2 marks**

b. The experiment was conducted over three days. Explain why this may lead to errors in the data obtained. **2 marks**

c. Students were only able to use a thermometer in which temperature could be measured to the nearest 0.5 °C. Identify the tolerance of this device and describe the uncertainty expected in the data. **2 marks**

d. Outline two limitations in the experimental method or data collection process that would affect the conclusions drawn. **2 marks**

e. Describe the most appropriate graph that the student should use to represent their data. Justify your choice. **2 marks**

f. Based on the student's results, write a conclusion for this investigation, linking back to the hypothesis. **2 marks**

Answers

1 Elements

1.2 Atomic structure

Practice problem 1

1 000 000 times larger.

1.2 Exercise

1. Helium atom, DNA molecule, bacteria, width of a human hair, flea.
2. 20 times smaller
3. 140 times smaller

Practice problem 2

1. H
2. Each atom has one proton, the atomic number $Z = 1$
3. hydrogen-1: 0
   hydrogen-2: 1
   hydrogen-3: 2
4. $^{1}_{1}\text{H}$, $^{2}_{1}\text{H}$ (deuterium), $^{3}_{1}\text{H}$ (tritium)

1.3 Representing atoms

1.3 Exercise

1. **a.** $A = 1$ **b.** $A = 10$ **c.** $A = 47$ **d.** $A = 79$
2. **a.** $A = 13$ **b.** $Z = 26$ **c.** Aluminium
3. **a.** B **b.** Mg **c.** Ar **d.** Ca
4. **a.** 22 **b.** 18 **c.** $^{40}_{18}\text{Ar}$
5. number of protons = 35,
   number of neutrons = 44
6. $^{2}_{1}\text{H}$
7.

Element	Number of protons	Number of electrons	Number of neutrons
$^{12}_{6}\text{C}$	6	6	6
$^{56}_{26}\text{Fe}$	26	26	30
$^{40}_{18}\text{Ar}$	18	18	22
$^{235}_{92}\text{U}$	92	92	143
$^{238}_{92}\text{U}$	92	92	146
$^{19}_{9}\text{F}$	9	9	10

   a. $^{235}_{92}\text{U}$ and $^{238}_{92}\text{U}$
   b. $^{238}_{92}\text{U}$ has 3 more neutrons than $^{235}_{92}\text{U}$.
8. The atomic number is always the same for a particular element, while the mass number can differ due to the existence of isotopes (an element with differing numbers of neutrons).

1.4 Electrons

Practice problem 3

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

Practice problem 4

Yes

1.4 Exercise

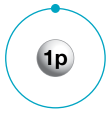
1.

Energy level (n)	Maximum number of electrons ($2n^2$)
1	$2 \times 1^2 = 2$
2	$2 \times 2^2 = 8$
3	$2 \times 3^2 = 18$
4	$2 \times 4^2 = 32$

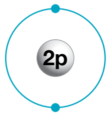
2. **a.** The Bohr model could explain the emission spectrum of hydrogen.
   b. The spectra of atoms more complex than hydrogen could not be explained satisfactorily.
3.

Element	Symbol	Atomic number	Electron configuration 1, 2, 3, 4
Hydrogen	H	1	1
Helium	He	2	2
Lithium	Li	3	2, 1
Beryllium	Be	4	2, 2
Boron	B	5	2, 3
Carbon	C	6	2, 4
Nitrogen	N	7	2, 5
Oxygen	O	8	2, 6
Fluorine	F	9	2, 7
Neon	Ne	10	2, 8
Sodium	Na	11	2, 8, 1
Magnesium	Mg	12	2, 8, 2
Aluminium	Al	13	2, 8, 3
Silicon	Si	14	2, 8, 4
Phosphorus	P	15	2, 8, 5
Sulfur	S	16	2, 8, 6
Chlorine	Cl	17	2, 8, 7
Argon	Ar	18	2, 8, 8
Potassium	K	19	2, 8, 8, 1
Calcium	Ca	20	2, 8, 8, 2

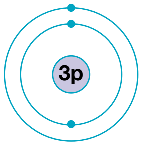
4. Hydrogen, H



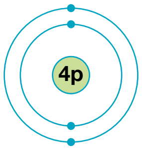
Helium, He



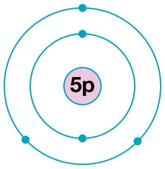
Lithium, Li



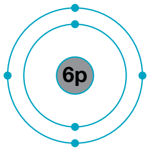
Beryllium, Be



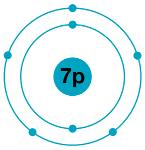
Boron, B



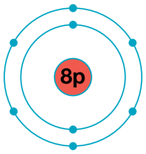
Carbon, C



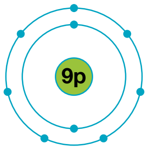
Nitrogen, N



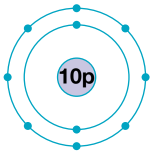
Oxygen, O



Fluorine, F



Neon, Ne



5. **a.** 2, 5. **b.** Ca.

6. ${}_{3}\text{Li}$: $1s^2 2s^1$
${}_{19}\text{K}$: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
${}_{6}\text{C}$: $1s^2 2s^2 2p^2$
${}_{17}\text{Cl}$: $1s^2 2s^2 2p^6 3s^2 3p^5$
${}_{18}\text{Ar}$: $1s^2 2s^2 2p^6 3s^2 3p^6$
${}_{7}\text{N}$: $1s^2 2s^2 2p^3$

7. **a.** ${}^{32}_{15}\text{P}$: $1s^2 2s^2 2p^6 3s^2 3p^3$
b. ${}^{32}\text{P}$ has 17 neutrons while ${}^{31}\text{P}$ has only 16.

8. **a.** Group 16 **b.** Period 3 **c.** Sulfur, S

9. **a.** Lithium **b.** Neon **c.** Sodium **d.** Aluminium **e.** Argon **f.** Calcium

10. **a.** $Z = 20$: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

$Z = 12$: $1s^2 2s^2 2p^6 3s^2$

$Z = 4$: $1s^2 2s^2$

$Z = 9$: $1s^2 2s^2 2p^5$

b. $Z = 5$: $1s^2 2s^2 2p^1$

$Z = 6$: $1s^2 2s^2 2p^2$

$Z = 8$: $1s^2 2s^2 2p^4$

$Z = 16$: $1s^2 2s^2 2p^6 3s^2 3p^4$

11. A shell is a region in which a group of electrons of similar energies moves around the nucleus. Subshells are found within any shell and contain orbitals with the same energy level.

12.

Element	Symbol	Atomic number	Ground state electron configuration
Hydrogen	H	1	$1s^1$
Helium	He	2	$1s^2$
Lithium	Li	3	$1s^2 2s^1$
Beryllium	Be	4	$1s^2 2s^2$
Boron	B	5	$1s^2 2s^2 2p^1$
Carbon	C	6	$1s^2 2s^2 2p^2$
Nitrogen	N	7	$1s^2 2s^2 2p^3$
Oxygen	O	8	$1s^2 2s^2 2p^4$
Fluorine	F	9	$1s^2 2s^2 2p^5$
Neon	Ne	10	$1s^2 2s^2 2p^6$
Sodium	Na	11	$1s^2 2s^2 2p^6 3s^1$
Magnesium	Mg	12	$1s^2 2s^2 2p^6 3s^2$
Aluminium	Al	13	$1s^2 2s^2 2p^6 3s^2 3p^1$
Silicon	Si	14	$1s^2 2s^2 2p^6 3s^2 3p^2$
Phosphorus	P	15	$1s^2 2s^2 2p^6 3s^2 3p^3$
Sulfur	S	16	$1s^2 2s^2 2p^6 3s^2 3p^4$
Chlorine	Cl	17	$1s^2 2s^2 2p^6 3s^2 3p^5$
Argon	Ar	18	$1s^2 2s^2 2p^6 3s^2 3p^6$
Potassium	K	19	$1s^2 2s^2 2p^6$ $3s^2 3p^6 4s^1$
Calcium	Ca	20	$1s^2 2s^2 2p^6$ $3s^2 3p^6 4s^2$

13. a. Ground state
    b. Excited state
    c. Ground state
14. a. The electrons are not in the lowest available energy levels.
    b. When an electron changes to the ground state energy would be emitted and a photon released.
    c. $1s^22s^22p^6$
15. a. Mn
    b. Ca

1.5 Review

1.5 Exercises

1.5 Exercise 1: Multiple choice questions

1. C.
2. D.
3. B.
4. A.
5. C.
6. D.
7. D.
8. C.
9. B.
10. D.
11. C.
12. A.
13. B.
14. C.
15. B.
16. C.

1.5 Exercise 2: Short answer questions

1.

Object	Size
Red blood cell	7000 nm
Virus	30 nm
Hydrogen molecule	0.15 nm

2. a. i. $Z = 11$
    ii. $A = 23$
    iii. $Z = 11$
    iv. Number of neutrons = 12
    v. Number of electrons = 11
    vi. Sodium

b. i. $Z = 9$
    ii. $A = 19$
    iii. $Z = 9$
    iv. Number of neutrons = 10
    v. Number of electrons = 9
    vi. Fluorine

c. i. $Z = 14$
    ii. $A = 28$
    iii. $Z = 14$
    iv. Number of neutrons = 14
    v. Number of electrons = 14
    vi. Silicon

d. i. $Z = 26$
    ii. $A = 56$
    iii. $Z = 26$
    iv. Number of neutrons = 30
    v. Number of electrons = 26
    vi. Iron

e. i. $Z = 79$
    ii. $A = 197$
    iii. $Z = 79$
    iv. Number of neutrons = 118
    v. Number of electrons = 79
    vi. Gold

f. i. $Z = 92$
    ii. $A = 235$
    iii. $Z = 92$
    iv. Number of neutrons = 143
    v. Number of electrons = 92
    vi. Uranium

3. See table at foot of the page.*
4. Isotopes
5. a. A, C and D.
    b. $A={}^{40}_{20}\text{Ca}$, $B={}^{37}_{17}\text{Cl}$, $C={}^{42}_{20}\text{Ca}$, $D={}^{46}_{20}\text{Ca}$.
6. $Z = 17$: chlorine ${}^{37}_{17}A$, ${}^{35}_{17}D$

    $Z = 12$: magnesium ${}^{26}_{12}B$, ${}^{25}_{12}E$, ${}^{24}_{12}G$

    $Z = 27$: cobalt ${}^{59}_{27}C$, ${}^{60}_{27}F$
7. a. It needs to absorb energy.
    b. A photon equal to the difference in energy of the two levels is emitted.
8. One colour
9. The ground state refers to the most stable state of an atom, where all electrons are occupying their normal energy levels or shells. When an atom is excited, the electrons can absorb energy and jump to higher energy levels (the excited states), depending on the amount of energy absorbed.

3*

Name of atom	Atomic number	Mass number	Protons	Neutrons	Electrons
Argon	18	36	18	36 − 18 = 18	18
Sulfur	16	34	16	34 − 16 = 18	16
Argon	18	38	18	38 – 18 = 20	18
Phosphorus	15	31	15	31 – 15 = 16	15
Lead	82	208	82	208 – 82 = 126	82
Potassium	19	39	19	39 – 19 = 20	19
Sulfur	16	36	16	36 – 16 = 20	16

10. a. A shell is the main energy level where an electron may be found. A shell is made up of subshells. Each shell has a specific number of subshells.
    b. An orbit is a regularly repeated path or circuit. An orbital is a region of space in which electrons move.
11. a. Half-filled and filled subshells are more stable than other partly filled subshells
    b. $1s^22s^22p^63s^23p^63d^54s^1$
    c. $1s^22s^22p^63s^23p^63d^{10}4s^1$

1.5 Exercise 3: Exam practice questions

1. Neutral atoms must have the same number of protons (positively charged) and electrons (negatively charged). If a species had 6 protons and 13 electrons, it would have an overall charge of +7, which would make it a charged species (ion).
2. a. 40 000 times smaller
    b. i. 60 000 nm^2 **ii.** 1 000 000 nm^3 **iii.** 0.06
    c. The surface area to volume ratio gets 10 times bigger.
3. a. i. $\text{Ar} = 1s^22s^22p^63s^23p^6$
        ii. $\text{Ca}^{(2+)} = 1s^22s^22p^63s^23p^6$
    b. The argon atom has 8 electrons in its outer third shell ($3s^23p^6$; $2 + 6 = 8$) so it is in group 18. The number of outer electrons in the atom determines its group number. It has 3 shells and so would be found in Period 3 because this corresponds to the number of shells present in the atom.

2 The periodic table

2.2 Organisation of the periodic table

2.2 Exercise

1. The word ‘ununseptium’ is another way of saying the number 117, from Latin for one-one-seven. Detailed responses can be found in the worked solutions in your digital formats.
2. Chemists study the physical and chemical properties of the elements (and their compounds) that the periodic table is based on. Detailed responses can be found in the worked solutions in your digital formats

2.3 Patterns and trends in the periodic table

Practice problem 1

The element is nitrogen as the atom belongs to period 3 and group 15. Therefore, the ion is N^{3-} (nitride).

2.3 Exercise

1. a. 8 electrons = 8 protons = period 2, group 16, oxygen
    b. 3 electrons = 3 protons = period 2, group 1, lithium
    c. 26 electrons = 26 protons = period 4, group 8, iron
    d. 36 electrons = 36 protons = period 4, group 18, krypton
    e. 21 electrons = 21 protons = period 4, group 3, scandium
    f. 30 electrons = 30 protons = period 4, group 12, zinc
2. a. The element is manganese and belongs to period 4, group 5.
    The original ion was Mn^{7+}.
    b. The element is chlorine, and belongs to period 3, group 17.
    The original ion was Cl^- (chloride).
3. a. $1s^22s^22p^6$, period 2, group 16
    b. $1s^22s^22p^63s^23p^1$, period 3, group 13
    c. $1s^22s^22p^63s^23p^64s^2$, period 4, group 2
    d. $1s^22s^22p^63s^23p^64s^23d^{10}4p^5$, period 4, group 17
4. a. $1s^22s^22p^63s^23p^64s^23d^{10}$.
    b. A transition metal is defined as any element with an incomplete *d* subshell that can form several stable ions. Zn fulfils neither of these requirements because it has a filled 3*d* subshell and forms only a 2+ ion.
5. a. The electronegativity increases from left to right across the periodic table.
    b. The atomic radii increase.
    c. The reducing strength decreases across a period.
6. a. Elements in the first period have electrons that fill the first shell. This shell contains a maximum of 2 electrons.
    b. Elements in the second period have electrons that fill the second shell, 2 in the *s*-subshell and 6 in the *p*-subshell.
    c. The first series of transitions metals have electrons in the 3*d* subshell. The 3*d* subshell is filled after the 4*s* subshell and hence the first series of transition metals are in the fourth period.
7. a. Ca is more electronegative than K.
    b. Be is more electronegative than Ca.
    c. Cl is more electronegative than Br.

2.4 Review

2.4 Exercises

2.4 Exercise 1: Multiple choice questions

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

2.4 Exercise 2: Short answer questions

1. a. Cl and I
    b. Na
    c. Nickel is a transition metal.
    d. Cl is the most reactive non-metal.
    e. Nitrogen is the most electronegative.
2. a. Mendeleev, having collected a large amount of experimental data on the chemical properties of elements and their compounds, organised the data in such a way that he could conclude that the chemical properties vary periodically with increasing atomic weight.
    b. Mendeleev recognised that elements in the same group have similar chemical properties. Tellurium, Te, was placed before iodine so that tellurium was in a group with elements that shared similar chemical and physical properties to it.
3. All elements have a unique atomic number whereas the mass of an element is determined from the weighted average of the masses of all the naturally occurring isotopes of that element.

4. a. Na and K would be the most similar because they are in the same group. Na and Cl would be the most different because Na is a metal and Cl is a non-metal.
   b. Cl and I would be the most similar because they are in the same group. Cl and Mg would be the most different because Mg is a metal and Cl is a non-metal.
5. Electrons are further away from the attracting power of the nucleus.
6. a. $Z = 15$: $1s^2 2s^2 2p^6 3s^2 3p^3$
   b. Phosphorous is a non-metal.
   c. Group 15
   d. Period 3
   e. *p* block
   f. Nitrogen or arsenic
   g. Phosphorous would have a lower electronegatively than chlorine.
   h. Phosphorous has a larger atomic size than nitrogen.
7. In the modern form of the periodic table, elements are placed in order of increasing atomic number, not increasing atomic mass.
8. a. The nuclear charge of an element is equal to the charges on the nuclei of all atoms of that element.
   b. Calcium has the higher nuclear charge.
9. a. Atomic size decreases going across a period.
   b. Atomic size increases going down a group.
10. Elements with the same outer electron configuration are in the same group and so have similar properties; e.g. oxygen $1s^2 2s^2 2p^4$, sulfur $1s^2 2s^2 2p^6 3s^2 3p^4$ and selenium $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$.

2.4 Exercise 3: Exam practice questions

1. a. The electronegativity of an element is a measure of the degree to which an atom can attract electrons towards itself.
   b. Both ionisation energy and electronegativity relate to electrons. Ionisation energy is the energy required to remove an electron from a gaseous atom or ion (1 mark), while electronegativity provides a comparative measure of the ability of the nucleus of an atom to attract an electron to itself.
2. Metals have low electronegativities because they have a low effective nuclear charge and, therefore, do not strongly attract the outermost electrons. Non-metals, on the other hand, have high electronegativities because of their high effective nuclear charge and so more strongly attract outershell electrons.
3. a. *D* or *X* b. *X* c. *B* d. *J*
   e. *A* f. *G* or *J* g. *T*

3 Metals

3.2 Properties of metals

3.2 Exercise

1. a. $1s^2 2s^1$.
   b. Li^+, $1s^2$
   c. Solid lithium contains lithium cations in a ‘sea’ of delocalised valence electrons. Non-directional electrostatic forces of attraction hold the lattice together.
2. b (Sodium), d (Potassium), e (Manganese). a (Hydrogen) and c (Chlorine) are non-metals.
3. The structure of a metal is stable because metallic atoms, when losing their valence electrons to form cations, attain a group 18 electron configuration.
4. a. Iron is hard because its structure is very dense and the electrostatic attraction between the cations and delocalised electrons is very strong.
   b. When iron is hammered into a sheet (malleable) or drawn into a wire (ductile), layers of atoms move past one another without disrupting the force between the cations and the negative ‘sea’ of electrons.
   c. Iron can conduct electricity because under the force of an electric field, the ‘sea’ of delocalised, outershell electrons is mobile whether the metal is in solid or molten form.
5. a. The metallic lattice consists of magnesium cations, Mg^{2+}, in a ‘sea’ of delocalised valence electrons.
   b. The delocalised valence electrons within the lattice can reflect light.
6. The ‘delocalised sea’ of electrons model does not explain the differences in magnetism, density and strength of metals.
7. A metal that has large grains (perfect close packing) will have fewer dislocations and will bend more easily than a metal with small grains, which has many dislocations.
8. a. +1 b. +2
9. Properties of sodium and potassium are found in table 3.1.
   For Rb, we would expect its melting point to be lower than 64 °C, and its thermal conductivity to be less than 100 $m^{-1}K^{-1}$.
10. Aluminium is easily moulded and has an inert aluminium oxide coating which is insoluble in water.
11. Gold and silver are used for jewellery because they are soft enough to be easily worked, are lustrous and are unreactive so will not corrode easily.
12. a. Gold is a coloured metal.
    b. Mercury is a liquid at room temperature.
    c. Zinc compounds are white.
13. a. Iron (steel) — it has high strength and is cheap.
    b. Aluminium — it is strong and resistant to corrosion.
    c. Copper — it is ductile and a good conductor of electricity,
    d. Lead — it is soft and dense.
14. Aluminium is cheaper and lighter than copper.
15. a. They have a mobile ‘sea’ of delocalised outershell electrons that can flow between all of the cations.
    b. Their metallic lattice bonds with mobile electrons in all directions.
    c. The attractive forces of the cations and the mobile ‘sea’ of delocalised electrons occurs in all directions.
    d. Metals appear shiny as the mobile ‘sea’ of delocalised, outershell electrons allows light to be reflected.
    e. The high density of metals is due to their closely packed metallic lattices.

3.3 Reactivity of metals

Practice problem 1

Ensuring that calcium is oil-free allows the water to come into contact with the pure metal, allowing a reaction to occur.

Practice problem 2

1. Hydrogen gas
2. Magnesium nitrate
3. All of the nitric acid had been used up.

3.3 Exercise

1. $Ca(s) + H_2O(aq) \rightarrow Ca(OH)_2 + H_2(g)$
2. a. $4K(s) + O_2(g) \rightarrow 2K_2O(s)$
   b. $Cs(s) + 2H_2O(l) \rightarrow Cs(OH)_2(aq) + H_2(g)$
   c. $Sn(s) + 2HNO_3(aq) \rightarrow Sn(NO_3)_2(aq) + H_2(g)$

3.4 Extraction of metal from its ore

3.4 Exercise

1. Magnetite: Fe_3O_4, iron(II, III) oxide
   Haematite: Fe_2O_3, iron(III) oxide
   Pyrite: FeS, iron(II) sulfide
   Goethite: FeO(OH) iron (III) oxide-hydroxide
   Limonite: $Fe_2O_3 \cdot nH_2O$, iron(III) oxide (*n*)hydrate
   Siderite: $FeCO_3$, iron(II) carbonate
2. Sample answer: Optimising the use of coal in blast furnaces and replacement by other fossil fuels, which can reduce carbon emissions but not eliminate them. Sulfur dioxides could be reduced through decreased use of oil.
3. Iron is used in all aspects of our daily lives. Detailed responses can be found in the worked solutions in your digital formats.

3.5 Modifying metals

3.5 Exercise

1. a. Tempering disrupts the metallic lattice, reducing brittleness while retaining hardness.
   b. Quenching is the process of plunging hot metal into water to rapidly cool it. Quenched metals are harder but more brittle. Annealing involves heating a metal in an inert gas to a moderate temperature, maintaining this temperature for several hours, and then cooling it slowly. Annealing results in softer, more ductile metals.
2. The work-hardened copper becomes more brittle; annealing it restores its ductility.
3. The metal paperclip has been work-hardened.
4. a. Heating 30% tin and 70% lead together.
   b. Substitutional
   c. Solder has a lower melting point than both tin and lead and, therefore, can be used to join metal parts together.
5. You would expect the melting point of the alloy to be over 200 °C. However, the alloy has a significantly lower melting temperature (70 °C), causing water to be released at a lower temperature. This shows that alloys can have different properties compared to the properties of constituent metals in the alloy.

3.6 Metallic nanomaterials

3.6 Exercise

1. a. Volume 1 mm^3: surface area 60 cm^2
   b. Volume 1 nm^3: surface area 60 000 000 cm^2
   c. 10 million times
2. The reactivity of a metal is dependent upon the size of its particles.
3. Quantum dots are semi-conducting nanocrystals that develop intense, long-lasting colours under UV light and light-emitting diodes such as lasers.

3.7 Review

3.7 Exercises

3.7 Exercise 1: Multiple choice questions

1. C.
2. C.
3. D.
4. A.
5. B.
6. B.
7. A.
8. B.
9. C.
10. B.
11. A.
12. D.
13. C.
14. B.
15. C.

3.7 Exercise 2: Short answer questions

1. *B*.
2. a. (lithium) and e. (calcium)
3. a. The metals are good conductors of heat and are corrosion resistant.
   b. Woks made from normal steel can rust on exposure to water after being washed. A coating of oil protects the steel from water and moisture.
   c. Responses will vary; for example, copper pots and aluminium tongs and spatulas.
4. a. Because gold is unreactive, it can be found in its natural state.
   b. Gold is a soft metal.
5. Silver bonds to the mercury and helps the amalgam set.
   Tin helps the amalgam stay free from tarnish.
   Mercury can be poured into the tooth easily.
   Copper and zinc give the amalgam strength.
   Gold will not tarnish or react with food or drink.
6. a. A metal is a pure substance, whereas an alloy is a mixture of metals or of a metal and a non-metal heated to a molten state and then cooled.
   b. Responses will vary depending on metal chosen.
7. a. Element — it contains only one type of atom.
   b. Alloy — it is a mixture of different metals.
   c. Element — it contains only one type of atom.
   d. Alloy — pure gold mixed with another metal (copper or silver).
   e. Alloy — it is a mixture of different metals.
   f. Element — it contains only one type of atom.
   g. Alloy — it is a mixture of different metals.
   h. Compound — it contains two or more different elements that are chemically bonded.
   i. Compound — it contains two or more different elements that are chemically bonded.
   j. Element — it contains only one type of atom.
   k. Alloy — it is a mixture of different metals.
   l. Alloy — it is a mixture of different metals.

m. Compound — it contains two or more different elements that are chemically bonded.

8. The properties of an alloy are different from the properties of the constituent metals.

9. a. *d* block **b.** *p* block **c.** *s* block **d.** *d* block
e. *d* block

10. Annealing is the process of heating a metal and then cooling it slowly. Tempering is the process of warming a quenched metal and allowing it to cool slowly.

11. Quenching is the process of cooling a red hot metal quickly by plunging into cold water.

12. a. Dish-draining rack on a sink
  b. Bike chain
  c. Tin-coated cans
  d. Galvanised steel roofing

13. To find the position of nickel in the activity series you could conduct an experiment where you observe its reaction with water or acid and compare the results with other metals with known positions on the activity series.

14.

	Nanogold	Bulk gold	Both	Neither
Reactive	√			
Magnetic	√			
Spreads through water	√			
Range of colours	√			
Good conductor			√	
very large surface area	√			
Insoluble		√		
Unreactive		√		
Does not conduct				√
Transition metal			√	

15. a. Iron is used in construction, machinery, transport and appliances — in any situation where strength and durability are required.
  b. Iron is malleable, ductile, has a high melting point and is hard.
  c. Iron is abundant, hard, has high tensile strength and is relatively easily worked.
  d. Iron is alloyed to improve properties
  e. 1. $C(s) + O_2(g) \rightarrow CO_2(g)$
    2. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
    3. $C(s) + CO_2(g) \rightarrow 2CO(g)$
    4. $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$

3.7 Exercise 3: Exam practice questions

1. a. C
  b. A
  c. B

2. a. Iron ores are abundant and iron is relatively easily extracted from its ore compared with some other metals.
  b. Stainless steel. It is more resistant to corrosion due to the addition of chromium.
  c. Uses of iron include construction, to make containers, to make vehicles, machinery and tools, and as a catalyst.

4 Ionic compounds

4.2 Structure and properties of ionic substances

Practice problem 1

a. Ga^{3+}, gallium ions
b. P^{3-}, phosphide ions

Practice problem 2

a. * See figure at foot of the page.

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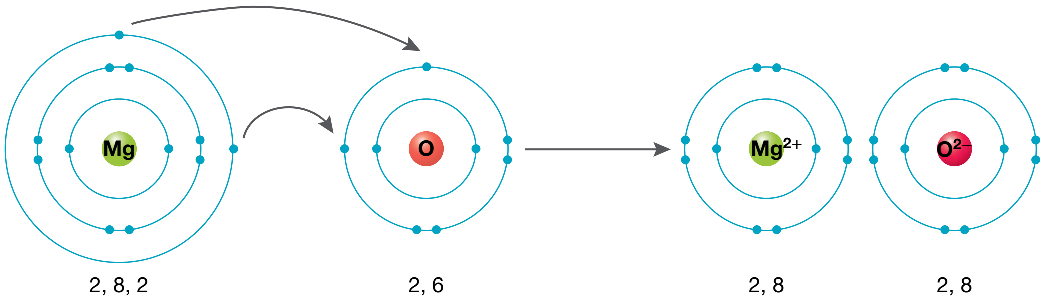
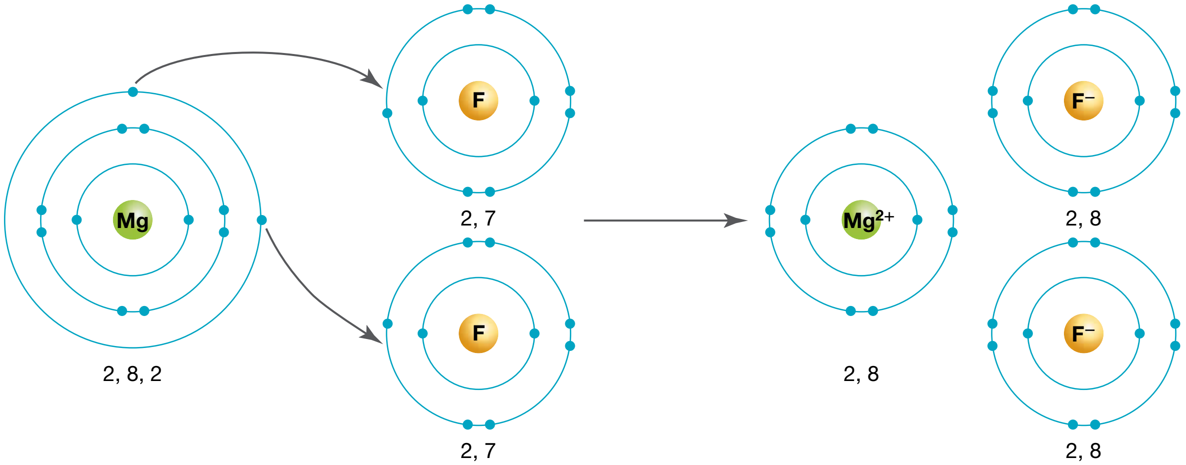
a*



$Mg^{2+} + O^{2-} \rightarrow MgO$

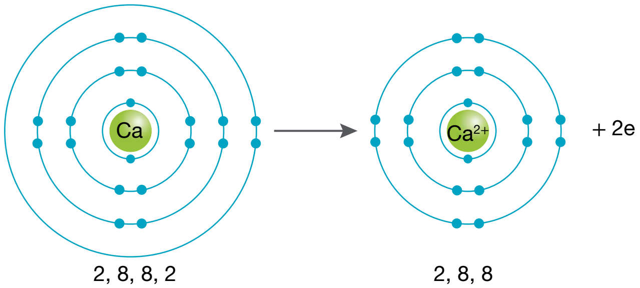
b.



$Mg^{2+} + 2F^{-} \rightarrow MgF^{2}$

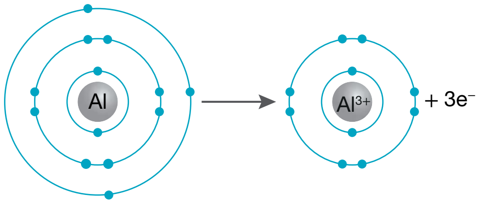
4.2 Exercise

1. Ca and Al each form an ion in order to attain a full outer shell and, hence, stability.
2. **a.**



Calcium Ca → $Ca^{2+} + 2e^{-}$

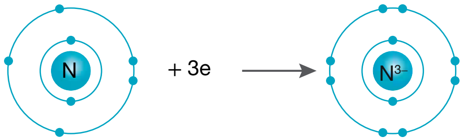
2, 8, 8, 2 2, 8, 8



Aluminium Al → $Al^{3+} + 3e^{-}$

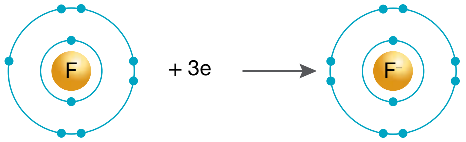
2, 8, 3 2, 8

b.



Nitrogen $N + 3e^{-}$ → N^{3-}(nitride)

2, 5 2, 8



Fluorine $F + e^{-}$ → F^{-}(fluoride)

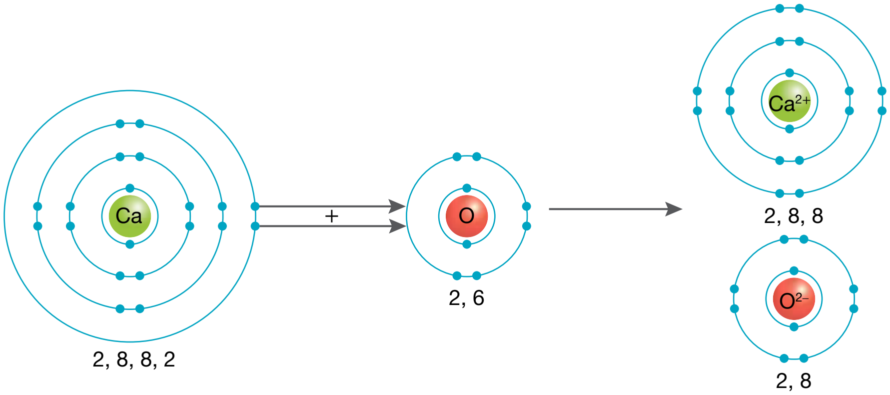
2, 7 2, 8

3. **a.** Barium ion, Ba^{2+}
b. Potassium ion, K^{+}
c. Phosphide ion, P^{3-}
d. Chloride ion, Cl^{-}
e. Sulfide ion, S^{2-}

4.

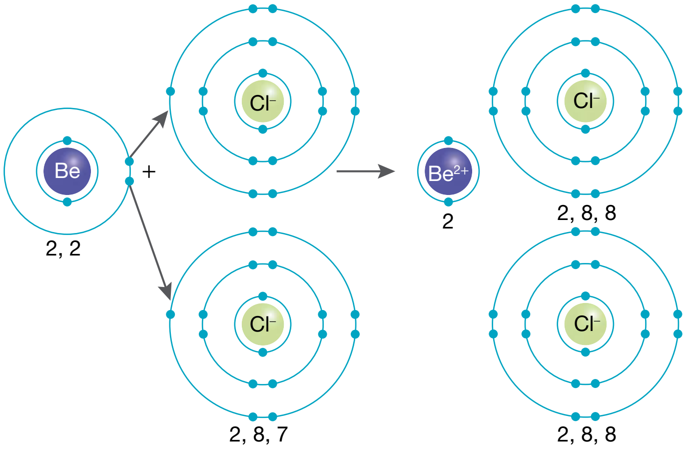
Period	Group 1	Group 2	Group 13	Group 15	Group 16	Group 17
2	1+	2+	3+	3–	2–	1–
3	1+	2+	3+	3–	2–	1–

5. **a.**



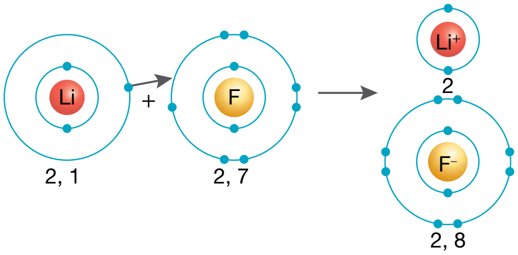
$Ca^{2+} + O^{2-} \rightarrow CaO$

b.



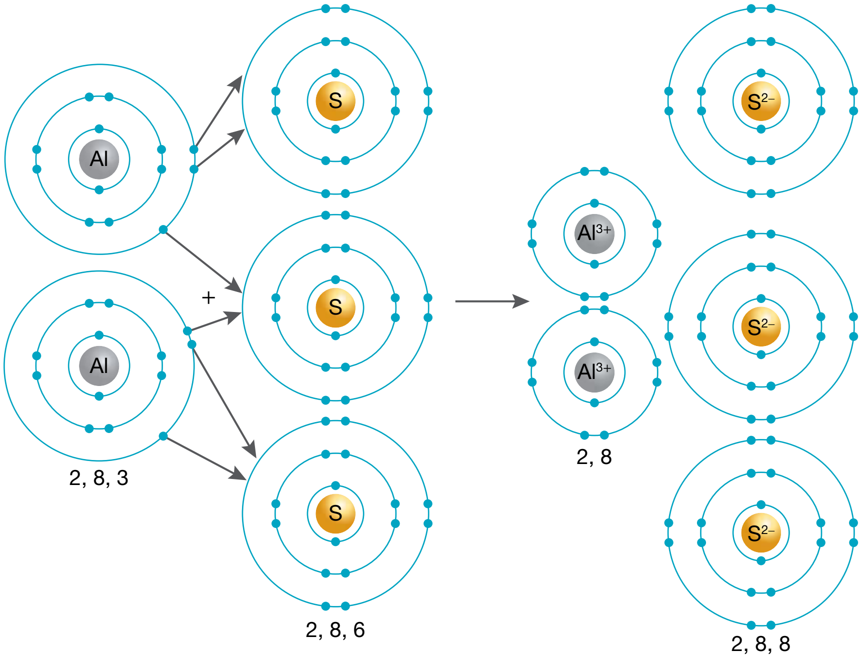
$Be^{2+} + 2Cl^{-} \rightarrow BeCl_2$

c.



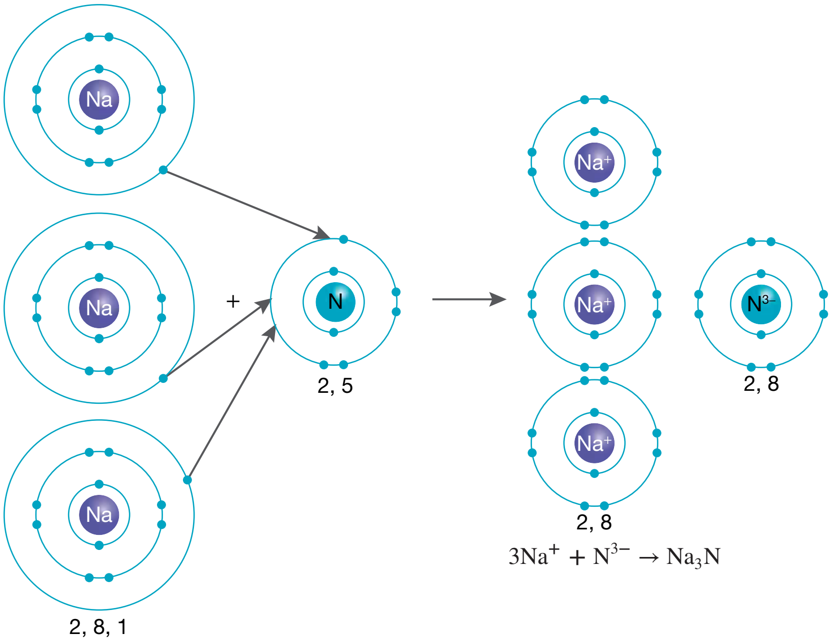
$Li^{+} + F^{-} \rightarrow LiF$

d.



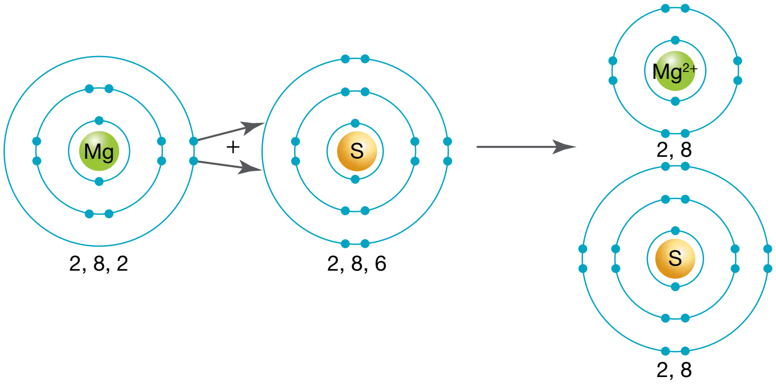
$2Al^{3+} + 3S^{2-} \rightarrow Al_2S_3$

e.



$3Na^{+} + N^{3-} \rightarrow Na_3N$

f.



$Mg^{2+} + S^{2-} \rightarrow MgS$

6. **a.** K^+: 2, 8, 8.
Ca^{2+}: 2, 8, 8.
Al^{3+}: 2, 8

b. F^-: 2, 8.
O^{2-}: 2, 8.
N^{3-}: 2, 8

7. Neon; Ne

8. **a.** The ions Mg^{2+} and $2Cl^-$ make up $MgCl_2$.
The ions K^+ and Cl^- make up KCl.
  b. Two Cl^- ions are needed to balance the one Mg^{2+} ion.
One K^+ ion balances one Cl^- ion.
  c. Ionic bonds

9. **a.** Calcium chloride has a high melting point because it is an ionic compound with strong electrostatic attractive forces holding the ions together in a three-dimensional lattice.
  b. Calcium chloride will shatter when pressure is applied because ions of like charge repel when aligned with one another.
  c. Calcium chloride will conduct electricity only in the liquid state where the ions are free to move. In the solid state, the ions are held strongly in place by the strong ionic bond and so will not conduct electricity as there are no mobile charged particles.

10. **a.** **i.** NaCl has one Na^+ ion for each Cl^- ion; therefore, the ratio is 1:1.
    ii. MgO has one Mg^{2+} ion for each O^{2-} ion; therefore, the ratio is 1:1.
  b. MgO has a higher melting point because the electrostatic attraction is greater between its ions due to their higher charges.

11. The substance is likely to be ionically bonded. Detailed responses can be found in the worked solutions in your digital formats.

12. **a.** X_2Y
  b. Ionic bond
  c. Ionic compounds are typically hard, brittle and have a high melting point.

4.3 Ionic nomenclature

Practice problem 3

Magnesium fluoride.

Practice problem 4

a. Strontium iodide. **b.** V_2O_5.

Practice problem 5

$Na_2Cr_2O_7$

4.3 Exercise

1. **a.** $AlCl_3$ **b.** BaO **c.** Na_2S **d.** Mg_3P_2

2. **a.**

Ions	K^+	Ca^{2+}	Al^{3+}
F^-	KF	CaF_2	AlF_3
O^{2-}	K_2O	CaO	Al_2O_3
N^{3-}	K_3N	Ca_3N_2	AlN

  b. KF: potassium fluoride
$CaF_{2:}$ calcium fluoride
AlF_3: aluminium fluoride
K_2O: potassium oxide
CaO: calcium oxide
Al_2O_3: aluminium oxide
K_3N: potassium nitride
Ca_3N_2: calcium nitride
AlN: aluminium nitride

3. **a.** Potassium chloride **b.** Silver sulfide
  c. Sodium nitride

4. **a.** Iron(II) sulfide **b.** Iron(III) sulfide
  c. Copper(II) chloride

5. **a.** Tin(II) oxide **b.** Copper(I) oxide
  c. Lead(II) bromide

6. **a.** SnF_4 **b.** PbS **c.** HgO
  d. FeN **e.** Cu_2S **f.** SnO

7. HSO_4^-, hydrogen sulfate
HSO_3^-, hydrogen sulfite
SO_4^{2-}, sulfate
SO_3^{2-}, sulfite
$S_2O_3^{2-}$, thiosulfate

8.

Cation	Anion	Empirical formula
Ag^+	Cl^-	AgCl
K^+	S^{2-}	K_2S
Mg^{2+}	O^{2-}	MgO
Al^{3+}	Br^-	$AlBr_3$
Fe^{3+}	CO_3^{2-}	$Fe_2(CO_3)_3$
Ba^{2+}	PO_4^{3-}	$Ba_3(PO_4)_2$
NH_4^+	SO_4^{2-}	$(NH_4)_2SO_4$

9. **a.** Na_2SO_3 **b.** $Ca(NO_2)_2$ **c.** $Cu(HCO_3)_2$

10. **a.** $Al_2(CO_3)_3$,
aluminium carbonate
  b. $NaNO_3$,
sodium nitrate
  c. $Hg_3(PO_4)_2$,
mercury(II) phosphate
  d. $PbSO_4$,
lead(II) sulfate

4.4 Crystal formation and uses of ionic compounds

Practice problem 6

a. Cobalt(II) bromide hexahydrate
b. $SnCl_2{\cdot}2H_2O$

4.4 Exercise

1. **a.** $Na_2CO_3{\cdot}10H_2O$
  b. Magnesium sulfate heptahydrate
  c. $MgCl_2{\cdot}6H_2O$
  d. Barium chloride dihydrate

4.5 Review

4.5 Exercises

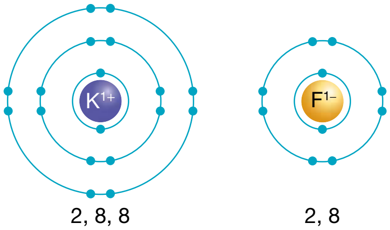
4.5 Exercise 1: Multiple choice questions

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

12. C.
13. C.

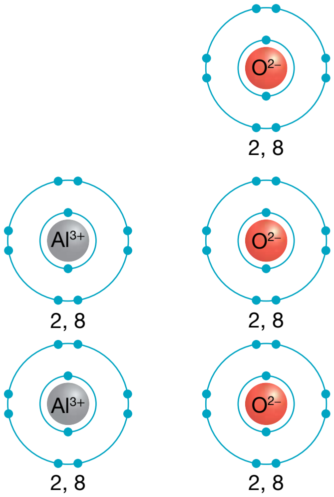
4.5 Exercise 2: Short answer questions

1. See table at foot of the page.*
2. a. 2, 8, 8, 2
   b. 2, 7
   c. Each of two fluorine atoms accepts one electron from a calcium atom, resulting in two fluoride ions, $2F^-$, and one calcium ion, Ca^{2+}.
   d. See figure at foot of the page.*
   e. Ionic bond
   f. Ca:F = 1:2
3. Mg (2, 8, 2) → Mg^{2+} (2, 8) + $2e^-$
4. KCl is composed of K^+ and Cl^-. Potassium oxide is composed of K^+ and O^{2-}. Because two K^+ ions are necessary to balance the O^{2-} charge, the ratio of the different ions in potassium oxide is 2:1.
5. a.


Potassium fluoride

$K^+ + F^- \rightarrow KF$
2, 8, 8 + 2, 8

b.

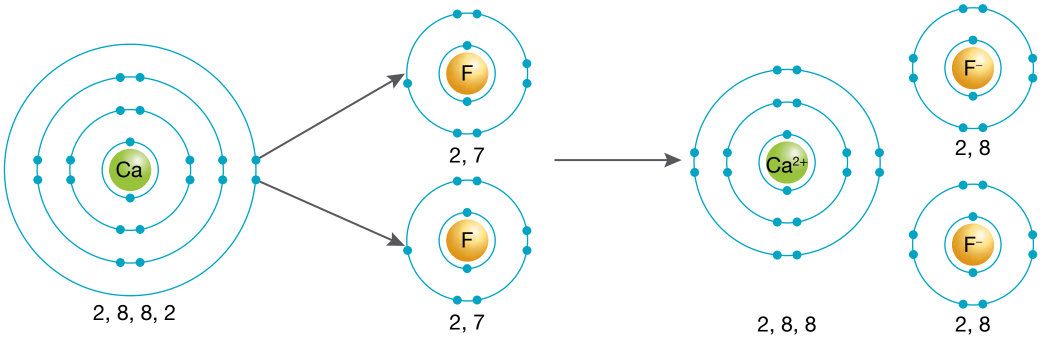

Aluminium oxide

$2Al^{3+} + 3O^{2-} \rightarrow Al_2O_3$
2, 8 + 2, 8

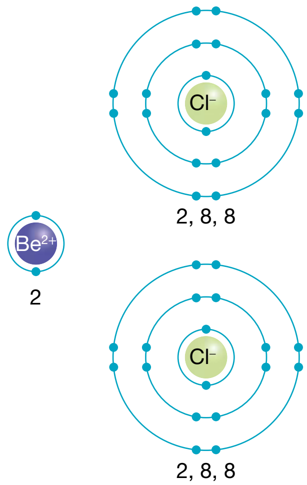
1*

Name of atom	Symbol for atom	Electron configuration of atom	Symbol for ion	Electron configuration of ion
Lithium	Li	2, 1	Li^+	2
Beryllium	Be	2, 2	Be^{2+}	2
Nitrogen	N	2, 5	N^{3-}	2, 8
Oxygen	O	2, 6	O^{2-}	2, 8
Fluorine	F	2, 7	F^-	2, 8
Sodium	Na	2, 8, 1	Na^+	2, 8
Magnesium	Mg	2, 8, 2	Mg^{2+}	2, 8
Aluminium	Al	2, 8, 3	Al^{3+}	2, 8
Phosphorus	P	2, 8, 5	P^{3-}	2, 8, 8
Sulfur	S	2, 8, 6	S^{2-}	2, 8, 8
Chlorine	Cl	2, 8, 7	Cl^-	2, 8, 8
Potassium	K	2, 8, 8, 1	K^+	2, 8, 8
Calcium	Ca	2, 8, 8, 2	Ca^{2+}	2, 8, 8

2d*



c. **Beryllium chloride**



Be^{2+} + $2Cl^{-} \rightarrow BeCl_2$
2 + 2, 8, 8

6. Ionic salts conduct electricity in the molten and aqueous states because the ions have dissociated and are now mobile. In the solid state, they are held rigidly by strong ionic bonds and, therefore, cannot conduct electricity.
7. See figure at foot of the page.*
8. Ionic substances are usually solids and have high melting points due to the strong attractive electrostatic forces holding the ions together in a lattice.
9. **a.** *XY*
   b. The bonding between *X* and *Y* must be ionic
   c. $X^{2+} + Y^{2-} \rightarrow XY$
10. **a.** NaF **b.** $Ba(NO_3)_2$ **c.** $Fe(OH)_3$ **d.** Na_2S
    e. Al_2O_3 **f.** CaH_2 **g.** $CuSO_4$ **h.** NH_4OH
    i. Cr_2O_3 **j.** $Ca(NO_3)_2$ **k.** LiCl **l.** KCN
    m. Na_2HPO_4
11. **a.** Zinc chloride **b.** Aluminium carbonate
    c. Sodium sulfate **d.** Silver nitrate
    e. Sodium hydroxide
12. **a.** Na_2CO_3
    Sodium carbonate
    b. $NaHCO_3$
    Sodium hydrogen carbonate
    c. $CaCO_3$
    Calcium carbonate
13. **a.** $MgSO_4{\cdot}7H_2O$ **b.** $Na_2CO_3{\cdot}10H_2O$.
    c. $ZnCl_2{\cdot}6H_2O$. **d.** $BaCl_2{\cdot}2H_2O$.
14. **a.** Barium chloride trihydrate
    b. Lithium chloride tetrahydrate
    c. Cobalt(II) chloride pentahydrate

4.5 Exercise 3: Exam practice questions

1. **a.** 2^{+} **b.** 3^{-} **c.** 3^{+}.
2. Oxygen atoms have the electron configuration 2, 6. For two atoms of oxygen to bond ionically, each would need to gain two electrons to attain a full outer shell. This is not possible.
3. **a.** When a metal atom forms a cation, it loses electrons and, therefore, the effective nuclear charge increases because the remaining electrons will be pulled in with a greater force by the nucleus. This will decrease the size of the metal ion when compared to the metal atom.
   b. When a non-metal atom forms an anion, it gains electrons and, therefore, the effective nuclear charge decreases because the remaining electrons will be pulled in with a weaker force by the nucleus. This will increase the size of the non-metal ion when compared to the non-metal atom.

5 Quantifying atoms and compounds

5.2 Relative isotopic mass and the carbon-12 scale

Practice problem 1

a. 309.3
b. 195.8
c. 262.0

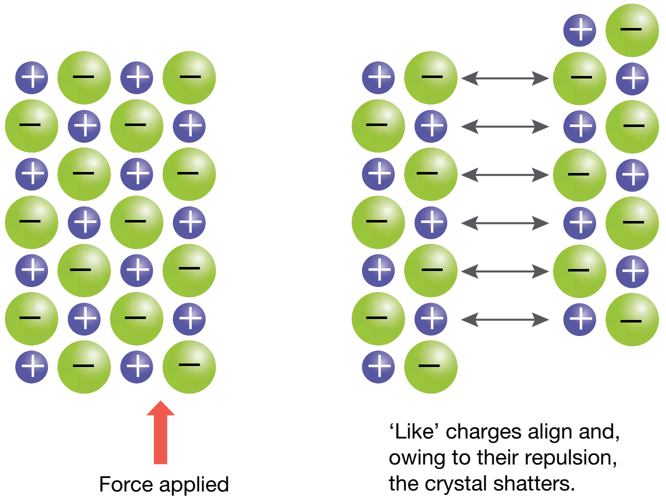
Practice problem 2

$A_r = 63.5$

5.2 Exercise

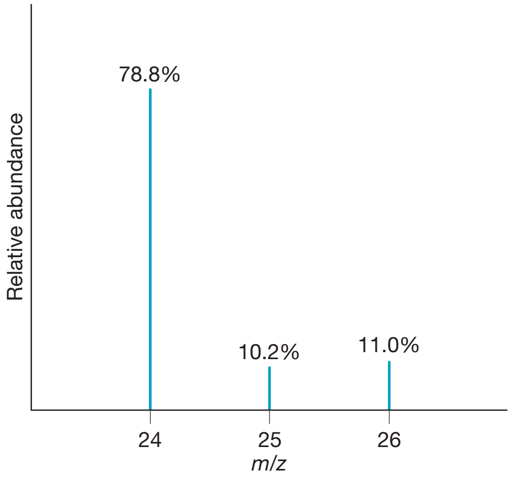
1. **a.** 44.0 **b.** 58.5 **c.** 34.0
   d. 98.1 **e.** 180.0

7*



'Like' charges align and, owing to their repulsion, the crystal shatters.

2. a.



b. $A_r(Mg) = 24.3$

3. 71 (2 significant figures)

4. $Ar(B) = 11$ (2 significant figures)

5.3 The mole

Practice problem 3

$M_r = 239.1$ g mol^{-1}

Practice problem 4

12×10^{23} atoms of oxygen

5.3 Exercise

1. a. 48.0 g mol^{-1} b. 253.8 g mol^{-1} c. 56.1 g mol^{-1}
   d. 36.5 g mol^{-1}
2. a. 1.57×10^3 g b. 0.140 g c. 2.60×10^{-8} g
   d. 5.23×10^6 g e. 5.53×10^2 g
3. 200 g of Mg (option a) has the greatest mass.
4. See table at foot of the page.*

5.4 Determination of the empirical formula

Practice problem 5

$Na_2CO_3 \cdot 10H_2O$.

5.4 Exercise

1. C_2OH_2
2. C_3OH_3
3. $CaSO_4 \cdot 2H_2O$
4. $MgSO_4 \cdot 7H_2O$
5. $Na_2B_4O_7$

5.5 Review

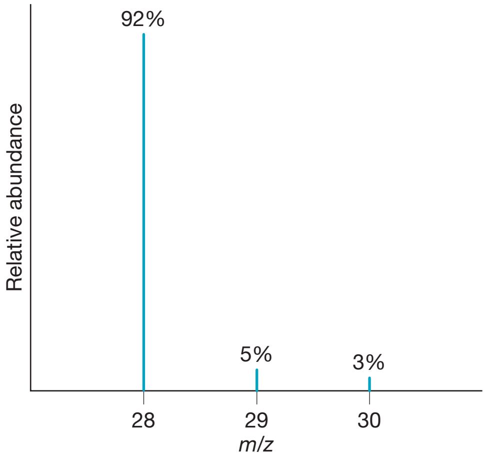
5.5 Exercises

5.5 Exercise 1: Multiple choice questions

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

5.5 Exercise 2: Short answer questions

1. a.



4*

	Substance	Amount (g)	Molar mass (M)	Number of atoms in the molecule	Number of moles (n) of substance	Number of molecules	Total number of atoms
1	Water, H_2O	3.2	18.0	3	0.18	1.1×10^{23}	3.3×10^{23}
2	Methane, CH_4	2.7	16.0	5	0.17	1.0×10^{23}	5.1×10^{23}
3	Ammonia, NH_3	0.056	17.0	4	0.0033	2.0×10^{21}	7.9×10^{21}
4	Ethanoic acid, CH_3COOH	1.64×10^3	60.0	8	27.3	1.64×10^{25}	1.31×10^{26}
5	Benzene, C_6H_6	44	78.0	12	0.56	3.4×10^{23}	4.0×10^{24}
6	Octane, C_8H_{18}	267	114.0	26	2.34	1.41×10^{24}	3.66×10^{25}
7	Ethanol, CH_3CH_2OH	4.6×10^2	46.0	9	10	6.0×10^{24}	5.4×10^{25}
8	Ozone, O_3	0.101	48.0	3	0.002 11	1.27×10^{21}	3.81×10^{21}
9	Sulfuric acid, H_2SO_4	7.0×10^3	98.0	7	71	4.3×10^{25}	3.0×10^{26}
10	Carbon dioxide, CO_2	1.8×10^3	44.0	3	42	2.5×10^{25}	7.5×10^{25}

- **b.** $Z = 14$
- **c.** Approximately 28, 29 and 30
- **d.** A_r (Si) $= 28$ (to 2 significant figures)

2. The abundance of the lighter isotope is 55%.

3.
- **a.** 6.0 mol nitrogen atoms
- **b.** 24.0 mol hydrogen atoms
- **c.** 8.0 mol oxygen atoms
- **d.** 2.0 mol phosphorus atoms
- **e.** 3.6×10^{24} nitrogen atoms
- **f.** 1.9×10^{2} g of phosphate ions
- **g.** 84.0 g of nitrogen atoms

4. See table at foot of the page.*

5.
- **a.** 1.35×10^{23} atoms
- **b.** 7.2×10^{22} atoms
- **c.** 5.71×10^{25} atoms
- **d.** 3.5×10^{23} atoms
- **e.** 3.2×10^{23} atoms

6.
- **a.** 8.93×10^{3} g
- **b.** 4.5 g
- **c.** 194 g

7.
- **a.** 0.000 31 mol
- **b.** 1.9×10^{20} ions

8.
- **a.** 3.4×10^{-4} mol
- **b.** 2.1×10^{20} molecules
- **c.** 5.0×10^{4} g (50 kg)

9.
- **a.** 2.841×10^{-3} mol
- **b.** 1.71×10^{21} molecules

10.
- **a.** 0.648 mol
- **b.** 17.5 mol
- **c.** 210 g
- **d.** 10.4 g

11.
- **a.** 1.00×10^{22} atoms
- **b.** 3.16×10^{22} atoms

12. $C_4H_5N_2O$

13. $C_3H_7NO_2S$

14. COH_4

5.5 Exercise 3: Exam practice questions

1.
- **a.** 114.0
- **b.** 365 g
- **c.** 1.3×10^{22}
- **d.** 6.9×10^{23} atoms
- **e.** 16%
- **f.** 126 g

2. C_2H_4O

3. Cu_2O

4. $x = 7$

5. $C_{19}OH_{38}$

6. Mass number = total number of protons + neutrons in the nucleus of an atom. For example, a particular neutral atom of carbon has six protons and eight neutrons. Therefore, it has a mass number of 14 (and an atomic number of six, and six electrons). The relative atomic mass of neon, A_r(Ne), is equal to the weighted average of the masses of neon's naturally occurring isotopes on a scale where the ^{12}C atom has a mass of 12 exactly. For example, Ne has two isotopes, ^{20}Ne ($M = 20.0$) with an abundance of 90.0% and ^{22}Ne with an abundance of 10.0%.

Consequently,

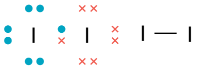
$$A_r\,(\text{Ne}) = \frac{(90.0 \times 20.0)\ +\ (10.0 \times 22.0)}{100} = 20.2$$

Mass numbers can be used as an approximation for relative isotopic masses because a proton and a neutron both have a relative mass very close to one. The molar mass, M, of an element is equal to the mass of one mole of the element; e.g. $M(\text{Ne}) = 20.2$ g mol^{-1}.

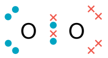
6 Materials from molecules

6.2 Representing molecules

Practice problem 1



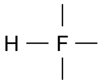
Practice problem 2



Practice problem 3

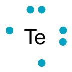
a.

b.



6.2 Exercise

1. **a.**



b.

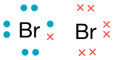
4*

Name	Formula	Molar mass (M)	Mole (n)	Mass (m)
Sodium hydroxide	NaOH	40.0 g mol^{-1}	0.0853	3.41 g
Ccarbon tetrachloride	CCl_4	154.0 g mol^{-1}	1.40	216 g
Sodium carbonate	Na_2CO_3	106.0 g mol^{-1}	1.00	106 g
Potassium chloride	KCl	74.6 g mol^{-1}	0.25	19 g
Ammonium phosphate	$(NH_4)_3PO_4$	149.0 g mol^{-1}	0.0568	8.46 g

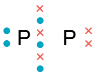
2. See table at foot of the page.*

3. Iodine, I_2

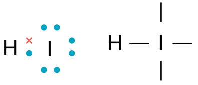
4.



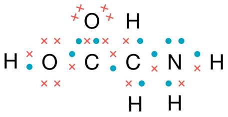
5.



6.

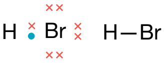


7. a.

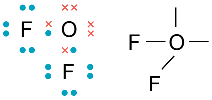


b. Showing this molecule as an electron dot diagram is cumbersome and slow to draw, and errors are easy to make. (*Note:* other answers are possible.)

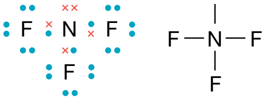
8. a.



b.

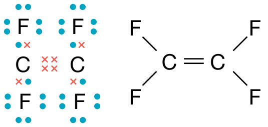


9.

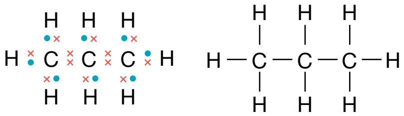


10. a.

b.



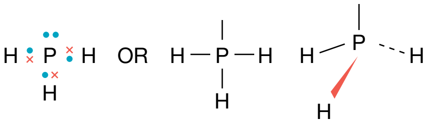
c.



d.

6.3 Predicting molecular shape

Practice problem 4



Note: the distribution of electron pairs around a central atom and molecular shape are not necessarily the same.

Practice problem 5

$H-C\equiv C-H$

Practice problem 6

a. The bonds are non-polar.

b. The bonds are polar.

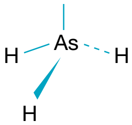
Practice problem 7

a. Non-polar

b. Polar

6.3 Exercise

1.



2. a. Sulfur to oxygen bonds are polar due to the electronegativity difference and lack of symmetry.

b. Bromine to bromine bonds are non-polar because the electronegativities are equal and they are symmetrical.

3. a. Cl_2 is non-polar because no difference exists between electronegativities of the atoms involved.

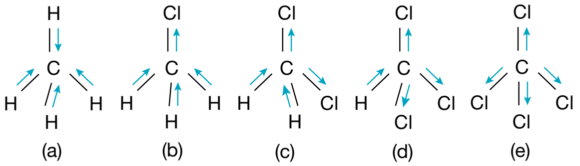
b. NH_3 is polar as the electronegativity difference between the nitrogen and hydrogen atoms causes dipoles that do not cancel as the molecule is not symmetrical.

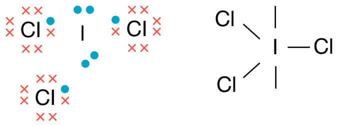
4. a. $\overset{\delta-}{H}-\overset{\delta+}{F}$ **b.** $\overset{\delta-}{O}-\overset{\delta+}{H}$ **c.** $\overset{\delta-}{C}-\overset{\delta+}{H}$

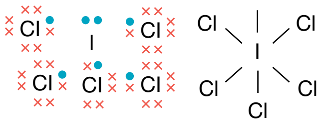
d. $\overset{\delta-}{N}-\overset{\delta+}{H}$ **e.** $C-C$

2*

Group	Element	Valence electrons	Dot diagram	Lone pairs	Bonding electrons
16	S	6		2	2
14	Si	4		0	4
15	P	5		1	3
17	Cl	7		3	1
17	Br	7		3	1
16	Se	6		2	2
18	Ar	8		4	0

5. a. HI has polar covalent bonding because no metals are present, and a difference exists in electronegativities.
   b. KCl has ionic bonding because it involves a metal and a non-metal.
   c. F_2 has non-polar covalent bonding because electrons are shared but no difference exists in electronegativities.
6. N–N, Cl–N, Li–F.
   Note: the representation Li–F should be Li^+F^-.
7. A large electronegativity difference means that the 'sharing' is so unequal it is essentially a 'donation' from one atom to the other. Positive and negative ions are the result. However, a small electronegativity difference means that a large amount of sharing is still occurring — it is just slightly uneven in favour of the more electronegative atom.
8. A tetrahedral shape does not necessarily mean the molecule is non-polar. For example, the molecules CH_4 and CH_3Cl are both tetrahedral. The former is non-polar while the latter is polar. It is better to consider whether the molecule is symmetrical or not.
9. 

   In the structures the bond dipoles cancel out for (a) and (e), so they are non-polar. In structures (b), (c) and (d), they do not cancel out so these molecules are polar.
10. a. 

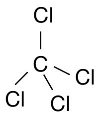
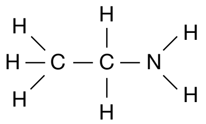
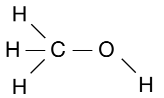
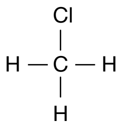
   b. The outer shell of iodine (shell number 5) can hold more than eight electrons.
   Note: part (a) represents a question at the upper levels of difficulty for Unit 1 VCE Chemistry.

6.4 Explaining the properties of molecular substances

Practice problem 8

a. CI_4 is a non-polar tetrahedral molecule. Only dispersion forces exist between the molecules.
b. CH_3NH_2 has a tetrahedral shape around the carbon atom and a trigonal planar shape around the second central atom, nitrogen, resulting in a polar molecule. It also has a highly electronegativity atom (nitrogen) bonded to hydrogen atoms. So, dispersion forces and hydrogen bonding exist between these molecules.

6.4 Exercise

1. a. 
   CCl_4 is a non-polar molecule, so the only forces between molecules will be dispersion forces.
   b. 
   $CH_3CH_2NH_2$ is a polar molecule, with nitrogen and hydrogen atoms. Therefore, the forces between molecules are dispersion forces and hydrogen bonding.
   c. Cl — Cl
   Cl_2 is a non-polar molecule, so the only forces between molecules will be dispersion forces.
   d. 
   CH_3OH is a polar molecule, with oxygen and hydrogen atoms. Therefore, the forces between molecules are dispersion forces and hydrogen bonding.
   e. 
   CH_3Cl is a polar molecule containing chlorine and has an overall dipole. It does not have hydrogen bonding. Therefore, the forces between molecules will be dispersion forces and dipole–dipole attractions.
2. Water has an extra form of intermolecular bonding (hydrogen bonding) between its molecules that H_2S does not. This makes the attractions between the molecules stronger and increases the boiling point.
3. Although dispersion and dipole–dipole intermolecular forces exist in all three substances, as the number of electrons in each molecule increases (from 18 in HCl, to 36 in HBr and 54 in HI), so does the strength of the dispersion forces. This is reflected in the increasing boiling points.
4. Br_2 is a non-polar molecule and has only dispersion forces between its molecules, ICl is a polar molecule and has dipole-dipole forces as well as dispersion forces between its molecules. This extra force causes the higher melting point as more heat energy is required to separate the molecules.
5. Cl_2 has the higher boiling point because dispersion forces increase with increasing number of electrons.
6. Helium atoms have only two electrons while argon atoms have 18. Thus, the dispersion forces acting on argon atoms are stronger and its boiling point is higher.
7. HBr is a polar molecule and experiences dipole–dipole forces as well as dispersion forces, while Kr experiences only dispersion forces.
8. HF has hydrogen bonding, which is stronger than other dipole–dipole forces.
9. They are all gases because only weak dispersion forces act between the molecules.

10. I_2 is a non-polar substance and, therefore, will dissolve in non-polar carbon tetrachloride more readily than in polar water.

11. a. Candle wax is relatively soft and can be scratched because it is a non-polar covalent molecular compound with weak dispersion forces between its molecules.
    b. Candle wax is not soluble in water because it is non-polar but water is polar.
    c. Candle wax is a non-conductor because it has no free-moving charged particles (electrons or ions).

12. a. Hydrogen bonding is an intermolecular force; it acts between molecules.
    b. Hydrogen bonding occurs when H atoms are bonded to N, O or F *within* a molecule. The large electronegativity difference between H and the atoms of N, O or F lead to polar bonds and polar molecules overall. If a molecule is polar, attractions will exist *between* it and other nearby polar molecules due to the small $\delta+$ and $\delta-$ charges that they contain.
    c. For hydrogen bonding to occur, a hydrogen atom must be covalently bonded to either an oxygen, nitrogen or fluorine atom within a molecule.

6.5 The relative strengths of bond types in molecular substances

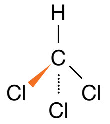
Practice problem 9

a. HF
b. HI
c. HF
d. HF

6.5 Exercise

1. dispersion forces < dipole–dipole attraction < hydrogen bonding < covalent bonding
2. a. No b. CF_4 c. CBr_4 d. CBr_4
3. a. Hydrogen bonding does not exist in any of these molecules.
   b. CH_3Br
   c. CH_3Br
4. Boiling points should be about the same.
5. a. O = O
   Linear

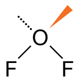
   b.



   Tetrahedral

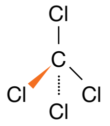
   c. Br — Br
   Linear

   d.



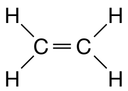
   V-shaped

   e.



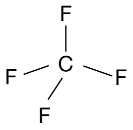
   Tetrahedral

   f.



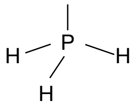
   Trigonal planar around each C atom

6. a.



   Tetrahedral

   b.



   Pyramidal

   c. H — C
   Linear

7. a. Polar b. Non-polar c. Non-polar
   d. Polar e. Polar f. Non-polar
8. CH_3F has dispersion and dipole–dipole forces acting between its molecules; however, dispersion forces and the stronger hydrogen bonding between molecules of CH_3OH cause it to have a higher boiling point.
9. H_2O, CH_3OH and CH_3CH_2OH exhibit hydrogen bonding because they are polar molecules that contain hydrogen bonded to oxygen. C_2H_6 does not have oxygen bonded to hydrogen.
10. Water expands on freezing. This is due to hydrogen bonding making hexagonal crystals. The water molecules are further apart in this structure than they are in the liquid. Thus the same quantity of water expands when it freezes, cracking the glass bottle as it does so.

6.6 Review

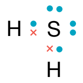
6.6 Exercises

6.6 Exercise 1: Multiple choice questions

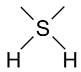
1. D.
2. A.
3. B.
4. D.
5. A.
6. A.
7. A.
8. D.
9. B.
10. C.
11. A.
12. A.
13. B.
14. A.
15. B.

6.6 Exercise 2: Short answer questions

1. Covalent molecular substances have low melting and boiling temperatures, do not conduct electricity and are soft.
2. a. Eight
   b. Group 18
   c. Noble gases
   d. The outer shell of He can only hold two electrons.
3. a. Atoms will bond in such a way as to obtain eight electrons in their outer shell.
   b. No; exceptions do exist to the octet rule.
4. a. Group 14 b. Group 18
   c. Group 17 d. Group 15
5. a.

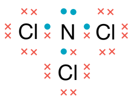


Electron dot diagram

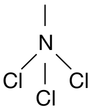


Structural formula

   b.

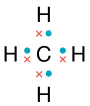


Electron dot diagram

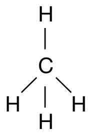


Structural formula

   c.

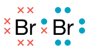


Electron dot diagram



Structural formula

   d.

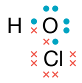


Electron dot diagram

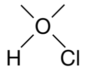
Br — Br

Structural formula

e.



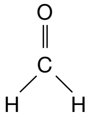
Electron dot diagram



Structural formula

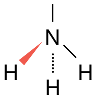
f.

Electron dot diagram



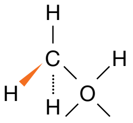
Structural formula

6. a.



N has a single lone pair, creating a pyramidal shape.

b.

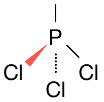


C has no lone pairs, so its structure is tetrahedral around the C atom; oxygen has two lone pairs, and hence structure is V-shaped around O atom.

c. H—C≡N

The triple bond between C and N creates a linear structure.

d.

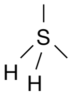


P has a single lone pair, creating a pyramidal shape.

e. —N≡N—

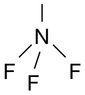
The triple bond between the N atoms creates a linear structure.

7. a.



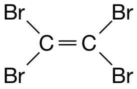
Sulfur has two lone pairs so structure is V-shaped.

b.



N has a single lone pair, creating a pyramidal shape.

c.



The double bond between the C atoms and the distribution of the Br atoms about each C creates a trigonal planar around each C atom.

8. a. Non-polar covalent **b.** Polar covalent
c. Non-polar covalent **d.** Polar covalent
e. Ionic

9. a. Responses will vary. Sample answer:

$\overset{\delta+}{H}-\overset{\delta-}{Cl}$

Hydrogen chloride molecule
Chlorine has a higher electronegativity than hydrogen and hence the two electrons in the bonding pair are not equally shared. They are more attracted to the Cl atom, resulting in the formation of a partial negative charge on the Cl and a partial positive charge on the H.

b. Each atom has the same attraction for the bonding pair of electrons (There is no difference in electronegativity).

c. Polarity depends on the shape of the molecule as well as bond dipoles. If the bond dipoles cancel each other out due to the shape of the molecule, it will be a non-polar molecule.

10. Polar: NH_3, CH_3OH, HCN, PCl_3, H_2S, and NF_3
N_2 and C_2Br_4 are non-polar.

11. Ammonia, NH_3, has a molecular dipole due to its pyramidal shape. Carbon dioxide, CO_2, is non-polar because its shape is linear and the bond dipoles cancel each other out.

12. a. Iodine has more electrons and hence stronger dispersion forces.

b. Fluorine is a gas at room temperature due to its weak dispersion forces. Iodine is a solid at room temperature because it has much stronger dispersion forces.

13. a. OCS has the stronger intermolecular forces (dispersion and dipole–dipole forces), while non-polar CO_2 has dispersion forces only.

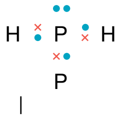
b. HF has both dispersion forces and hydrogen bonding, which is stronger than other dipole–dipole forces, while HBr has dispersion and dipole–dipole forces.

14. **c** and **d** exhibit hydrogen bonding

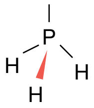
15. Although both methanol and ethanol have hydrogen bonding and dispersion forces, ethanol has the higher boiling point because it has more electrons and, therefore, stronger dispersion forces.

6.6 Exercise 3: Exam practice questions

1. a. Yes (one pair).



b.

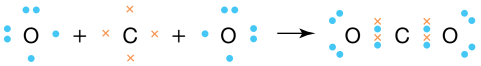


c. Pyramidal

d. Responses will vary but should include a logical explanation that mentions three bonds and one lone pair giving a tetrahedral arrangement around P but the shape, therefore, being described as pyramidal.

e. PH_3 is such a molecule. Four electron pairs may be made up of three bonding and one lone pair. These will achieve maximum separation in a tetrahedral configuration around the central atom. As they do this, the bonding pairs take the other atoms they are joined to with them. The arrangement of *atoms* in the molecule is therefore pyramidal.

2. a.



b. CO_2 is a linear molecule.

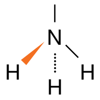
c. Four lone pairs — two on each oxygen atom.

d. Polar

e. Non-polar

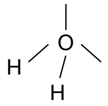
f. Dispersion forces

3. a. Pyramidal



b. Ammonia is a polar molecule because it is asymmetrical, and the ‘H end’ of the molecule is slightly positive while the ‘N end’ of the molecule is slightly negative.

c. While dispersion forces and hydrogen bonds exist between ammonia molecules, ammonia exists as a gas because, at room temperature, these forces of attraction can be overcome by the kinetic energy (energy of motion) of the ammonia molecules.

4. a. 

   There are four electron pairs that repel each other to give a tetrahedral arrangement of electron pairs around the central O atom.

   The repulsion from the two lone pairs results in the two bonding pairs being ‘pushed’ into a V-shape.

   b. Intramolecular: the H and O atoms are strongly covalently bonded *within* the molecule.

   Intermolecular: both dispersion forces and hydrogen bonds exist *between* the H_2O molecules.

   c. When ice freezes, each water molecule forms hydrogen bonds with four other water molecules.

   This results in ice forming a very open, widely spread hexagonal arrangement of water molecules, where the density of ice is less than that of liquid water.

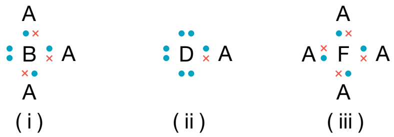
   d. Answers will vary. Sample answer:

   Half-fill a plastic drink bottle with water and mark the level of the water. Place the bottle upright in a freezer overnight. Remove the bottle from the freezer and mark the level of the ice. Comparing these two levels will show an approximately 10% increase in the volume of ice compared with water.

5. a. Elements D and E are in same group (group 17)

   b. C would represent any group 16 element

   c. A–C bonds

   d. 

   e. i. BA_3 and AD will be polar.

   ii. FA_4 will be non-polar.

   f. BA_3 is pyramidal.

7 Carbon lattices and carbon nanomaterials

7.2 Covalent lattice bonding

Practice problem 1

a. Silicon atoms will be covalently bonded to four other silicon atoms, which in turn will be bonded to four more each, and so on. The arrangement of bonds around each silicon atom will be tetrahedral.

b. Silicon is in period 3 compared to carbon in period 2 so its bonding electrons are further from the nucleus than carbon, which makes the bond length longer.

7.2 Exercise

1. a. Silicon atoms will be covalently bonded to four carbon atoms, which in turn will be bonded to four silicon atoms. This alternating pattern will continue throughout the crystal. In this way, the overall formula will be SiC (1:1 ratio). The arrangement of bonds around each carbon and each silicon atom will be tetrahedral.

   b. The bonds are longer in silicon dioxide than in diamond as silicon is in period 3 while carbon is in period 2.

2. Every carbon atom has four outershell electrons that can be used to form strong covalent bonds with other atoms. In doing so, carbon can form molecules containing single, double and triple bonds as well as straight chain, branched chain and cyclic structures. Carbon can also form layer and network structures and a more recently discovered form of carbon, fullerenes.

3. In diamond, each carbon atom is covalently bonded to four other carbon atoms in a three-dimensional network lattice. In graphite, each carbon atom forms covalent bonds with three other carbon atoms to form layers of hexagonal sheets. These layers are stacked on top of each other and are held together by dispersion forces. The distance between the layers is much greater than the carbon to carbon bond length in diamond. Hence, the atoms in diamond, being packed more closely than those in graphite, give diamond a greater density.

4. a. The weak dispersion forces between the layers of hexagonal sheets of carbon atoms in graphite allow the carbon atoms to be readily layered onto paper to form a mark.

   b. In diamond, each carbon atom is covalently bonded to four other carbon atoms in a three-dimensional network lattice. Hence, diamond is an exceptionally hard and strong material and is, therefore, useful for drilling.

5. In graphite. each carbon atom forms covalent bonds with three other carbon atoms to form layers of hexagonal sheets. One electron from each carbon atom is delocalised and is free to move across the sheets of carbon atoms. It can therefore conduct electricity. In diamond, each carbon atom is covalently bonded to four other carbon atoms in a three-dimensional network lattice; that is, all of carbon’s outershell electrons are localised in covalent bonds and no electrons are free to conduct electricity.

7.3 Carbon nanomaterials

Practice problem 2

100 nm

7.3 Exercise

1. a. 1 000 000 (10^6)   b. 1 000 000 000 (10^9)

   c. 1000

2. 2.5×10^4 nm

3. In graphite, each carbon atom forms covalent bonds with three other carbon atoms to form layers of hexagonally bonded atoms. These layers are stacked on top of each other and are held together by dispersion forces. Carbon nanotubes comprise single sheets of carbon atoms similar to those in graphite; however, in nanotubes, the sheets are rolled onto a cylinder. Electrons can travel along the single sheet in a nanotube more easily than in the multiple sheets in graphite, hence they are better conductors of electricity. The layers in graphite mean that they can slide easily over each other but carbon nanotubes are one of the strongest substances known.

4. Buckyballs are cage-like molecules of carbon, with the most common being a C_{60} molecule. Nanotubes comprise sheets of carbon atoms similar to those in graphite that are rolled onto a cylinder. They may form into single or multiple layers. Graphene is a flat layer of carbon hexagonal sheets that is only one atom thick.

7.4 A review of bonding

Practice problem 3

a. Substance *C* is metallic.
b. Substance *D* is covalent layer.

Practice problem 4

a. Copper: metallic bonding
b. Copper(II) oxide: ionic bonding
c. Oxygen gas: covalent bonding within its molecules and dispersion forces between them

7.4 Exercise

1. *E*: covalent molecular bonding.
F: covalent lattice.

2. a. Covalent molecular, due to its poor conductivity in its solid and liquid state.
b. Metallic, due to its conductivity in solid and liquid form.
c. Covalent molecular, due to its poor conductivity of electricity and heat.
d. Ionic, due to its conductivity of electricity in its liquid, but not its solid state.

3. a. Substance *O* **b.** Substance *M*
c. Substance *L* **d.** Substance *N*

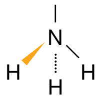
4. a. Metallic bonding **b.** Ionic bonding
c. Dispersion forces between atoms **d.** Covalent bonding
e. Covalent bonding

5. a. Metallic bonding **b.** Covalent network lattice
c. Covalent layer lattice **d.** Hydrogen bonding
e. Ionic bonding

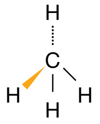
6. a. Covalent molecular bonding
b. Covalent network
c. Ionic network lattice
d. Metallic lattice
e. Metallic lattice
f. Covalent molecular bonding

7. i. a. Carbon dioxide is a covalent molecular substance with weak dispersion forces between its molecules.
b. O=C=O

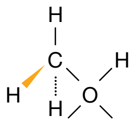
ii. a. Ammonia is a covalent molecular substance with hydrogen bonding and dispersion forces between its molecules.
b.



iii. a. Methane is a covalent molecular substance with weak dispersion forces between its molecules.
b.



iv. a. Methanol is a covalent molecular substance with hydrogen bonding and dispersion forces between its molecules.
b.



8. D. H—F has the strongest dipole because the electronegativity difference between hydrogen and fluorine is greater than any of the other electronegativity differences (it is a hydrogen bond).

9. a. Sodium chloride has ionic bonding, which is characterised by strong electrostatic attractions between oppositely charged ions extend throughout the lattice.
b. Salt crystals are brittle because of their structure. Ionic substances are made of lattices of cations and anions. When pressure forces like charges to align, repulsion occurs and the crystal shatters.
c. Salt is useful as a flavouring because it is soluble in water. Water molecules can move between the ions forming ion–dipole bonds with the ions, hence freeing them by disrupting the rigid ionic lattice.

10. a. Gold is used in jewellery because it is shiny, durable, highly ductile and malleable. These properties are due to its metallic lattice. The non-directional nature of the metallic bond allows distortion without disrupting the lattice, while the lustre is due to light being reflected from the delocalised, outershell electrons.
b. Graphite has a covalent layer lattice with weak dispersion forces between the layers. This accounts for its use as a lubricant; it is slippery because the layers slide readily over one another.

7.5 Review

7.5 Exercises

7.5 Exercise 1: Multiple choice questions

1. D.
2. D.
3. D.
4. D.
5. C.
6. A.
7. D.
8. B.
9. C.
10. A.
11. C.
12. i. B. **ii.** C.
13. D.
14. D.
15. C.

7.5 Exercise 2: Short answer questions

1. a. 1000 (10^3) **b.** 1 000 000 (10^6)
c. 1 000 000 000 (10^9)

2. a. Diamond has a covalent network lattice. Each carbon is bonded to four other carbon atoms in a tetrahedral arrangement. Graphite has a covalent layer lattice. Each carbon is bonded to three other carbons in a trigonal planar arrangement (forming layers). There are delocalised electrons between the layers.
b. Both have strong covalent bonding between the carbon atoms in the lattice; however, graphite has strong covalent bonding in only two dimensions, with weak dispersion forces between layers, while diamond has strong covalent bonding in all three dimensions.
c. Both consist of atoms held together by covalent bonds.
d. Graphite is used as a lubricant, in pencil 'lead' and electrodes; diamond is used in jewellery and drilling and as an abrasive (for cutting glass).

3. a. Graphite is acting as a lubricant.

b. Weak dispersion forces exist between the layers.

4. Diamonds can cut glass because they are very hard due to strong covalent bonding in three dimensions.

5. Diamonds are brittle because they are held together in a strong covalent lattice. When enough pressure is applied to break the strong covalent bonds, the lattice is distorted and the diamond shatters.

6. A C_{60} molecule (buckyball) is made of alternating hexagonal and pentagonal rings of covalently bonded carbon atoms that form a ball similar to the shape of a soccer ball.

7. a. Ionic **b.** XY_2

8. a. C. **b.** E. **c.** A. **d.** D. **e.** B.

9. a. $CaCl_2$ is an ionic substance that will contain ionic bonds due to the calcium donating its two spare electrons, and each chlorine atom accepting one electron to fill its outer shell.

b. In CCl_4, the carbon will form four covalent bonds with chlorine to create a covalent molecule.

c. MgO contains ionic bonding because the magnesium atom donates its two outershell electrons and the oxygen atom accepts two electrons to fill its outer shell.

d. Water contains covalent bonding within molecules because the oxygen and hydrogen atoms share electrons, and dispersion forces and hydrogen bonding between molecules.

e. The bonding within $CO_2(g)$ is covalent and the intermolecular bonding is dispersion forces.

f. In helium, only weak dispersion forces exist between atoms.

g. C_2H_2 contains covalent bonds within molecules and dispersion forces between molecules.

h. HNO_3 contains covalent bonds within molecules and dispersion forces between molecules.

i. H_2S contains covalent bonding within molecules.

j. $CaCO_3$ contains ionic bonding within molecules and dispersion forces between molecules.

k. Ionic bonding occurs in KOH because it contains K^+ ions and OH^- ions.

l. In 18 carat gold, metallic bonding occurs, in which the valence electrons delocalise and form a sea of electrons around the metal ions.

10. Discrete molecules are often volatile (evaporate) at room temperature due to weak intermolecular forces. Metallic, ionic and covalent network substances have strong bonds and remain solid at room temperature; hence, they do not have a smell.

11. a. i.

Compound	Electrical conductivity	Bonding
Benzoic acid	Poor	Covalent molecular
Magnesium sulfate	Good	Ionic
Sodium fluoride	Good	Ionic
Paraffin wax	Poor	Covalent molecular
Aluminium	Good	Metallic
Sugar	Poor	Covalent molecular

ii. Aluminium would be expected to conduct electricity in the solid state as it is a metallic element. Metallic bonding consists of positive cations surrounded by a sea of delocalised electrons which are always able to carry charge.

iii. See table at foot of the page.*

iv. See table at foot of the page.*

11 iii*

Compound	Electrical conductivity	Bonding	Bond/force breaking at melting
Benzoic acid	Poor	Covalent molecular	Weak dispersion forces between molecules
Magnesium sulfate	Good	Ionic	Strong electrostatic forces of attraction between oppositely charged ions
Sodium fluoride	Good	Ionic	Strong electrostatic forces of attraction between oppositely charged ions
Paraffin wax	Poor	Covalent molecular	Weak dispersion forces between molecules
Aluminium	Good	Metallic	Strong electrostatic forces of attraction between the metal ions and the delocalised sea of outershell electrons
Sugar	Poor	Covalent molecular	Weak dispersion forces between molecules and hydrogen bonds

11 iv*

Compound	Electrical conductivity	Bonding	Expected melting point
Benzoic acid	Poor	Covalent molecular	Low
Magnesium sulfate	Good	Ionic	High
Sodium fluoride	Good	Ionic	High
Paraffin wax	Poor	Covalent molecular	Low
Aluminium	Good	Metallic	High
Sugar	Poor	Covalent molecular	Low

- **b.** Covalent molecular substances do not conduct electricity in the molten state because there are no mobile (free moving) electrons or ions.
- **c.** Some molecular substances react with water to form ions (ionise) while other do not. For example, non-polar molecules do not react with water.

12. Diamond has a covalent network structure with strong covalent bonds in three dimensions. Tungsten has a metallic lattice structure consisting of close-packed cations in a 'sea' of electrons. The bonding is strong in both substances, and extreme heat is needed to break the bonds.

13. a. *G*
- **b.** *H*
- **c.** *E* and *F*

14. *A* Covalent molecular
B Metallic
C Ionic
D Covalent network

15. a. Lightweight and strong; also resists corrosion.
- **b.** Lightweight (saves on transport costs); can be recycled.
- **c.** Very good conductor of electricity; also resists corrosion.
- **d.** Good conductor of electricity; chemically inert.

7.5 Exercise 3: Exam practice questions

1. Silicon dioxide has a covalent network structure. Its strong three-dimensional bonding results in a rigid solid at room temperature. Carbon dioxide is a discrete covalent molecular compound. It has weak dispersion forces between its molecules and hence is a gas at room temperature.

2. The properties of graphite and graphene depend on their respective structures. Because their structures vary considerably, so do their properties. In graphite, each carbon atom forms covalent bonds with three other carbon atoms to form layers of hexagonal sheets. These layers are stacked on top of each other and are held together by dispersion forces. By comparison, graphene is flat layer of carbon hexagonal sheets that is only one atom thick. There are no weak dispersion forces holding layers together, meaning that all bonding within the network is strong.

3. a. The type of bonding in each substance is covalent network bonding.
- **b.** Germanium would be expected to be hard. This is because the strong covalent bonds between its atoms are hard to disrupt.
- **c.** C is diamond because the dark grey colour is consistent with it turning into graphite. Of the others, the melting temperature represents the increasing strength of their covalent bonds. This is caused by the bonds becoming shorter, due to the period on the periodic table to which each belongs. Therefore, D is Ge (longest and weakest bonds), B is silicon and A is silicon carbide.

4. a. HF
b. Na
c. KF
d. He
e. Si
f. H_2S

8 Organic compounds

8.2 Grouping hydrocarbons into families

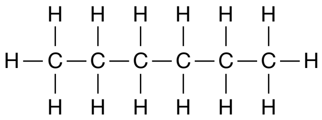
Practice problem 1

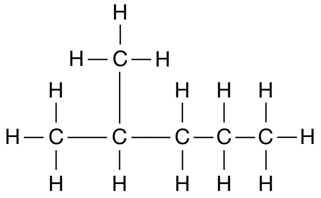
$C_{16}H_{34}$

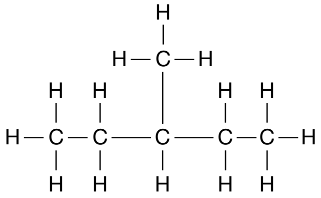
Practice problem 2

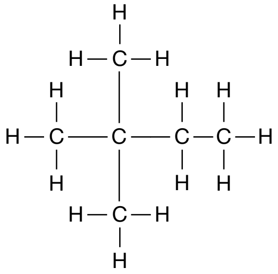
C_6H_{12}

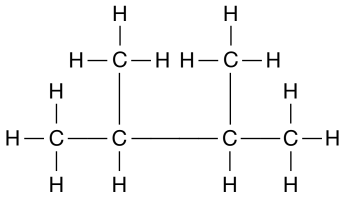
Practice problem 3







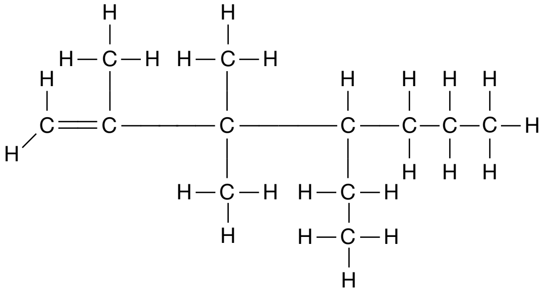




Practice problem 4

2,2,3-trimethylbutane

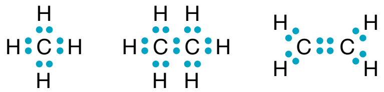
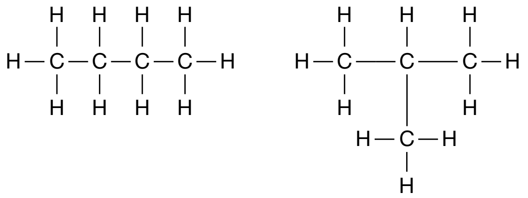
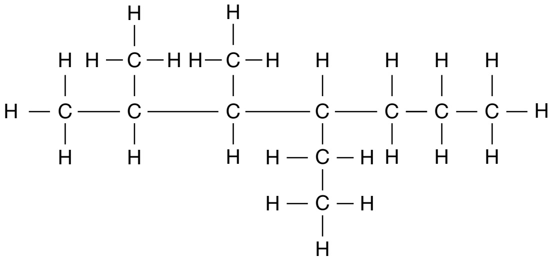
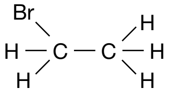
Practice problem 5

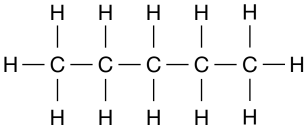
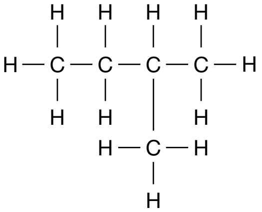
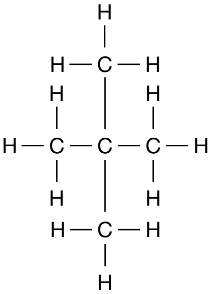
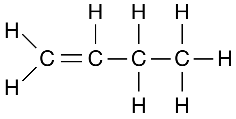
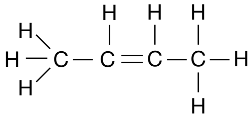
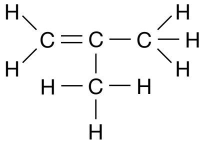
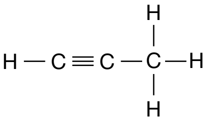


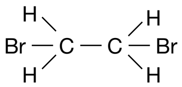
Practice problem 6

5-methylhept-2-ene

8.2 Exercise

1. $C_{10}H_{22}$
2. $C_{13}H_{26}$
3. 
methane ethane ethene
4. a. Carbon dioxide and water
   b. Carbon monoxide or carbon, and water
5. a. Combustion and substitution reactions
   b. Combustion and addition reactions
6. As the intermolecular forces between molecules increase, the melting and boiling points increase. The intermolecular forces between alkanes are dispersion forces. Dispersion forces increase with molecular size (with an increase in the number of electrons in a molecule). So as the molecular size of an alkane increases, the forces between molecules increases, meaning that the alkane will be more solid than a smaller alkane, which has weaker dispersion forces.
7. Butane has a higher boiling point than propane because it is a larger molecule (contains more electrons) and, hence, stronger dispersion forces will exist between butane molecules.
8. Although both molecules have dispersion forces of about the same magnitude, molecules of hexane can pack closer together than molecules of 2,3-dimethylbutane. They, therefore, have a stronger effect in the former due to the molecules being closer. This means hexane has the higher boiling point.
9. 
10. See figure at foot of the page.*
11. a. The *only* position for the methyl group is carbon 2. Hence numbers are not required.
    b. Propan-3-ol is incorrect because it has been numbered from the wrong end. It should be propan-1-ol.
12. 3,3 dimethylhexane
13. 
14. a. 
    Note: other answers showing multiple Br substitution are possible.

10*	(i) pentane	(ii) butene	(iii) propyne
(a)	3 isomers	3 isomers	1 isomer
(b)	  	  	

- **b.** 
- **c.** Part **a**. substitution reaction. Part **b**. addition reaction.

15. a. See figure at foot of the page.*
- **b. i.** 2,2,4,4-tetramethylheptane
  - **ii.** Octane
- **c. a. i.** $CH_2CHCHCH_2CH_2CH_3$
  - **ii.** $CH_2C(CH_3)CH_3$
  - **iii.** $CH_3CCCH_2CH_3$
  - **b. i.** $CH_3C(CH_3)_2CH_2C(CH_3)_2CH_2CH_2CH_3$
  - **ii.** $CH_3(CH_2)_6CH_3$

8.3 Functional groups

Practice problem 7

- **a.** Octan-1-ol
- **b.** Propan-1-ol

Practice problem 8

- **a.** 2-methylpentan-3-ol
- **b.** $CH_3CH(CH_3)CH(OH)CH_2CH_3$

Practice problem 9

- **a.** The longest carbon chain is five carbons long, with the carboxyl group attached to the first carbon. A methyl group is also attached to the third carbon, so the name of the molecule is 3-methylpentanoic acid.
- **b.** $CH_3CH_2CH(CH_3)CH_2COOH$

Practice problem 10

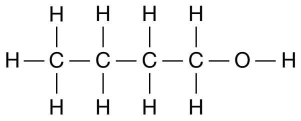
- **a.** Methyl propanoate: $C_4H_8O_2$
- **b.** Propyl methanoate: $C_4H_8O_2$
- **c.** Ethyl ethanoate: $C_4H_8O_2$

All of the esters listed are isomers.

8.3 Exercise

1. $CH_3CH_2CH_2CH_2CH_2CH_2CH_2OH$ or $CH_3(CH_2)_6OH$
2. $CH_3CH_2CH_2CH_2CH_2COOH$ or $CH_3(CH_2)_4COOH$
3. A functional group is an atom, group of atoms or a bond that, when inserted into a molecule, gives that molecule a characteristic set of properties. Two examples of functional groups are the hydroxyl group (−OH) and the carboxyl group (−COOH).
4. **a.** Methylpropan-2-ol
   - **b.** $CH_3C(CH_3)(OH)CH_3$
   - **c.** The CH_3 group does not need a number because it can only go on carbon 2. (If it is placed on carbon 1 or carbon 3, it makes a butanol.) However, the −OH group can go on either carbon 1 or carbon 2, hence the number must be specified.
5. **a.** 3,4-dimethylpentanoic acid
   - **b.** $CH_3CH(CH_3)CH(CH_3)CH_2COOH$
6. **a.** A homologous series is a series of similar compounds, the formulas of which differ by CH_2 for successive members.
   - **b.** The alcohols are a homologous series because the formulas of successive members differ by CH_2. For example, the first three members are methanol, ethanol and propanol. The respective formulas of these are CH_3OH, CH_3CH_2OH and $CH_3CH_2CH_2OH$.
   - **c.** The carboxylic acids are a homologous series because the formulas of successive members differ by CH_2. For example, the first three members are methanoic acid, ethanoic acid and propanoic acid. The respective formulas of these are HCOOH, CH_3COOH, CH_3CH_2COOH.
   - **d.** The alkanes are a homologous series because the formulas of successive members differ by CH_2. For example, the first three members are methane, ethane and propane. The respective formulas of these are CH_4, C_2H_6 and C_3H_8.
7. **a.** Molecular formula: $C_4H_{10}O$

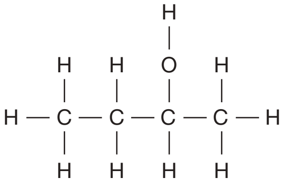
   Structural formula:

   $$\begin{array}{ccccccccccccc} & & \mathrm{H} & & \mathrm{H} & & \mathrm{H} & & \mathrm{H} & & & & \\ & & | & & | & & | & & | & & & & \\ \mathrm{H} & - & \mathrm{C} & - & \mathrm{C} & - & \mathrm{C} & - & \mathrm{C} & - & \mathrm{O} & - & \mathrm{H} \\ & & | & & | & & | & & | & & & & \\ & & \mathrm{H} & & \mathrm{H} & & \mathrm{H} & & \mathrm{H} & & & & \end{array}$$

   Condensed formula: $CH_3CH_2CH_2CH_2OH$
   - **b.** Molecular formula: $C_4H_{10}O$

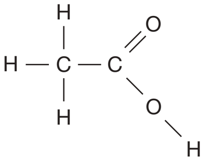
     Structural formula:

     $$\begin{array}{ccccccccccc} & & & & & & \mathrm{H} & & & & \\ & & & & & & | & & & & \\ & & \mathrm{H} & & \mathrm{H} & & \mathrm{O} & & \mathrm{H} & & \\ & & | & & | & & | & & | & & \\ \mathrm{H} & - & \mathrm{C} & - & \mathrm{C} & - & \mathrm{C} & - & \mathrm{C} & - & \mathrm{H} \\ & & | & & | & & | & & | & & \\ & & \mathrm{H} & & \mathrm{H} & & \mathrm{H} & & \mathrm{H} & & \end{array}$$

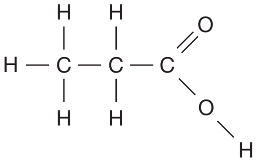
     Condensed formula: $CH_3CH_2CH(OH)CH_3$
8. **a.** Molecular formula: $C_2H_4O_2$

   Structural formula:

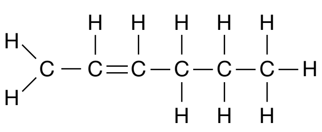
   Condensed formula: CH_3COOH
   - **b.** Molecular formula: $C_3H_6O_2$

     Structural formula:

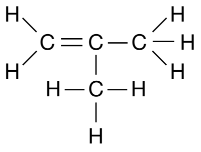
     

     Condensed formula: CH_3CH_2COOH

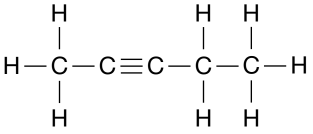
15a*



(i)



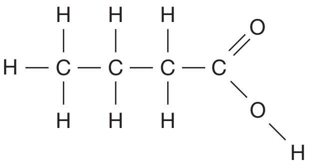
(ii)



$$\begin{array}{ccccccccccccc} & & \mathrm{H} & & & & & & \mathrm{H} & & \mathrm{H} & & \\ & & | & & & & & & | & & | & & \\ \mathrm{H} & - & \mathrm{C} & - & \mathrm{C} & \equiv & \mathrm{C} & - & \mathrm{C} & - & \mathrm{C} & - & \mathrm{H} \\ & & | & & & & & & | & & | & & \\ & & \mathrm{H} & & & & & & \mathrm{H} & & \mathrm{H} & & \end{array}$$

(iii)

c. Molecular formula: $C_4H_8O_2$
Structural formula:

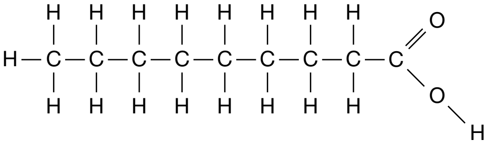


Condensed: $CH_3CH_2CH_2COOH$

9. a. **i.** Methyl butanoate
ii. Ethyl propanoate
iii. Methyl methanoate

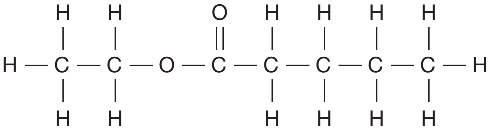
b. For those listed, a. and b. are isomers that share the molecular formula $C_5H_{10}O_2$.

10.



$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2COOH$ or $CH_3(CH_2)_7COOH$

11.



$CH_3CH_2CH_2CH_2COOCH_2CH_3$
Note: drawing/writing this in opposite direction (from right to left) is equally correct.

12. Pentanoic acid exhibits hydrogen bonding between its molecules while ethyl propanoate does not. This results in stronger intermolecular bonding and, hence, a higher boiling point because the molecules are harder to separate.

13. a. Propan-1-ol and ethanoic acid
b. Ethanol and propanoic acid
c. Pentan-1-ol and decanoic acid

14. Methanol and ethanoic acid

15. a. $CH_2(OH)CH(OH)CH_2(OH)$
b. $CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2COOH$ or $CH_3(CH_2)_{16}COOH$

8.4 The origin and use of crude oil

Practice problem 11

The molecules making up bottled gas are smaller than those making up petrol. Because dispersion forces are the only intermolecular attraction between all of these molecules, a weaker force of attraction will exist between the molecules in bottled gas compared to those in petrol. Bottled gas will, therefore, have the lower boiling point (a lower condensation point) and so will be collected higher in the column.

8.4 Exercise

1. Crude oil is formed from the remains of marine animal and plant life that existed millions of years ago. When these died, they fell to the bottom of the sea and became covered by sediment. Heat and pressure over the subsequent millions of years changed their remains into crude oil (and natural gas).

2. The molecules making up petrol are smaller than those making up kerosene. Because dispersion forces are the only intermolecular attraction between all of these molecules, a weaker force of attraction will exist between the molecules in petrol compared to those in kerosene. Petrol will, therefore, have the lower boiling point (a lower condensation point) and so will be collected higher in the column.

3. Lighter fractions are either gases or liquids with low boiling points that are easy to vaporise. It is the vapours that burn.

4. The fractions recovered lower contain larger hydrocarbon molecules. Viscosity increases with increasing molecular size because of the stronger intermolecular dispersion forces and the increased possibility of entanglement of the long molecules.

5. Sample answer: $C_{12}H_{26}(g) \rightarrow C_{10}H_{22}(g) + C_2H_4(g)$
A correct answer must show a smaller alkane and a small alkene. So a number of different answers will be correct.

8.5 Determining the empirical and molecular formulas of organic compounds

Practice problem 12

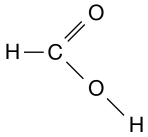
Empirical formula: C_2H_5O

Practice problem 13

Molecular formula: C_6H_{14}

Practice problem 14

a. Empirical formula: CH_2O_2
b. Molecular formula: CH_2O_2
c.



Condensed formula: HCOOH

8.5 Exercise

1. a, b, c, e, g and h are empirical formulas.

2. If only one hydroxy (−OH) group is in the alcohol molecule, only one oxygen will be in its molecular formula. This means that the formula must also be an empirical formula because more than one oxygen per molecule is not possible.

3. Molecular formula: $C_2H_6O_2$

4. a. Molecular formula: C_5H_{12}
b. Molecular formula: C_6H_{10}
c. Molecular formula: C_6H_8
d. Molecular formula: C_6H_{12}
e. Molecular formula: $C_{10}H_{20}$
f. Molecular formula: $C_6H_{12}O_6$
g. Molecular formula: $C_2H_4Cl_2$

5. Molecular formula: $C_{10}H_{14}N_2$

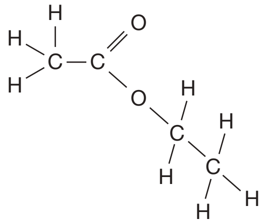
6. Molecular formula: $C_8N_4O_2H_{10}$

7. Molecular formula: $C_5H_{10}O_2$

8. Molecular formula: $C_3H_6O_3$.

9. Molecular formula: $C_2H_4Br_2$.

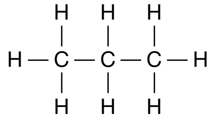
10. Molecular formula: $C_4H_8O_2$



Note: this is the actual structure referred to in the question. However, on the basis of the information given, any other ester with the formula $C_4H_8O_2$ would also be correct.

11. a. Molecular formula: C_3H_8

b.



$$\begin{array}{ccccccccc} & & H & & H & & H & & \\ & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & \\ & & H & & H & & H & & \end{array}$$

c. This molecule belongs to the alkane homologous series.

12. Molecular formula: C_5H_{10}

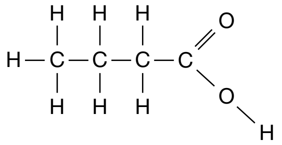
13. a. Compounds that contain only carbon, hydrogen and oxygen are the alcohols, carboxylic acids and esters.

b. If a metal or carbonate is added to an acid, gas of hydrogen or carbon dioxide respectively would form. This test would differentiate a carboxylic acid from an alcohol or an ester.

c. Empirical formula: CH_4O

d. Molecular formula: CH_4O

14.



The compound is butanoic acid.

8.6 Review

8.6 Exercises

8.6 Exercise 1: Multiple choice questions

1. C.
2. B.
3. C.
4. B.
5. D.
6. C.
7. C.
8. C.
9. D.
10. D.
11. D.
12. C.
13. D.
14. C.
15. A.

8.6 Exercise 2: Short answer questions

1. a. Hydrocarbons are chemical compounds that contain only carbon and hydrogen.

b. Propane, C_3H_8, propene, C_3H_6, and propyne, C_3H_4

2. Alkenes and alkynes need at least two carbons to have a double or triple bond between them. This is not possible with just one carbon.

3. a. Alkanes are produced by the fractional distillation of petroleum.

b. The uses of alkanes include bottled gas, plastics, chemicals, fuel, lubricants, waxes and road surfaces

4. Isomers are molecules with the same molecular formula but different structures. For example, C_5H_{12} has three isomers — pentane, 2-methylbutane and 2,2-dimethylpropane — and C_4H_8 has three isomers — but-1-ene, but-2-ene and 2-methylpropene.

5. a. Saturated hydrocarbons contain only single carbon–carbon bonds. Unsaturated hydrocarbons contain one or more carbon–carbon double (or triple) bonds.

b. Unsaturated hydrocarbons

6. a. $C_{22}H_{46}$ **b.** $C_{17}H_{34}$ **c.** $C_{13}H_{24}$

7. Successive members of the alcohol series differ by a CH_2 group.

8. a. Substitution reactions are where one or more hydrogen atoms in an alkane are replaced by another atom or atoms; e.g. $C_2H_6 + Cl_2 \rightarrow C_2H_5Cl + HCl$.

b. Addition reactions are where a small molecule adds across the double bond of an alkene; e.g. $C_2H_4 + HCl \rightarrow C_2H_5Cl$.

c. Functional group is an atom or group of atoms that gives a molecule a characteristic set of properties; e.g. in CH_3OH, the functional group is the hydroxy group $(-OH)$.

d. Fractional distillation is the separation of a mixture of liquids on the basis of their varying boiling temperatures; e.g. petroleum is fractionally distilled to produce gas, petrol, kerosene, diesel oil and lubrication oil.

e. Catalytic cracking is the heating of larger, less useful hydrocarbon molecules, in the presence of a catalyst, to produce smaller, more useful hydrocarbon molecules; e.g. $C_{12}H_{26} \rightarrow C_{10}H_{22} + C_2H_4$.

9. a. Methyl butanoate **b.** Propyl propanoate

c. Ethyl methanoate

10. a. $CH_3CH_2CH_2COOCH_2CH_3$

b. $CH_3COOCH_2CH_2CH_3$

c. $CH_3CH_2COOCH_2CH_2CH_2CH_3$

11. Multiple answers are possible for each part. One example is provided for each part.

a. $CH_3CHCHCH_3$, but-2-ene

b. $CH_3CH_2CH_2CH_2OH$, butan-1-ol

c. $CH_3CH_2CH_2COOH$, butanoic acid

d. $CH_3COOCH_2CH_3$, ethyl ethanoate

12. a. $CH_2C(CH_3)CH_2CH(CH_3)CH_2CH_2CH_3$

b. $CHCCH_2CH_2CH_3$

c. $CH_3CH_2CHCHCH_2CH_2CH_2CH_3$

d. $CH_3CH(CH_3)CH(CH_3)CHCH_2$

e. $CH_3CH(CH_3)CH(CH_3)CH_3$

f. $CH_3CH_2CH(CH_3)CH(CH_2CH_3)CH_2CH_2CH_3$

13. a. Methylbutane

b. 3-methylpentane

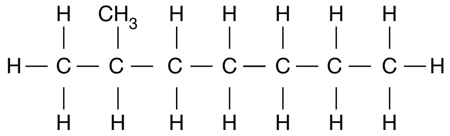
c. 3-ethylhexane

d. 3,3-dimethylhexane

e. 2,4-dimethylhexane

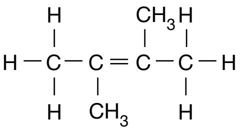
f. 2,2,4,4-tetramethylpentane

14. a.

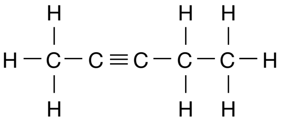


$$\begin{array}{ccccccccccccccccc} & & H & & CH_3 & & H & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & C & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & | & & | & & | & & | & & \\ & & H & & H & & H & & H & & H & & H & & H & & \end{array}$$

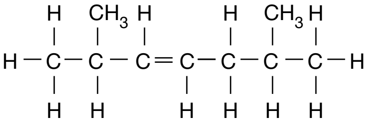
b.



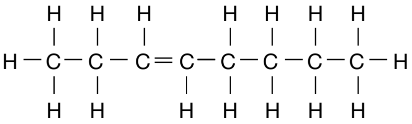
c.



d.



e.



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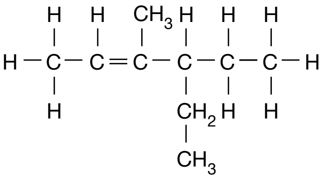
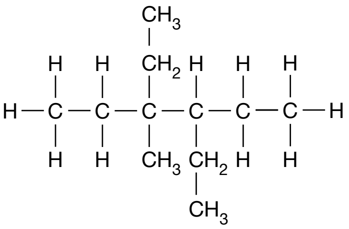
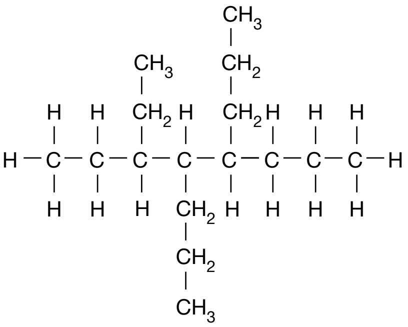


g.

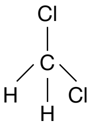


h.

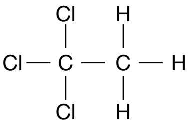


15. The longest chain has seven carbon atoms. The compound should have been named 3,3-dimethylheptane.

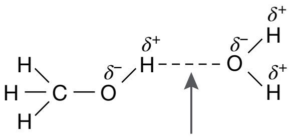
16. a.



b.



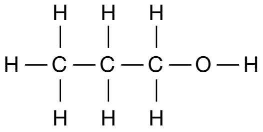
17.



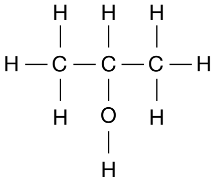
hydrogen bond between methanol and water

18. Molecular formula: CH_3H_8O

The alcohol is either propan-1-ol



or propan-2-ol

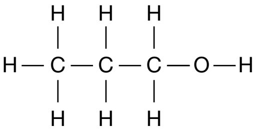


19. a. Empirical formula: C_3H_8O

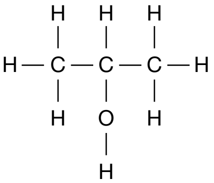
b. Molecular formula: CH_3H_8O

c. Alcohols, because they react with a carboxylic acid (ethanoic acid) to produce a pleasant smelling product — presumably an ester. Both compounds contain only 1 oxygen atom. As the hydroxyl group (−OH) uses only one oxygen, they are alcohols.

d. Propan-1-ol



Propan-2-ol



e. Propan-1-ol would have a higher boiling point due to its molecules being able to get closer together. This would increase the effect of the dispersion forces between its molecules, making the intermolecular bonding stronger in propan-1-ol than in propan-2-ol.

8.6 Exercise 3: Exam practice questions

1. a. Boiling points reflect the strength of the intermolecular forces operating. The intermolecular forces operating are dispersion forces, which increase with molecule size.

b. The easier it is to vaporise, the more flammable a hydrocarbon is. Petrol has smaller molecules than paraffin wax and, hence, has weaker dispersion forces between its molecules, making it easier to vaporise.

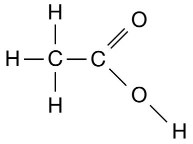
c. These two substances would have roughly equal dispersion forces. However, hydrogen bonding exists between molecules of ethanol, but not between molecules of propane. The stronger intermolecular forces in ethanol mean that it is a liquid at room temperature, whereas propane with the weaker intermolecular forces is a gas.

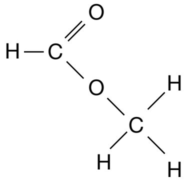
2. a. Empirical formula: CH_2O

b. Molecular formula: $C_2H_4O_2$

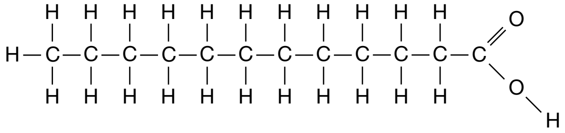
$$\text{The molecular formula} = 2 \times \text{empirical formula}$$
$$= C_2H_4O_2.$$

c. Either of the following structural formulas are correct for this compound.





3. **a.** Carboxylic acids homologous series
   b. 6-methyl-5-propyloctanoic acid
   c. An example is shown. *Note:* many other isomers are valid answers.



4. **a.** Petrol
   b. Petrol
   c. Paraffin
   d. If the fractions collected were still mixtures, further refining in an additional fractional distillation process would separate these mixtures.
5. **a.** Carboxylic acid
   b. Alcohols and esters
   c. **i.** 5.420 g **ii.** 0.903 g **iii.** 2.894 g
   d. Empirical formula: $C_5H_{10}O_2$
   e. Molecular formula: $C_5H_{10}O_2$

9 Polymers

9.2 Formation of polymers

Practice problem 1

Approximate molar mass = 70 000 g mol^{-1}

9.2 Exercise

1. Monomers for addition polymerisation require a carbon–carbon double bond. Alkanes do not possess a double carbon–carbon bond but alkenes do.
2. **a.** Examples of natural polymers include proteins, starch, cellulose, wool, cotton and many others.
   b. Examples of synthetic polymers include polyethene, polypropene, polyvinyl chloride, nylon, PTFE, PVA, PET, polystyrene and many others.
3. Approximate molar mass = 70 000 g mol^{-1}
4. Like most addition monomers, special conditions are required before polymerisation can occur. These often involve the use of a catalyst and elevated temperatures and pressures. If these conditions are not present, polymerisation will not take place.
5. **a.** ≈ 16 000
   b. 32 000
6. Alkynes can act as addition monomers because they have a triple carbon−carbon bond. The polymer produced uses electrons from one of the triple bonds to link the monomers together. The resulting polymer would contain residual carbon–carbon double bonds.

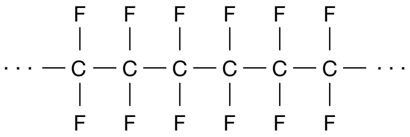
9.3 Linear and cross-linked polymers

9.3 Exercise

1. The bonding between chains in thermosetting polymers is covalent bonding. This is much stronger than the weak intermolecular forces that exist between the chains in thermoplastic polymers.
2. **a.** Thermoplastic polymers only have weak forces, most often dispersion forces, between their chains.
   Elastomers have a small amount of covalent cross-linking between their chains.
   Thermosetting polymers have a large number of covalent cross-links between their chains.
   b. Thermoplastic polymers are able to be softened and melted and then re-moulded.
   Elastomers are characterised by their ability to be stretched and to then return to their original shape once the stretching force is removed.
   Thermosetting polymers are not able to be softened by heating and many will instead char.
3. **a.** Making an article from a thermosetting polymer involves two main steps. The first is to place the preliminary linear polymers (called the resin) into the mould. These polymers have functional groups attached to the long polymer backbone. The second step is to induce cross-linking by making the functional groups from one chain react with functional groups from a neighbouring chain. This may be done by heat and pressure, or by adding a chemical to cause this reaction.
   b. A plastic is defined as a substance that can be moulded at some stage in its production. Thermosetting polymers are moulded at the stage before the cross-linking is initiated. Hence, by definition, they can be called plastics.
4. A thermoplastic article is able to be softened and melted. After processing (such as moulding and extruding), it will retain its new shape upon cooling. Thermosetting polymers cannot do this because they cannot be melted and re-shaped.
5. It would be expected that polymer *A* would be harder, less flexible and more heat-resistant.

9.4 Polymer selection

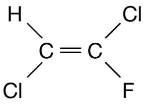
Practice problem 2



OR

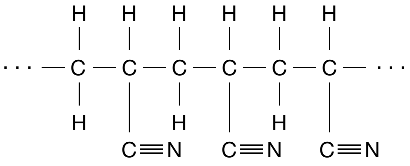
$-\!\left[CF_2 \right]\!-_n$

Practice problem 3

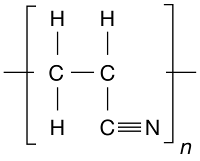


9.4 Exercise

1\.



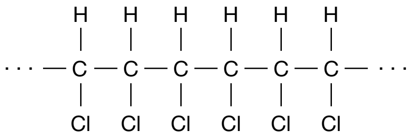
OR



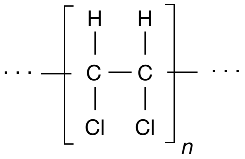
Note: a structure similar to the above but where each successive monomer is randomly inverted or flipped would also be considered a correct answer for this question.

2\. a. Monomer II is the correct monomer.

b\.



OR



3\. Chain length: In a linear polymer, longer chains will be harder, less flexible and more heat resistant.
Degree of branching: As the amount of branching increases, polymers will be more flexible, softer and not as dense. Detailed responses can be found in the worked solutions in your digital formats.

4\. Plasticisers are small molecules that fit between the chains of a linear polymer and push them further apart. This makes the dispersion forces acting between these chains less effective. As a result, the polymer will be softer and more flexible.

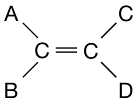
5\. Ethene is water repellent and polyvinyl alcohol is water soluble due to the monomers from which each is made. Ethene is non-polar due to its C–H bonds and its symmetry. Ethene, and consequently polyethene, are thus insoluble in water and water repellent. Vinyl alcohol, and subsequently, polyvinyl alcohol, have O–H groups, which make them polar and soluble in water.

6\. Crystallinity is caused by polymer chains being able to pack together in an orderly pattern. This is more likely to occur when the side groups attached to these chains are in an orderly or predictable pattern. If the pattern of these side groups is more random, it will make it more difficult for the chains to pack together.

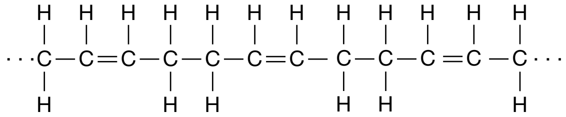
7\. Polypropene, polystyrene and polytetrafluoroethene are all made from non-polar monomers. As water is polar, the polymers will be water insoluble and water repellent.

8\. As the required polymer needs to soften under moderate heat, it cannot be thermosetting. Hence, the polymer cannot be cross-linked. It needs to be made from a monomer that is non-polar so that it will be water repellent. For it to be rigid, the movement of its long polymer chains needs to be restricted in the solid state. This is possible if the chains are long and able to tangle due to their length. Alternatively, if the chains can be encouraged to align more often with each other so that the degree of crystallinity is increased, this will also improve the polymer's hardness. However, if the use specifies continual cycles of melting and solidifying this last effect may not be reproducible.

9\.



10\.



11\. a. Sample answer:
Advantages: polymers are lighter than metal, making them more fuel efficient, and are easier to produce.
Disadvantage: polymers are not as strong as metal, and hence are more easily damaged.

12\. Detailed responses can be found in the worked solutions in your digital formats.

13\. PVC is not recycled to the same extent for two main reasons. Firstly, it is often used for a long time (e.g. pipes, gutters), and is not replaced often. This reduces the amount of recycling. Secondly, it contains chlorine and additives (plasticisers) which make the recycling process more complicated.

9.5 Review

9.5 Exercises

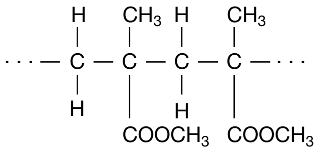
9.5 Exercise 1: Multiple choice questions

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

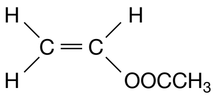
9.5 Exercise 2: Short answer questions

1. Polymers are very large molecules formed by joining together large numbers of small molecules (called monomers). Monomers are small molecules that can react together to form polymers. They must contain a carbon–carbon double bond.

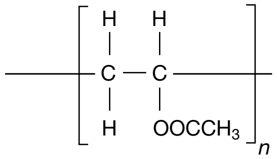
2. Advantages: can be moulded into any shape, don't corrode, are cheap to produce and lighter than metallic options, are good insulators of heat and electricity, some are recyclable, most are chemically inert, are strong and relatively cheap to produce. Detailed responses can be found in the worked solutions in your digital formats.
   Disadvantages: nonrenewable/finite resource (petroleum), take a long time to decompose (if at all) and contribute to landfill. Detailed responses can be found in the worked solutions in your digital formats.
3. Alkanes are represented by the formula, C_nH_{2n+2}. (a), (c) and (g) follow this relationship and therefore can not act as monomers in a polymerisation reaction. (b) and (e) are alkenes, as they follow the formula C_nH_{2n}. As these molecules contain double bonds, they can act as monomers. (d) and (f) can also act as monomers in polymerisation reactions as they are alkynes and follow the formula C_nH_{2n-2}.
4. Thermoplastic polymers, such as polyethene, polypropene and polystyrene, can be repeatedly melted, reshaped and hardened by cooling. Thermosetting polymers, such as Bakelite, Formica and rubber, do not melt when heated but instead char, and cannot be reshaped through melting.
5. a. Long polymer chains made by polymerising monomers.
   b. Has cross-links between its long polymer chains (but not as many).
6. The forces between its polymer chains are much stronger as the result of strong covalent bonds.
7. In LDPE, the extensive branching results in the polymer chains being kept further apart, resulting in less effective dispersion forces. In HDPE there is minimal branching so the polymer chains can pack together much more closely, resulting in more effective and hence stronger dispersion forces.
8. Recycling is important because it reduces both the use of valuable finite resources to produce plastics and the quantity of non-biodegradable material going into landfill.
9. Many recycling methods are dependent on plastic being sorted according to type. The recycling code both assists and reduces errors in this procedure.
10. a.



   b. Perspex is stronger and more impact-resistant than glass.
   c. Perspex is more expensive than glass and melts when exposed to high temperatures; it can also scratch and may not be as clear as glass.
11. a.



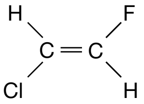
   b. Dispersion forces exist between these chains as well as dipole–dipole forces, due to the polar carbon–oxygen bonds.
   c.



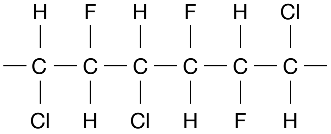
12. In low-density polyethene (LDPE), the polymer chains are highly branched, producing a low-density amorphous solid with very few crystalline regions that can scatter light. In high-density polyethene (HDPE), the polymer chains have minimal branching and a more ordered solid form that has many crystalline regions that can scatter light.
13. a. A flexible wrapper that dissolves in water is required. A monomer should be polar. Chains that are short and branched would assist in producing flexibility.
   b. A strong thermoplastic polymer with longer polymer chains where the entanglement and stronger dispersion forces between the chains add to the strength. Some degree of crystallinity might also be desirable, but not too much as you do not want the toy to shatter. Detailed responses can be found in the worked solutions in your digital formats.
   c. These are single-use bags that need to be very flexible, only moderately strong and moisture resistant. Non-polar monomers which polymerise to produce branched chains that cannot pack tightly together. Such a polymer should also be able to recycled.
   d. A garden hose needs to be flexible yet strong. A polymer with long chains should be chosen, and to overcome the stiffness it should also contain plasticisers.
14. See table at foot of the page.*

9.5 Exercise 3: Exam practice questions

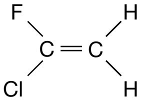
1. a.



   b.



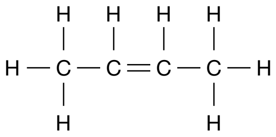
   c.



14*

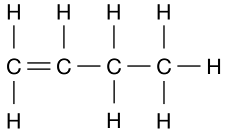
	Bond type	
Polymer	**Intramolecular**	**Intermolecular**
Diamond	Covalent	None exist
Graphite	Covalent	Dispersion forces between layers
Polyethene	Covalent	Dispersion forces between chains
Vulcanised rubber	Covalent	Disulfide covalent bonds between chains

2. a. i.



$H-C-C=C-C-H$

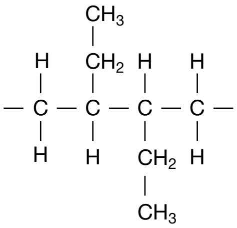
ii.



$C=C-C-C-H$

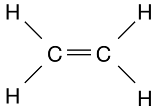
b. This polymer is made from the monomer but-2-ene.

c.



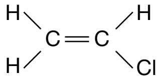
$-C-C-C-C-$

3. a. Monomer for polyethene (HDPE)



ethene

Monomer for polyvinyl chloride (PVC)



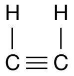
b. Both chains are about the same length and have roughly equal amounts of branching. They would, therefore, be expected to show the same degree of entanglement. Differences could be caused by the presence of chlorine atoms in place of some hydrogen atoms. This would increase the dispersion forces between the chains. Acting against this might be the fact that chlorine atoms are larger than hydrogen atoms, so the chains might not be able to pack as closely together. This would decrease the effect of the dispersion forces.

c. Plasticisers push the polymer chains a little further apart. This means the dispersion forces between the chains are not quite as effective and results in a polymer that is softer and more flexible.

d. The main difference between LDPE and HDPE is the amount of branching that exists in the polymer chains. This results in chains that cannot pack as effectively together in LDPE as they can in HDPE, resulting in LDPE being less dense, softer and more flexible than HDPE.

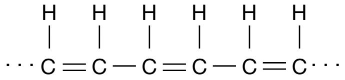
4. a. $H-C\equiv C-H$

Or



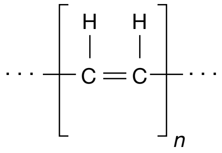
$C\equiv C$

b.



$\cdots C=C-C=C-C=C\cdots$

c.



$\cdots\left[C=C\right]_n\cdots$

d. Polyethyne would be more reactive than polyethene due to the presence of double bonds.

10 Research investigation

10.2 What is a research investigation?

10.2 Exercise

1. Responses should elaborate on how conducting research investigations allows us to gain an understanding of different aspects of science. Detailed responses can be found in the worked solutions in your digital formats.
2. A poster allows for trends and patterns to easily be seen, by making sure that information is succinct, coherent and relevant to the research question. Posters can include a variety of forms of media, including text, photographs, tables and graphs. Detailed responses can be found in the worked solutions in your digital formats.
3. Responses need to include the meanings of BY and NC. BY is referring to the need for attribution of the resource. NC refers to the fact that this image is only permitted for non-commercial uses and no financial gain is able to be made. You would be able to use this image in your report, because it is non-commercial; however, you would be required to attribute the work.
4. Responses will need to contain two examples of primary and secondary data. Detailed responses can be found in the worked solutions in your digital formats.

10.3 Effective science communication

10.3 Exercise

1. Responses will need to explain how communicating effectively in science allows trends and patterns to be clearly seen, provides support for research, deepens conceptual understanding, and provides context in regard to the importance and implication of findings. Detailed responses can be found in the worked solutions in your digital formats.
2. An international set of units (SI) for measurements allows for easy comparisons between different sets of data. This allows for consistency and provides a logical way to present data.
3. Theories are well-supported explanations of phenomena that have been based on investigations, research and observations. Models, on the other hand, are representations of observed phenomena. Detailed responses can be found in the worked solutions in your digital formats.
4. Responses will need to include one of each of the social, economic, environmental and ethical factors involved. Detailed responses can be found in the worked solutions in your digital formats.

10.4 The nature of evidence and information

10.4 Exercise

1. Responses will need to describe three aspects found in strong evidence as opposed to weak evidence. Detailed responses can be found in the worked solutions in your digital formats.
2. Responses will need to describe the difference between scientific and non-scientific ideas. Detailed responses can be found in the worked solutions in your digital formats.
3. Peer-reviewed work has higher validity and reliability compared to work that is not peer-reviewed. Responses will

need to explain this point. Detailed responses can be found in the worked solutions in your digital formats.

4. This is an example of selection bias. Responses will need to describe this type of bias. Detailed responses can be found in the worked solutions in your digital formats.

11 Properties of water

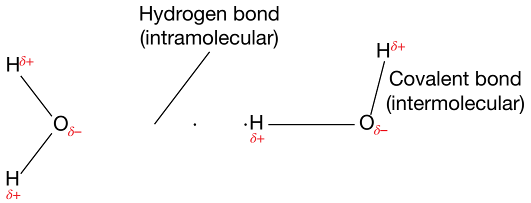
11.2 Intermolecular forces in water

Practice problem 1

$m(H_2O) = 150\,g$

11.2 Exercise

1. Water contains hydrogen–oxygen bonds, which allow for the formation of comparatively strong permanent dipoles called hydrogen bonds.
2. Weakest to strongest: dispersion forces, dipole–dipole forces, hydrogen bonding, covalent bonding.
3. a. A water molecule consists of two hydrogen atoms covalently bonded to a single oxygen atom. Two lone pairs are also present and the repulsion of the bonding pairs and the lone pairs results in the water molecule being bent.
   b. Oxygen is more highly electronegative than hydrogen. The shared electrons will, therefore, be more closely attracted to the oxygen, the oxygen atom becomes slightly negatively charged and the hydrogen slightly positively charged. This is a polar covalent bond. The molecule can also be described as polar because the bent shape means that there is an uneven charge distribution overall.
   c.



4. Gas particles have higher kinetic energy than liquid particles so energy needs to be removed, so changing state from gas to liquid (such as in condensation) is an exothermic process.
5. When liquid water evaporates into steam the intermolecular bonds, hydrogen bonds, are broken. The covalent bonds between the atoms in the molecule are still intact.
6. a. Mass and volume are required to calculate density.

   $$d = \frac{m}{V}$$

   b. Units of density: g mL^{-1} or kg m^{-3}
   c. Higher
7. $m(H_2O) = 65.3\,g$
8. When water freezes the density decreases because as the molecules lose energy they form a hexagonal crystalline lattice, which occupies more volume than the random arrangement of molecules in the liquid form.
9. Polonium's melting point can be approximated as –25 °C to –35 °C, the boiling point would be approximately 25 °C to 35 °C.
10. The hydrogen bonds between the H_2O molecules are stronger than the dipole–dipole forces between the H_2S molecules. (The electronegativity difference between S and H is much smaller than that between O and H so H_2S is considerably less polar.)

11.3 Latent heat and specific heat capacity

Practice problem 2

$q(H_2O) = 226\,kJ$

Practice problem 3

$q(H_2O) = 87.7\,kJ$

11.3 Exercise

1. Latent heat of vapourisation.
2. The heat energy is used to break the intermolecular hydrogen bonds between the water molecules while the temperature remains constant.
3. The energy (measured in joules) needed to raise the temperature of 1 g of water by 1 °C is 4.18 joules.
4. Water has a higher specific heat capacity than sand so it requires more heat energy to raise its temperature. This means the temperature of the sand increases at a faster rate than that of the water.
5. It would require more heat to raise the temperature of copper because it has higher specific heat capacity. (0.39 J g^{-1} $°C^{-1}$ for copper and 0.13 J g^{-1} $°C^{-1}$ for gold.)
6. $q(H_2O) = 37.9\,kJ$
7. Final temperature = 25.0 + 30.0 = 55.0 °C
8. $c(Cu) = 0.392$ J $g^{-1}°C^{-1}$
9. $q(Al) = 5.6\,kJ$
10. $q = 50.7\,kJ$

11.4 Review

11.4 Exercises

11.4 Exercise 1: Multiple choice questions

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

11.4 Exercise 2: Short answer questions

1. Usually substances become denser as the temperature decreases until a solid is formed. Water, however, has a lower density in the solid form because of the more ordered regular arrangement of molecules in the solid form compared to the random arrangement in liquid form. The ordered arrangement takes up more space so the density is less.
2. a. Latent heat refers to the energy that is required to change the state of a substance.
   b. The latent heat of vapourisation for water is much greater than the latent heat of fusion because when water is boiled the molecules have to break all of the hydrogen bonds and separate completely as they turn into a gas. The process of melting only partially separates the molecules and so requires less energy.

3. a. The energy needed to raise the temperature of 1g of a substance by 1 °C.
   b. Water has a relatively high specific heat capacity compared to other substances because it is capable of absorbing a relatively large amount of heat before its temperature rises significantly.
   c. Sample response: Water is used as a coolant in car radiators to keep the engine from overheating.
4. A large amount of energy is required to evaporate the water from clothing because of water's high latent heat of vapourisation value.
5. When a lettuce leaf is frozen the water in its cells expands and breaks the cell walls. This means the cell loses its shape, leaving the lettuce limp when defrosted.
6. Water has the highest melting and boiling temperatures because of the hydrogen bonding between the molecules. Other group 16 hydrides depend on dispersion forces, which are relatively weak but increase as the size of the molecule increases.
7. a. $q = 4.97 \times 10^2$ kJ
   b. A large amount of additional heat energy is released when the steam condenses into a liquid.
8. a. $q(H_2O) = 96.2$ kJ   b. 44.3 g
9. 20.7 °C
10. 11.5 kJ

11.4 Exercise 3: Exam practice questions

1. a. The density of water increases as the temperature decreases until maximum density at 4 °C because the water molecules start to form a hexagonal crystalline lattice in ice held together by hydrogen bonds instead of the closer random arrangement found in liquid water.
   b. Because ice is less dense than water, it floats on top of the water insulating the water and trapping oxygen below so organisms can survive below the ice. If ice were denser than water, it would freeze from the bottom up and sea creatures would not survive.
   c. Sample answer: Place 9.0 mL of water in a measuring cylinder. Place 9.0 mL of oil in a measuring cylinder. Place both in a freezer overnight. Observe the volume of each the next day. If the volume has increased this means that the density has decreased.
2. NASA scientists are looking for the presence of water because all organisms require water to survive and is essential for transportation of chemicals in organisms. Most biochemical reactions occur in aqueous solutions. So, if water is present, some life may also be present.
3. a. The damp cloth absorbs heat energy from the water bottle to evaporate/change state, and this keeps the bottle and water inside cool.
   b. The water vapour from the air condenses when it contacts the cold can. Energy goes into the can from the water in the condensation process.
   c. The fish will remain at 0 °C while ice is present. If the temperature of the fish were higher than 0 °C, heat energy would be transferred to the ice to melt the ice, therefore cooling the fish.
4. a. 197 g
   b. Specific heat capacity is the energy (measured in joules) needed to raise the temperature of 1 g of a pure substance by 1°C.
   c. Water has the higher capacity, which means that it can absorb more heat energy (1 mark), than ethanol and so the ethanol would show the higher temperature. (1 mark)
   d. 7.7 kJ
5. a. E.   b. C.   c. B.
   d. E.   e. B.   f. B.
   g. D.   h. A.

12 Water as a solvent

12.2 Solution processes

Practice problem 1

Ethanol dissolves in water because its molecules are both polar molecules and have H bonded to O. Hydrogen bonds are formed between the water molecules and the hydrogen atom bonded to the oxygen atom in ethanol.

12.2 Exercise

1. The solute is the dissolved substance in solution. The solvent is the substance that the solute dissolves in. For example, salt (solute) dissolves in water (solvent).
2. Water molecules have a polar arrangement of the oxygen and hydrogen atoms. One side (hydrogen) has a positive electrical charge and the other side (oxygen) has a negative charge. This can disrupt attractive forces that hold the ionic compounds together and thus dissolve it. Detailed responses can be found in the worked solutions in your digital formats.
3. The ionic bonds between the Li^+ and Br^- ions must be broken, and the hydrogen bonds between the water molecules must be broken. New ion–dipole interactions are formed between the ions and water.
4. The negative ends of water molecules are attracted to the sodium ions and the positive ends of the water molecules are attracted to the hydroxide ions. This pulls the ions from the lattice and the water molecules. Detailed responses can be found in the worked solutions in your digital formats.
5. a. $CuCl_2(s) \xrightarrow{H_2O(l)} Cu^{2+}(aq) + 2Cl^-(aq)$
   b. $Na_3PO_4(s) \xrightarrow{H_2O(l)} 3Na^+(aq) + PO_4^{3-}(aq)$
   c. $(NH_4)_2SO_4(s) \xrightarrow{H_2O(l)} 2NH_4^+(aq) + SO_4^{2-}(aq)$
6. a. $HNO_3(l) + H_2O(l) \rightarrow H_3O^+(aq) + NO_3^-(aq)$
   b. $H_2SO_4(l) + H_2O(l) \rightarrow H_3O^+(aq) + HSO_4^-(aq)$
   c. $HBr(l) + H_2O(l) \rightarrow H_3O^+(aq) + Br^-(aq)$
7. Oil is insoluble in water because it consists of non-polar molecules with weak dispersion forces but water is polar with strong hydrogen bonds between its molecules. This hydrogen bonding holds the water molecules together and so they do not interact with the oil molecules.
8. Miscible: Can be mixed to form a solution.
9. Cordial, vinegar, wine would be miscible with water.
10. a. Ion–dipole interactions
    b. Ion–dipole interactions
    c. Hydrogen bonding
11. Both NH_3 and HCl are polar molecules but ammonia, NH_3, dissolves by forming hydrogen bonds with water and HCl ionises in water and the ions become hydrated.
12. A solution or liquid that is able to conduct electricity because it has mobile ions.

12.3 Precipitation reactions

Practice problem 2

$MgCl_2$ is soluble.

Practice problem 3

$2KOH(aq) + Ca(NO_3)_2(aq) \rightarrow 2KNO_3(aq) + Ca(OH)_2(s)$

Practice problem 4

$Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$
Spectator ions: Na^+, Cl^-

12.3 Exercise

1. a. When two solutions are mixed and a solid is formed.
   b. Two solutions are mixed together and a cloudy solid forms.
2. a. Potassium ion K^+, sulfate ion SO_4^{2-}
   b. Ammonium ion $NH_4{}^+$, phosphate ion PO_4^{3-}
   c. Silver ion Ag^+, nitrate ion NO_3^-
3. a. Insoluble b. Soluble
   c. Insoluble d. Soluble
4. a. Not a precipitation reaction as no precipitate formed.
   b. Not a precipitation reaction as two solutions have not been mixed to form this solid.
   c. This is a precipitate reaction because two solutions have been combined and a solid product has been formed.
5. a. All sodium compounds are soluble.
   b. No potassium compounds are insoluble.
   c. These ions combine with silver to form soluble salts: nitrate, nitrite, ethanoate.
6. a. There will be a precipitate of iron (lll) hydroxide.
   b. There will be a precipitate of barium sulfate.
   c. There will be no precipitate.
7. a. $2LiOH(aq) + Ca(NO_3)_2(aq) \rightarrow Ca(OH)_2(s) + 2LiNO_3(aq)$
   b. $Na_2S(aq) + (CH_3COO)_2Pb(aq) \rightarrow PbS(s) + 2(CH_3COO)Na(aq)$
   c. $2(NH_4)_3PO_4(aq) + 3CaCl_2(aq) \rightarrow Ca_3(PO_4)_2(s) + 6NH_4Cl(aq)$
8. a. $BaCl_2(aq) + MgSO_4(aq) \rightarrow BaSO_4(s) + MgCl_2(aq)$
   $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$
   Spectator ions: Mg^{2+}, Cl^-
   b. $Ca(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow 2NaNO_3(aq) + CaCO_3(s)$
   $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$
   Spectator ions: Na^+, $NO_3{}^-$
9. Combine the two solutions, filter the mixture and place the filter paper with the zinc carbonate precipitate in an oven at 110°C to dry overnight.
10. Silver nitrate solution
    $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

12.4 Solvent properties of water in biological, domestic or industrial concepts

12.4 Exercise

1. Allows the transportation of nutrients and other soluble chemicals around the body to different cells and also carries waste away.
2. Oxygen and carbon dioxide are slightly soluble in water so fish can have oxygen for respiration and plants can have carbon dioxide for photosynthesis.
3. Example: Vinegar is ethanoic acid in water. Soft drink has dissolved sugar, colouring and carbon dioxide.
4. Example: Fertilisers can be sprayed on crops and then the solution can be absorbed through plant roots.
   Problem: Excess nutrients can be washed into rivers and lakes, promoting the growth of algae. These use up oxygen in the decay process so fish are unable to survive.
5. Example: Chemicals can be diluted with water for reasons of safety. Water is economical to use as a solvent, and can be used to transport waste materials.
6. Example: Contaminated water may be difficult to dispose of or it might enter waterways and cause pollution. Detailed responses can be found in the worked solutions in your digital formats.

12.5 Review

12.5 Exercises

12.5 Exercise 1: Multiple choice questions

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

12.5 Exercise 2: Short answer questions

1. a. The hydrogen bonded to the oxygen atom in ethanol has a slight positive charge due to the highly polar nature of the hydrogen to oxygen bond, so it is this hydrogen that will be attracted to the negative end of the water molecule. The electronegativities of hydrogen and carbon in the hydrogen–carbon bond are very close so those bonds have little polarity and so are not attracted to the water molecules.
   b. Hydrogen bonds can form between the hydroxy group in ethanol and water so ethanol is soluble in water. Ethane is a non-polar molecule and so does not have significant interactions with polar water.
2. The negative ends of the polar water molecule are attracted to the silver ions and the positive ends of the water molecules are attracted to the negative nitrate ions in the ionic lattice. Ion–dipole bonds are formed and the ions are separated by the water molecules that surround each ion. The ions are now described as being hydrated.
3. a. $KOH\,(s) \xrightarrow{H_2O(l)} K^+(aq) + OH^-(aq)$
   b. $Na_3PO_4\,(s) \xrightarrow{H_2O(l)} 3Na^+(aq) + PO_4^{3-}(aq)$
4. There is more oxygen dissolved in cold water because gases are more soluble at lower temperatures. Aquatic organisms are less likely to survive if the temperature increases and less is oxygen available.

5.

Description	Compound
This compound is insoluble	CH_4
This compound ionises when dissolved in water	HNO_3
This compound forms hydrogen bonds when dissolved in water	CH_3OH
This compound dissociates when dissolved in water	KI

6. a. Yes, all nitrite compounds are soluble.
   b. Yes, all sodium compounds are soluble.
   c. No, most sulfates are soluble but lead sulfate is an exception.
   d. No, most carbonates are insoluble.
   e. Yes, all ammonium compounds are soluble.
   f. No, most sulfides are insoluble.
7. a. Potassium chloride is ionic.
   b. Ammonia is polar.
   c. Methane is non-polar.
8. a. CaI_2 is soluble; most iodide compounds are soluble.
   b. KOH is soluble; all potassium compounds are soluble.
   c. AgCl is insoluble; it is an exception because most chlorides are soluble.
   d. AgI is soluble; most iodides are soluble.
   e. $BaSO_4$ is insoluble; it is an exception because most sulfates are soluble.
   f. NH_4Cl is soluble; all ammonium compounds are soluble.
   g. Na_2S is soluble; all sodium compounds are soluble.
   h. $MgCl_2$ is soluble; most magnesium compounds are soluble.
   i. CuS is insoluble; most sulfides are insoluble except Na_2S and K_2S.
   j. $Al(OH)_3$ is insoluble; most hydroxides are insoluble.
9. a. $BaCl_2(aq) + (NH_4)_2CO_3(aq) \rightarrow BaCO_3(s) + 2NH_4Cl\ (aq)$
   b. $2AgNO_3(aq) + H_2S(g) \rightarrow Ag_2S(s) + 2HNO_3(aq)$
   c. $K_2CrO_4(aq) + Pb(NO_3)_2(aq) \rightarrow 2KNO_3(aq) + PbCrO_4(s)$
10. a. Yes, AgCl will precipitate.
    b. Yes, $BaSO_4$ will precipitate.
    c. No, $NaSO_4$ is soluble so no precipitate will form.
    d. No, the products would be NaCl and KNO_3 which are both soluble and so all of the ions would be dissociated.
    e. Yes, $Ca_3(PO_4)_2$ will precipitate.
    f. Yes, PbS will precipitate.
    g. No, all nitrate and ammonium compounds are soluble.
    h. No, most chlorides and bromides are insoluble.
11. AgCl: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

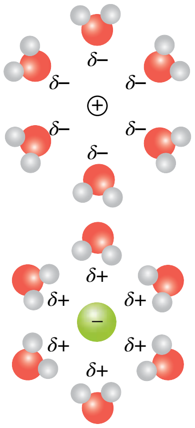
    $BaSO_4$: $Ba^{2+}(aq) + SO_4{}^{2-}(aq) \rightarrow BaSO_4(s)$

    $Ca_3(PO_4)_2$: $3Ca^{2+}(aq) + 2PO_4{}^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$

    PbS: $Pb^{2+}(aq) + S^{2-}(aq) \rightarrow PbS(s)$
12. a. $K_2SO_4(aq) + Ca(NO_3)_2(aq) \rightarrow 2KNO_3(aq) + CaSO_4(s)$
       $SO_4{}^{2-}(aq) + Ca^{2+}(aq) \rightarrow CaSO_4(s)$
    b. $Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow 2KNO_3(aq) + PbI_2(s)$
       $Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$
    c. $3MgCl_2(aq) + 2Na_3PO_4(aq) \rightarrow 6NaCl(aq) + Mg_3(PO_4)_2(s)$
       $3Mg^{2+}(aq) + 2PO_4{}^{3-} \rightarrow Mg_3(PO_4)_2(s)$
    d. $K_2S(aq) + ZnCl_2(aq) \rightarrow ZnS(s) + 2KCl(aq)$
       $S^{2-}(aq) + Zn^{2+}(aq) \rightarrow ZnS(s)$
    e. No reaction
13. a. Ba^{2+} b. Pb^{2+} c. Ba^{2+} d. Ba^{2+}
14. a. Chlorides, bromides and iodides will precipitate silver but not zinc.
    b. Sulfates precipitate barium but not magnesium.
    c. Sulfates precipitate lead but not silver.
    d. Barium precipitates sulfates but not chlorides.

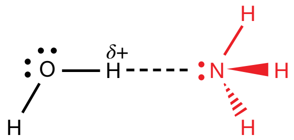
12.5 Exercise 3: Exam practice questions

1. Detailed responses can be found in the worked solutions in your digital formats.
   a. The glucose molecule has many –OH groups that are able to form hydrogen bonds with water. A glucose solution does not conduct electricity because glucose does not form ions.
   b. Sodium chloride can form ion–dipole interactions; however, benzene is non-polar.
   c. The hydrogen bonds between water molecules are stronger than the dipole–dipole forces between hydrogen chloride molecules.
2. When hydrogen bromide dissolves in water it ionises to form hydronium and bromide ions, which then become hydrated.



$HBr(s) + H_2O(l) \rightarrow H_3O^+(aq) + Br^-(aq)$

When ammonia dissolves in water it forms hydrogen bonds with water.



$NH_3(g) \rightarrow NH_3(aq)$

3. a. The breaking up or separation of an ionic compound into its component ions.
   b. The reaction of a molecular substance with water to produce ions.
   c. When sodium chloride is added to water, it dissociates into $Na^+(aq)$ and $Cl^-(aq)$ ions: i.e.
      $NaCl(s) \xrightarrow{H_2O(l)} Na^+(aq) + Cl^-(aq)$.
      Such solutions are good conductors of electricity because the ions are free to move.
   d. When hydrochloric acid is added to water it is ionised into $H_3O\ ^+(aq)$ and $Cl^-(aq)$ ions: i.e.
      $HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$.
      Such solutions are good conductors of electricity because the ions are free to move.

4. a. $BaSO_4$
   b. Sample answer:
   $Ba(OH)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaOH(aq)$
5. a. Test tube B (barium chloride)
   b. $BaCl_2(aq) + K_2CO_3(aq) \rightarrow BaCO_3(s) + 2KCl(aq)$
   c. $Ba^{2+}(aq) + CO_3{}^{2-}(aq) \rightarrow BaCO_3(s)$
   d. Spectator ions are present in the reaction but do not take part in the formation of the product $BaCO_3$. Spectator ions are $Cl^-(aq)$ and K^+.
   e. KCl(s): As the water is evaporated, the solution will form solid crystals of potassium chloride.

13 Reactions in water

13.2 Acids and bases

13.2 Exercise

1. Sample answers: ethanoic acid in vinegar, citric acid in citrus fruit and carbonic acid in soft drinks.
2. Covalent bonding
3. Hydrogen
4. Test with litmus paper; litmus paper is red in acid and blue in base.
5. Cleaners: ability to dissolve grease.
6. An alkali is a base that dissolves in water. A base is a chemical species that accepts a proton.
7. Ammonia (hydrogen bonding and some ionisation); sodium hydroxide, sodium carbonate, ammonium hydroxide (dissociation).
8. Electrolytes are solutions or liquids that conduct electricity because of the presence of ions.
9. Acids are corrosive so safety glasses, laboratory coat and gloves are recommended when working with acids.
10. Acids should be added to water and not the reverse because then the acids might splatter and cause injury.

13.3 Brønsted–Lowry theory of acids and bases

Practice problem 1

HNO_3 has become NO_3^-; it has lost a proton (H^+).
NH_3 has become NH_4^+; it has gained a proton (H^+).
Conjugate pairs are HNO_3/NO_3^- and NH_4^+/NH_3.

13.3 Exercise

1. Acid: H_2SO_4, base: H_2O
2. a. HBr b. HSO_4^- c. HSO_4^- d. H_3O^+
3. (b), (c), (e)
4. a. HSO_4^- b. HS^- c. S^{2-}
   d. OH^- e. NH_3
5. a. H_2O b. H_2CO_3 c. HS^-
   d. H_3O^+ e. HCN
6. $H_2O(l)$ and $NH_4^+(aq)$
7. a. Base b. Amphiprotic
   c. Amphiprotic d. Acid
   e. Amphiprotic
8. a. $H_3PO_4(aq) + H_2O(l) \rightleftharpoons H_2PO_4^-(aq) + H_3O^+(aq)$
   $H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$
   $HPO_4^{2-}(aq) + H_2O(l) \rightleftharpoons PO_4^{3-}(aq) + H_3O^+(aq)$
   b. $H_3PO_4(aq)/H_2PO_4^-(aq)$, $H_3O^+(aq)/H_2O(l)$
   $H_2PO_4^-(aq)/HPO_4^{2-}(aq)$, $H_3O^+(aq)/H_2O(l)$
   $HPO_4^{2-}(aq)/PO_4^{3-}(aq)$, $H_3O^+(aq)/H_2O(l)$
9. a. $HS^-(aq) + H_2O(l) \rightleftharpoons S^{2-}(aq) + H_3O^+(aq)$
   $HS^-(aq) + H_2O(l) \rightleftharpoons H_2S(aq) + OH^-(aq)$
   b. $HSO_4^-(aq) + H_2O(l) \rightleftharpoons SO_4^{2-}(aq) + H_3O^+(aq)$
   $HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_2SO_4(aq) + OH^-(aq)$
10. a. Triprotic b. Monoprotic
    c. Monoprotic d. Diprotic
    e. Monoprotic f. Diprotic

13.4 Concentrating on acids and bases

Practice problem 2

$n(NaOH) = 0.135$ M

13.4 Exercise

1. a. 0.130 mol b. 0.240 mol c. 0.0475 mol
2. 0.882 M
3. a. 0.714 L b. 5.00 L c. 10.4 L
4. 0.598 M
5. 0.00189 mol
6. A weak acid only partially ionises in water (only a small number of the acid molecules become ions).
7. Strong acids completely ionise to produce solutions that are good conductors of electricity;
   e.g. $HBr(aq) + H_2O(l) \rightarrow Br^-(aq) + H_3O^+(aq)$.
   Weak acids only partially ionise to produce solutions that are poor conductors of electricity;
   e.g. $HS^-(aq) + H_2O(l) \rightleftharpoons S^{2-}(aq) + H_3O^+(aq)$.
8. $HN(CH_3)_2(aq) + H_2O(l) \rightleftharpoons H_2N^+(CH_3)_2(aq) + OH^-(aq)$
9. a. HNO_3/NO_3^- and HCl/Cl^-
   b. Wear protective clothing, safety glasses and safety gloves, and work in a well-ventilated area.

13.5 pH and acidity

Practice problem 3

pH = 2

Practice problem 4

pH = 1.4

Practice problem 5

$[HNO_3] = 1.6 \times 10^{-2}$ M

Practice problem 6

$pH = -\log\left[4.2 \times 10^{-13}\right] = 12.4$

Practice problem 7

[KOH] = 0.0063 M

Practice problem 8

$c_2 = 1.02$ M

Practice problem 9

Water added = 450 mL

Practice problem 10

pH = 1.1

13.5 Exercise

1. a. $[OH^-] = 10^{-9}$ M
   b. $[H_3O^+] > [OH^-]$ so the solution is acidic
2. The pH scale provides a measure of the acidity of a substance. The formula for calculating pH is $pH = -\log_{10}[H_3O^+]$
3. A dilute solution has a small amount of solute in a particular volume of solvent whereas a concentrated solution has a large amount of solute in a given volume of solvent.
4. a. pH = 5 b. pH = 1
5. pH = 2.8
6. a. pH = 10 b. pH = 13
7. pH = 1.2
8. a. pH = 3.04 b. pH = 9.25
9. In a neutral solution $[H_3O^+] = [OH^-] = 10^{-7}$ M
   pH = 7
10. 31.3 mL

13.6 Reactions of acids

Practice problem 11

$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

Practice problem 12

$2CH_3COOH(aq) + Na_2CO_3(s) \rightarrow 2CH_3COONa(aq) + CO_2(g) + H_2O(l)$

$2H^+(aq) + Na_2CO_3(s) \rightarrow 2Na^+(aq) + CO_2(g) + H_2O(l)$

Practice problem 13

$2HCl(aq) + Mg(OH)_2(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

13.6 Exercise

1. See table at foot of the page.*
2. a. acid + metal → salt and hydrogen
   b. acid + metal carbonate → salt + water + carbon dioxide
   c. acid + metal hydroxide → salt + water
3. a. Detailed responses can be found in the worked solutions in your digital formats.
   $Zn(s) + 2HBr(aq) \rightarrow ZnBr_2(aq) + H_2(g)$
   b. $Ca(s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + H_2(g)$
   c. $2K(s) + 2HNO_3(aq) \rightarrow 2KNO_3(aq) + H_2(g)$
4. a. Detailed responses can be found in the worked solutions in your digital formats.
   $MgCO_3(s) + 2CH_3COOH(aq) \rightarrow Mg(CH_3COO)_2 + CO_2(g) + H_2O(l)$
   b. $Na_2CO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$
   c. $MnCO_3(s) + 2HNO_3(aq) \rightarrow Mn(NO_3)_2(aq) + CO_2(g) + H_2O(l)$
5. a. Detailed responses can be found in the worked solutions in your digital formats.
   $2NaOH + H_2SO_4 \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
   b. $Ca(OH)_2 + 2HCl \rightarrow CaCl_2(aq) + 2H_2O(l)$
   c. $Mg(OH)_2 + 2HNO_3 \rightarrow Mg(NO_3)_2(aq) + 2H_2O(l)$
6. a. $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$; spectator ion Br^-
   b. $MgCO_3(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + CO_2(g) + H_2O(l)$; spectator ion CH_3COO^-
   c. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$; spectator ions Na^+, SO_4^{2-}
7. a. $Ca(s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + H_2(g)$
   $Ca(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$
   b. $ZnCO_3(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + CO_2(g) + H_2O(l)$
   $ZnCO_3(s) + 2H^+(aq) \rightarrow Zn^{2+} + CO_2(g) + H_2O(l)$
   c. $Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$
   $CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$
8. a. Collect the gas from above the reaction flask in a test tube. Place a lit match at the mouth of the test tube. If a pop is heard then the gas is hydrogen.
   b. $2H_2(g) + O_2(g) \rightarrow H_2O(l)$
   c. No, there may not be enough gas produced or collected.
9. a. Direct the gas downwards into a test tube containing limewater ($Ca(OH)_2$). If a white precipitate is observed, the gas is likely to be carbon dioxide.
   b. $CO_2(g) + Ca(OH)_2(aq) \rightarrow CaCO_3(s) + 2H_2O(l)$
10. a. $Ni(s) + H_2SO_4(aq) \rightarrow NiSO_4(aq) + H_2(g)$
    $Ni(s) + 2H^+(aq) \rightarrow Ni^{2+}(aq) + H_2(g)$
    Spectator ions SO_4^{2-}
    b. $2Al(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$
    $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$
    Spectator ions SO_4^{2-}
    c. $CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + CO_2(g) + H_2O(l)$
    $CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+} + CO_2(g) + H_2O(l)$
    Spectator ions NO_3^-

13.7 Acids and the environment

13.7 Exercise

1. Acid rain is rain made acidic by mixing with atmospheric pollution, resulting a pH of less than 5.
2. Carbon dioxide and water produces carbonic acid
   $CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$
3. Carbonic acid is a weak acid and partially ionises.
   $H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$
4. $SO_2(g) + H_2O(l) \rightleftharpoons H_2SO_3(aq)$
   $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$
   $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$
5. Burning of coal and other fossil fuels together with active volcanoes produce SO_2.

1*

Acid	Formula of acid	Type of salt formed from the acid	Ion present
Hydrochloric	HCl	Chloride	Cl^-
Nitric	HNO_3	Nitrate	NO_3^-
Sulfuric	H_2SO_4	Sulfate	SO_4^{2-}
Ethanoic acid	CH_3COOH	Ethanoate	CH_3COO^-

6. Possible control methods include:
   - using natural gas as it produces less SO_2
   - using coal with a lower SO_2 content
   - removing SO_2 before emitting waste gases and using it in other industries.
7. Oceans are slightly alkaline but the added CO_2 being absorbed by the oceans is lowering the pH.
8. 1.3×10^{-6} M
9. Problems that might occur as oceans become less basic include:
   - coral bleaching
   - interference with marine organisms
   - reproductive disorders in some organisms
   - interference with shell building.
10. Calcium carbonate is a constituent of corals and the shells of shellfish. The presence of carbon dioxide decreases the already small concentration of carbonate ions present:
    $CO_2(g) + H_2O(l) + CO_3^{2-}(aq) \rightarrow 2HCO_3^-(aq)$
    If the carbonate ion concentration decreases too much, the shells of marine organisms may be affected.

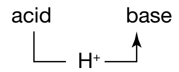
13.8 Review

13.8 Exercises

13.8 Exercise 1: Multiple choice questions

1. C.
2. B.
3. D.
4. C.
5. C.
6. B.
7. B.
8. C.
9. B.
10. A.
11. D.
12. D.
13. B.
14. A.
15. A.

13.8 Exercise 2: Short answer questions

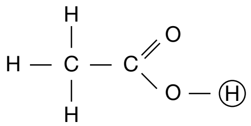
1. a. Acids taste sour, are corrosive, dissolve in water to produce a solution that conducts electricity, form solutions with a pH less than 7, change the colour of indicators and are neutralised by bases.
   b. Corrosive: $Fe(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2(g)$
   Neutralised by bases: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
2. a. A base reacts with an acid to neutralise it; a base is a proton acceptor. For example: MgO, Na_2CO_3.
   b. An alkali is a base that is soluble; for example, NaOH, KOH.
3. a. Acidic
   b. Acidic
   c. Both acidic and basic
   d. Acidic
   e. Acidic
   f. Basic
   g. Basic
   h. Both acidic and basic
   i. Basic
   j. Acidic
   k. Basic
4. a. A Brønsted–Lowry acid is a proton donor and a Brønsted–Lowry base is a proton acceptor.
   $HCl(aq) + H_2O \rightarrow H_3O^+(aq) + Cl^-(aq)$
   

   b. The gas hydrogen chloride is a molecular compound and, therefore, has no free ions or electrons to conduct a current. However, when hydrogen chloride is dissolved in water, it ionises to form H^+ and Cl^- ions, which are able to conduct a current.
5. $Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$
   $Al(OH)_3(s) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$
   $NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$
6. a. Conjugate acid–base pairs differ by one proton. The conjugate base of an acid is formed in a reaction where the acid loses a proton. The conjugate acid of a base is formed in a reaction where the base gains a proton.
   b. i. HCl ii. HCO_3^- iii. H_2SO_4 iv. H_3O^+ v. H_2O
   c. i. F^- ii. CO_3^{2-} iii. SO_4^{2-} iv. OH^- v. HS^-
7. $HCO_3^-(aq) + H_2O(l) \rightleftharpoons CO_3^{2-}(aq) + H_3O^+(aq)$
   $HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$
8. $H_2SO_4(aq) + H_2O(l) \rightarrow HSO_4^-(aq) + H_3O^+(aq)$
   $HSO_4^-(aq) + H_2O(l) \rightleftharpoons SO_4^{2-}(aq) + H_3O^+(aq)$
9. The strength of an acid is a measure of its ability to donate protons — strong acids are good proton donors while weak acids are poor proton donors. The concentration of an acid is a measure of the amount of solute in a unit volume of solution; for example, 4 M HCl is a concentrated solution while 0.010 M HCl has a low concentration. Glacial acetic acid (ethanoic acid) is a concentrated weak acid. Freyja is correct.
10. a. A weak base should be added, such as sodium carbonate.
    b. Alkaline
11. a. Iron, and some boron, manganese
    b. Calcium, magnesium, nitrogen, potassium
    c. Phosphorus
    d. Boron, iron, manganese,
    e. Boron, iron, manganese and potassium
    f. Phosphorus
    g. The soil for the lettuces may be lacking a particular nutrient. To test this, set up nine identical pots of seedlings using the gardener's soil and add a specified volume of pure water daily to pot 1 (as the control). For the remaining pots, add the same volume of solutions containing the same concentration of each of the elements listed. Maintain the same conditions of light and temperature for all pots. Observe over a specified period of time and measure the heights of the seedlings to see which has grown successfully.
12. a. pH = 3.0 b. pH = 3.3
13. pH is a logarithmic function meaning it changes by factors of 10. The effect of a dilution by a factor of 2 is shown in the worked solutions in your digital formats.
14. Detailed responses can be found in the worked solutions in your digital formats.
    a. 11 b. 10.7
15. a. $3Na_2S(s) + 2H_3PO_4(aq) \rightarrow 2Na_3PO_4(aq) + 3H_2S(g)$
    $2H^+(aq) + Na_2S(s) \rightarrow 2Na^+(aq) + H_2S(g)$

b. $2HCl(aq) + CuO(s) \rightarrow CuCl_2(aq) + H_2O(l)$
$2H^+(aq) + CuO(s) \rightarrow Cu^{2+}(aq) + H_2O(l)$
c. $H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + H_2O(l)$
$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

13.8 Exercise 3: Exam practice questions

1. D.
2. a. Monoprotic means it can donate only one proton.
b. $CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(s) + H_3O^+(aq)$
c. $CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COO^-(aq) + H_2O(l) + Na^+(aq)$
d. The equation in (b) does not react completely, ethanoic acid is only partially ionised in water.
e. Measure the conductivity. Weak acids only partly ionise, so there will not be a significant increase in conductivity.
f.



3. a. Amphiprotic means it can donate or receive a proton.
b. $HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$
$HCO_3^-(aq) + H_2O(l) \rightleftharpoons CO_3^{2-}(aq) + H_3O^+(aq)$
c. The first reaction occurs to a greater extent than the second.
4. a. Ammonium sulfate could be produced by combining ammonia and sulfuric acid.
b. $2NH_3(g) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq)$
c. Full ionic equation: $2\ NH_3(g) + 2H^+(aq) \rightarrow 2NH_4^+$
d. Ammonium sulfate would be obtained by evaporating the water from the solution.
e. $(NH_4)_2SO_4(s) \rightarrow 2NH_4^+(aq) + SO_4^{2-}(aq)$
5. a. $CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$
b. $H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+ + HCO_3^-$
$H_2CO_3(aq)$ is a weak acid that reacts with water to produce the hydronium ion and the hydrogen carbonate ion. The acidity is due to the presence of the hydronium ion.
c. $H_2SO_4(aq) + CaCO_3(s) \rightarrow CaSO_4(aq) + H_2O(l) + CO_2(g)$
d. Sulfur dioxide (SO_2), sulfur trioxide (SO_3) and nitrogen dioxide (NO_2)

14 Redox reactions in water

14.2 Redox reactions and agents

Practice problem 1

$2Br^-$ has lost two electrons to become Br_2.
Cl_2 has gained two electrons to become $2Cl^-$.
Br^- is oxidised, while Cl_2 is reduced; that is, the chlorine molecule is reduced.

Practice problem 2

a. $2Ag^+(aq) + Cu(s) \rightarrow Cu^{2+}(aq) + 2Ag(s)$

b. $Ag^+(aq) + e^- \rightarrow Ag(s)$

$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

c. Conjugate pairs: $Ag^+(aq)/Ag(s)$ and $Cu^{2+}(aq)/Cu(s)$

Practice problem 3

$SO_2(g) + 2H_2O(l) \rightarrow SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$

14.2 Exercise

1. Sample answer: fireworks, explosions, combustion, photosynthesis, respiration and metal corrosion.
2. a. Oxidation b. Reduction c. Oxidation
d. Oxidation e. Reduction
3. a. $Pb^{2+}(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + Pb(s)$
b. $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$ and $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
c. Zinc metal
d. Lead ions
e. Lead ions are the oxidising agents.
f. Zinc metal is the reducing agent.
4. This reaction is not an acid–base reaction because no protons have been transferred. Instead, it is a redox reaction because electrons have been transferred. Magnesium atoms have been oxidised and the hydrogen ions have been reduced.
$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$ and $2H^+(aq) + 2e^- \rightarrow H_2(g)$
5. a. Reduction: $Ag^+(aq) + e^- \rightarrow Ag(s)$.
Oxidation: $Pb(s) \rightarrow Pb^{2+} + 2e^-(aq)$.
b. Silver ions accept electrons to form silver metal.
c. Lead metal, Pb(s), is oxidised to lead ions, $Pb^{2+}(aq)$.
d. Lead metal
e. Lead metal
f. $Ag^+(aq)/Ag(s)$ and $Pb^{2+}/\ Pb(s)$
6. a. $2H^+$ accepts electrons to form $H_2(g)$.
b. H^+ is reduced to H_2 because it is the species accepting electrons.
c. Al(s) is the reducing agent.
d. Al(s). The reducing agent has been oxidised.
e. Al is the reductant as it donates electrons, $Al \rightarrow Al^{3+}$.
f. The oxidising agent is H^+ because it accepts electrons.
g. $Al^{3+}(aq)/Al(s)$ and $H^+(aq)\ /H_2(g)$
h. $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$
7. a. $Ca(s) \rightarrow Ca^{2+}(s) + 2e^-$
$Cl_2(g) + 2e^- \rightarrow 2Cl^-(s)$
$Ca^{2+}(s)/Ca(s)$ and $Cl_2(g)\ /Cl^-(s)$
b. $Mg(s) \rightarrow Mg^{2+}(s) + 2e^-$
$S(s) + 2e^- \rightarrow S^{2-}(s)$
$Mg^{2+}(s)/Mg(s)$ and $S(s)/S^{2-}(s)$
c. $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$
$Ag^+(aq) + e^- \rightarrow Ag(s)$
$Al^{3+}(aq)\ /Al(s)$ and $Ag^+(aq)/Ag(s)$
8. a. $NO_3^-(aq) + 2H^+(aq) + e^- \rightarrow NO_2(g) + H_2O(l)$
b. $ClO^-(aq) + 2H^+(aq) + 2e^- \rightarrow Cl^-(aq) + H_2O(l)$
c. $BrO_3^-(aq) + 2H^+(aq) + e^- \rightarrow BrO_2(aq) + H_2O(l)$
9. Responses should include the steps towards these half-equations. Fully worked solutions are available in your digital formats.
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$:
$NO_2^-(aq) + H_2O(l) \rightarrow NO_3^-(aq) + 2H^+(aq) + 2e^-$:
10. a. $I_2(s) + H_2S(g) \rightarrow 2I^-(aq) + S(s) + 2H^+$
b. $5SO_3^{2-}(aq) + 2MnO_4^- + 6H^+ \rightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(l)$
c. $3Cu(s) + 2NO_3^-(aq) + 8H^+(aq) \rightarrow 3Cu^{2+}(aq) + 2NO(g) + 4H_2O(l)$

14.3 Oxidation numbers

Practice problem 4

a. The oxidation number of phosphorus is +5.
b. The oxidation number of nitrogen is +5.
c. The oxidation number of Cl^- is –1.
d. The oxidation number of K is 0.

Practice problem 5

Oxidation numbers: $\overset{0}{Pb}$, $\overset{0}{Br}$, $\overset{2+}{Pb}$, $\overset{-1}{Br}$
Lead is the reducing agent.
Br_2 is the oxidising agent.
Electrons have been transferred from the lead to the bromine and so this is a redox reaction.

14.3 Exercise

1. **a.** Oxidation number P = 0
   b. Oxidation number $F_2 = 0$
   c. Oxidation number is $Zn_{2+} = 2+$.
   d. Oxidation number of H = +1, O = −2, C = +4
   e. Oxidation number of Na = +1 (ion), O = −2, N = +5
   f. Oxidation number of O = −2, Be = +2
   g. Oxidation number of $Ag^+ = +1$, S = −2
2. NH_4^+, N_2, NO_2, NO_3^-
3. The maximum oxidation number sulfur can have is six.
4. **a.** $\overset{0}{Br_2}(l) + \overset{+4}{S}\overset{-2}{O_2}(g) \rightarrow 2\overset{-1}{Br^-}(aq) + \overset{+6}{S}\overset{-2}{O_4}{}^{2-}(aq)$
   is a redox reaction.
   b. $\overset{+1}{K}\overset{-2}{O}\overset{+1}{H}(aq) + \overset{+1}{H}\overset{+5}{N}\overset{-2}{O_3}(aq) \rightarrow \overset{+1}{K}\overset{+5}{N}\overset{-2}{O_3}(aq) + \overset{+1}{H_2}\overset{-2}{O}(aq)$
   is not a redox reaction.
   c. $\overset{+3}{Fe^{3+}}(aq) + 3\overset{-2}{O}\overset{+1}{H^-}(aq) \rightarrow \overset{+3}{Fe}(\overset{-2}{O}\overset{+1}{H})_3(s)$
   is not a redox reaction.
   d. $5\overset{0}{Zn}(s) + 2\overset{+5}{N}\overset{-2}{O_3^-}(aq) + 12\overset{+1}{H^+}(aq) \rightarrow 5\overset{+2}{Zn^{2+}}(aq)$
   $+\overset{0}{N_2}(g) + 6\overset{+1}{H_2}\overset{-2}{O}(1)$
   is a redox reaction.
5. **a.** **i.** $\overset{0}{Ca}(s) + 2\overset{+1}{H_2}\overset{-2}{O}(1) \rightarrow \overset{+2}{Ca}(\overset{-2}{O}\overset{+1}{H})_2(aq) + \overset{0}{H_2}(g)$
   ii. Calcium has been oxidised and water has been reduced.
   iii. Oxidising agent: H_2O; reducing agent: Ca.
   b. **i.** $\overset{+4}{Pb}\overset{-2}{O_2}(s) + \overset{-1}{I^-} \rightarrow \overset{+2}{Pb^{2+}}(aq) + \overset{0}{I_2}(g)$
   ii. Lead oxide has been reduced and iodide ions have been oxidised.
   iii. Oxidising agent: PbO_2; reducing agent: I^-.
   c. **i.** $\overset{-1}{I^-}(aq) + \overset{+3}{N}\overset{-2}{O_2^-}(aq) \rightarrow \overset{0}{I_2}(s) + \overset{+2}{N}\overset{-2}{O}(g)$
   ii. The nitrite ions, NO_2^-, have been reduced and iodide ions have been oxidised.
   iii. Oxidising agent: NO_2^-; reducing agent: I^-.
   d. **i.** $\overset{+1}{H_2}\overset{-1}{O_2}(aq) + \overset{+7}{Mn}\overset{-2}{O_4^-} \rightarrow \overset{+2}{Mn^{2+}}(aq) + \overset{0}{O_2}(g)$
   ii. The permanganate ions, MnO_4^-, have been reduced and the hydrogen peroxide molecules have been oxidised.
   iii. Oxidising agent: MnO_4^-; reducing agent: H_2O_2.
6. **a.** Oxidation number of O in H_2O_2 is –1
   b. Sn^{2+} is oxidised.
   c. H_2O_2 is reduced.
   d. Hydrogen peroxide is reduced, so it is the oxidising agent.
   e. Sn^{2+} (tin ll ion) is oxidised, so it is the reducing agent.
7. See table at foot of the page.*

14.4 Reactivity series of metals

Practice problem 6

a. $Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)$
b. $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$
c. Zn is the reducing agent.
Pb^{2+} is the oxidising agent.

14.4 Exercise

1. **a.** Silver (Ag), gold (Au)
   b. Iron (Fe) will displace tin ions (Sn^{2+}) but will be displaced by zinc (Zn)
   c. Potassium (K), sodium (Na), calcium (Ca)
   d. Tin (Sn) and lead (Pb)
   e. Calcium (Ca)
   f. Copper (Cu) and mercury (Hg)
2. **a.** Group 1 metals
   b. The reactivity increases as you go down a group.

7*

Change	Oxidation	Reduction	Neither
$O_2 \rightarrow O_3$			✓
$MnO_2 \rightarrow MnO_4^-$	✓		
$P_2O_5 \rightarrow P_4H_{10}$			✓
$NH_3 \rightarrow NO_2$	✓		
$Fe^{2+} \rightarrow Fe^{3+}$	✓		
$S_2O_3^{2-} \rightarrow S_2O_4^{2-}$	✓		

3. a. The blue colour, $CuSO_4$, of the solution fades as the copper metal forms. This displacement reaction occurs because copper (Cu) is more reactive than iron (Fe).
   b. The products are copper and iron sulfate solution.
4. a. $Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$
   Reducing agent: Mg; oxidising agent: Cu^{2+}.
   b. $2Al(s) + 3Pb^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Pb(s)$.
   Reducing agent: Al; oxidising agent: Pb^{2+}.
   c. No reaction because copper is less reactive than iron.
   d. No reaction because silver is less reactive than zinc.
5. a. $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$
   $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
   $Pb^{2+}(aq) + Fe(s) \rightarrow Pb(s) + Fe^{2+}(aq)$
   b. $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$
   $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
   $Sn^{2+}(aq) + Zn(s) \rightarrow Sn(s) + Zn^{2+}(aq)$
   c. No reaction because iron is less reactive than aluminium and so will not displace aluminium from solution.
   d. $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$
   $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$
   $3Zn^{2+}(aq) + 2Al(s) \rightarrow 3Zn(s) + 2Al^{3+}(aq)$
6. a. The aluminium metal may have had an unreactive oxide coating.
   b. Use sand paper or emery paper to strip off the coating to expose the metal before repeating the experiment.
7. $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$
   The oxidising agent: H_2O
   The reducing agent: Na
8. Aluminium oxide
   $2Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$
9. a. $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
   Magnesium chloride solution and hydrogen gas
   b. $Mg(s) + 2H_2O(g) \rightarrow Mg(OH)_2(aq) + H_2(g)$
   Magnesium hydroxide and hydrogen
   c. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$
   Magnesium oxide
10. Gold would be too expensive and too soft, and magnesium would be too reactive. Copper would be a better choice or perhaps an alloy of copper. Factors to be considered include availability and cost of metal, lack of reactivity, hardness, ability to be shaped, and ease and cost of manufacture.

14.5 Corrosion

14.5 Exercise

1. a. Corrosion is the oxidation of metals by materials in their environment.
   b. Corrosion is a redox reaction.
2. Metals have different levels of reactivity, which results in some metals reacting with air and moisture to corrode, while other metals do not react.
3. Dry corrosion occurs when metals react directly with oxygen. For example, when sodium reacts directly with oxygen and so must be stored in oil. Wet corrosion involves the presence of water. For example, iron in a damp environment.
4. Rusting is the wet corrosion of iron. Corrosion refers to the oxidation of metals in general.
5. Water and oxygen are necessary for rusting to occur, and this process is speeded up in the presence of salty water.
   $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
   $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$
   $2Fe(s) + O_2(g) + 2H_2O(l) \rightarrow 2Fe(OH)_2(s)$
6. Set up two identical test tubes with the same volume of each type of water at the same temperature and place identical iron nails in each. Observe over the same period of time.
7. Iron is versatile, relatively abundant and relatively economically extracted; it is strong and as a steel alloy is less susceptible to corrosion.
8. Corrosion can cause weaknesses in constructions that can result in breakage, corroded metal cannot conduct electricity effectively and corroded pipes may leak.
9. a. A bike chain is protected by grease.
   b. A metal dish-draining rack is protected by a waterproof plastic coating.
   c. A metal roof is protected by painting.
   d. A ship's hull is protected by sacrificial protection.
10. By attaching a block of a slightly more reactive metal like zinc to less reactive metals like iron, the more reactive metal oxidised instead of the iron.

14.6 Review

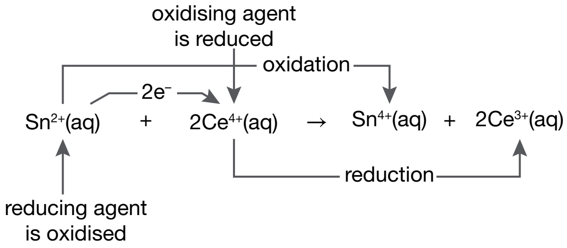
14.6 Exercises

14.6 Exercise 1: Multiple choice questions

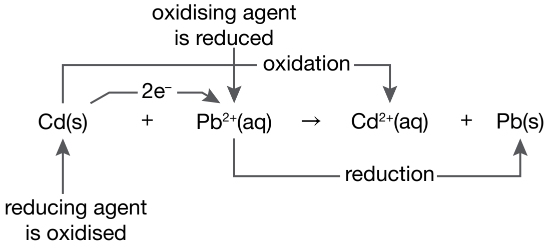
1. D.
2. B.
3. D.
4. D.
5. D.
6. A.
7. D.
8. A.
9. C.
10. B.
11. A.
12. A.

14.6 Exercise 2: Short answer questions

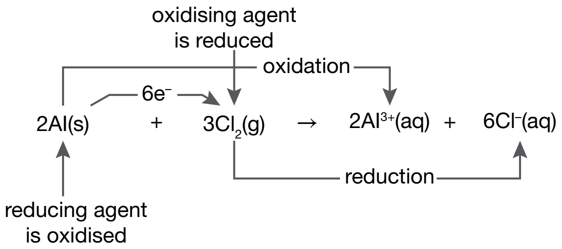
1. Oxidation is the gain of oxygen; reduction is the loss of oxygen.
2. Oxidation is the loss of electrons; reduction is the gain of electrons. (OIL RIG)
3. An oxidising agent is an electron acceptor, it is reduced.
4. a.



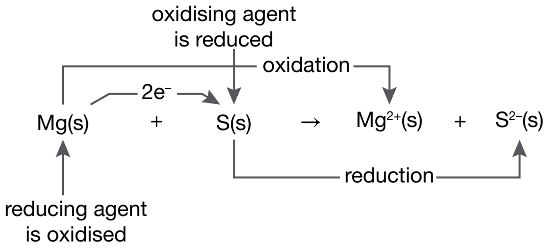
   b.



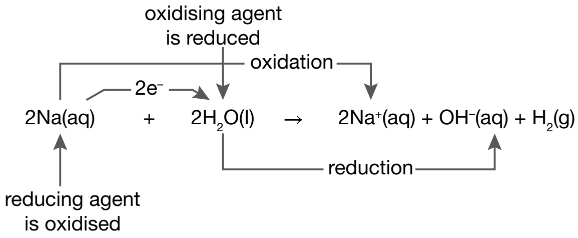
c.



d.



e.



5. a. $2Br^-(aq) + SO_4^{2-}(aq) + 4H^+(aq) \rightarrow Br_2(l) + SO_2(g) + 2H_2O(l)$

b. $2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$

c. $I_2(s) + H_2S(g) \rightarrow 2I^-(aq) + S(s) + 2H^+(aq)$

d. $3Cu(s) + 2HNO_3(aq) + 6H^+(aq) \rightarrow 3Cu^{2+}(aq) + 2NO(g) + 4H_2O(l)$

e. $3CuO(s) + 2NH_3(g) \rightarrow 3Cu(s) + 3H_2O(l) + N_2(g)$

f. $2Cr_2O_7^{2-}(aq) + 16H^+(aq) + 3CH_3CH_2OH(g) \rightarrow 4Cr^{3+}(aq) + 11H_2O(l) + 3CH_3COOH(aq)$

6. Displacement occurs when metal ions are converted into their metallic forms by another metal. For example, zinc displaces copper ions in the reaction:
$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$.

7. Groups 1 and 2

8. $Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)$

9. a. $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

b. $Fe(s) + Sn^{2+}(aq) \rightarrow Fe^{2+}(aq) + Sn(s)$

c. $Pb(s) + 2Ag^+(aq) \rightarrow Pb^{2+}(aq) + 2Ag(s)$

d. $2Al(s) + 3Zn^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Zn(s)$

In each case, the reducing agent has been placed first in the equation and the oxidising agent is second.

10. a. $Sn(s) + Cu^{2+}(aq) \rightarrow Sn^{2+}(aq) + Cu(s)$
$Cd(s) + Sn^{2+}(aq) \rightarrow Cd^{2+}(aq) + Sn(s)$

b. Cd > Sn > Cu (in order of decreasing reactivity).

c. Only reaction (iv) will occur.

11. a. $E > A > F > C > D > B$.

b. Sample answer: Mg, Zn, Fe, Sn, Pb, Cu

12. Balanced reaction:
$4Fe(OH)_2(s) + O_2(g) + 2H_2O(l) \rightarrow 4Fe(OH)_3(s)$
The corrosion of iron metal is a redox reaction because it involves the transfer of electrons. Iron loses electrons to form iron ions and oxygen in the presence of water gains electrons to form hydroxide ions.

13. Salt forms an electrolyte on the surface of the metal.

14. a. The steel will rust underneath the plastic.

b. The steel will rust underneath the paint.

c. The steel will rust underneath the tin.

d. No effect will occur — the remaining zinc will continue to protect the steel.

15. a. Coating the metal with a less reactive metal

b. If the tin is scraped away as the screw is tightened, the screw could corrode.

16. a. The block of zinc will corrode instead of the iron.

b. No protection would be provided, as lead is less reactive than iron.

c. Painting or coating the statue with an unreactive material.

14.6 Exercise 3: Exam practice questions

1. a. Redox reaction **b.** Not a redox reaction
c. Redox reaction **d.** Redox reaction

2. a. Reduction: $Fe_2O_3(s) + 6H^+(aq) + 6e^- \rightarrow 2Fe(s) + 3H_2O(l)$
Oxidation: $(CO(g) + H_2O(l) \rightarrow CO_2(g) + 2H^+(aq) + 2e^-) \times 3$
Balanced equation: $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(s)$

b. Oxidation: $(P_4(s) + 16H_2O(l) \rightarrow 4H_3PO_4(aq) + 20H^+(aq) + 20e^-) \times 3$
Reduction: $(HNO_3(aq) + 3H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(l)\ ) \times 20$
Balanced equation: $3P_4(s) + 20HNO_3(aq) + 8H_2O(l) \rightarrow 12H_3PO_4(aq) + 20NO(g)$

c. Reduction: $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$
Oxidation: $(H_2C_2O_4(aq) \rightarrow 2CO_2(g) + 2H^+(aq) + 2e^-) \times 3$
Balanced equation: $Cr_2O_7^{2-}(aq) + 3H_2C_2O_4(aq) + 8H^+(aq) \rightarrow 2Cr^{3+}(aq) + 6CO_2(g) + 7H_2O(l)$.

d. Oxidation: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
Reduction: $(HNO_3(aq) + H^+(aq) + e^- \rightarrow NO_2(g) + H_2O(l)) \times 2$
Balanced equation: $Cu(s) + 2HNO_3(aq) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$

e. Oxidation: $PbS(s) + 4H_2O(l) \rightarrow PbSO_4(s) + 8H^+(aq) + 8e^-$
Reduction: $(H_2O_2(l) + 2H^+(aq) + 2e^- \rightarrow 2H_2O(l)\ ) \times 4$
Balanced equation: $PbS(s) + 4H_2O_2(l) \rightarrow PbSO_4(s) + 4H_2O(l)$

3. a. $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$

b. $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$

c. Pb^{2+}

d. Sn

e. No. A further redox reaction will not occur because Mg^{2+} is a weaker oxidising agent than Sn^{2+}.

4. a. A^+, B^{2+}, C^+, D^{3+} **b.** $B > D > A > C$
c. $B > D > A > C$ **d.** B^{2+}

5. a. Rusting requires water and oxygen (air).

b. Salt, acids and pollutants

c. **i.** Very little water
  ii. No oxygen
  iii. Very little oxygen

15 Water sample analysis

15.2 The states of water and sampling protocols

15.2 Exercise

1. Responses should include details of why Australia has the highest per capita water storage rate in the world due to geography and climate. Responses should include explanations of rainfall and water management. Detailed responses can be found in the worked solutions in your digital formats.
2. If reticulated water is not available, water can be sourced from:
   - rainfall collected off roofs into rainwater tanks
   - catchment run off into dams
   - pumping water from natural waterbodies, such as rivers
   - accessing groundwater via a bore pump.
3. Strict protocols exist for the collection of water samples in order to produce accurate analytical results and to ensure consistency of sampling methods and result reporting. These apply before, during and after sampling takes place.
4. Duplicates are taken to increase the precision of the analysis. The results are accepted only if they are within a range of 10 to 20 per cent of each other.
5. a. Sample collected in polyethylene or PTFE bottle, or borosilicate glass of 100 mL; fill container completely to exclude air; transport under ice; maximum holding time 28 days; store in refrigerator ($\leq$ 6°).
   b. Collected in polyethylene or PTFE of 500 mL; fill container completely to exclude air; transport under ice; acidify to preserve with nitric acid to pH $<$ 2; maximum holding time $<$ 6 months; or 7 days without acidifying sample; store in refrigerator ($\leq$ 6°C).
   c. Collected in polyethylene, PTFE or glass of 500 mL; transport under ice away from light; if no interfering compounds present, add sodium hydroxide solution to pH $\geq$ 12; maximum holding time 24 hours if sulphide present, otherwise 14 days; store in refrigerator ($\leq$ 6°C) in dark.
   d. Collected in polyethylene, PTFE or glass (PTFE preferred) in 500 mL; transport under ice; fill container completely to exclude air; acidify to preserve with hydrochloric acid to pH $<$ 2; maximum holding time $<$ 6 months.
   e. Responses should include details of containers, holding conditions, collection of blanks and holding times. Detailed responses can be found in the worked solutions in your digital formats.
   f. Responses should include details of containers, holding conditions, collections and holding times. Detailed responses can be found in the worked solutions in your digital formats.

15.3 Chemical contaminants

15.3 Exercise

1. Responses should include details of industrial and agricultural practices. Detailed responses can be found in the worked solutions in your digital formats.
2. a. Chemical contaminants are unwanted chemicals that make water unsuitable for its intended use.
   b. Responses should include details of a situation such as sodium ions in bore water. Detailed responses can be found in the worked solutions in your digital formats.
3. POPs are contaminants that are difficult to eradicate from the environment. Two examples of POPs are DDT and PCBs.
4. Substances that undergo biomagnification are persistent in the environment, fat-soluble, and difficult to eliminate because they are stored in the fatty tissue of the animals. This results in concentration as the substance moves up the food chain.
5. If POPs are water soluble, they can be transported away from the source of the contamination as a dissolved contaminant in water. If they are not water soluble, they may be adsorbed onto particles (from soil), which can then be washed away in rivers.
6. a. Lead is from dissolution from natural sources or household plumbing containing lead (e.g. pipes, solder).
   b. Nickel is from prolonged contact of water with nickel-plated fittings.
   c. From natural stores in water or from blockages or corrosion of pipes.

15.4 Review

15.4 Exercise

15.4 Exercise 1: Multiple choice questions

1. B.
2. D.
3. B.
4. A.
5. C.

15.4 Exercise 2: Short answer questions

1. Of all the water on Earth, only 2.6 per cent is fresh water.
2. a. A biodegradable substance is able to break down via natural processes over the medium term.
   b. A non-biodegradable substance cannot break down over the short to medium term.
3. a. $C_{12}O_2Cl_4$
   b. Dioxins are formed as unintentional by-products of many industrial processes involving chlorine. This includes waste incineration, chemical and pesticide manufacturing, and pulp and paper bleaching.
   c. Dioxins have a low polarity.
   d. Dioxins accumulate due to their chemical stability and their ability to be absorbed by fat tissue.
4. Responses should include details of a contaminant, the impacts and treatment. Sample responses are available in your digital formats. Detailed responses can be found in the worked solutions in your digital formats.
5. There are five typical water treatment steps: coagulation, flocculation, sedimentation, filtration and disinfection.

15.4 Exercise 3: Exam practice questions

1. Responses should include details of containers, holding conditions, collection of blanks and holding times. Detailed responses can be found in the worked solutions in your digital formats.
2. Responses should include details of the laboratory protocols not properly followed. Detailed responses can be found in the worked solutions in your digital formats.

3. Responses should include details of the contaminant chosen. Detailed responses can be found in the worked solutions in your digital formats.

16 Measurement of solubility and concentration

16.2 Measuring solubility and the use of solubility curves

Practice problem 1

a. See figure at foot of the page.*
b. From the graph, at 35 °C, solubility is around 36.4 g/100 g water.
c. 53.7 g

Practice problem 2

Step 1 Find the mass of KCl in 0.700 mols.

$m = 52.5$ g

Step 2 At SLC 25 °C, the solubility of KCl is around 35 g/100 g. Therefore, a sample of 52.2 g will form a supersaturated solution.

Step 3 Obtain an unsaturated solution:

  a. Increase temperature of the solution to a minimum of 80 °C, or
  b. Continue working at SLC, but dilute the solution. At 25 °C, solubility is 35 g/100 g; we have 52.2 g, thus we need:

$$52.2 \times \frac{100}{35} = 149.14 \text{ g water or } 150 \text{ g water}$$

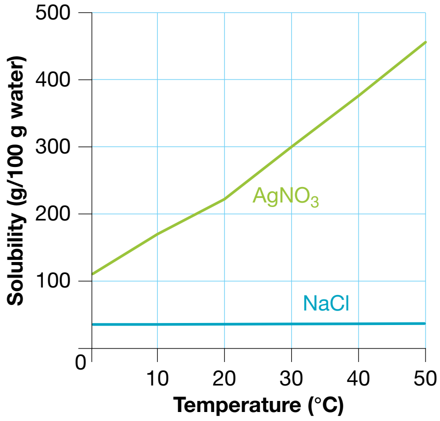
Practice problem 3

a. Both substances are fully dissolved.
b. All the potassium sulfate is still dissolved, and no crystals form.

   Predict that 14 g of pure potassium chloride crystals are formed.
c. Below 20 °C, potassium sulfate crystals would then start to form as well, and the potassium chloride crystals would no longer be pure.

16.2 Exercise

1. 12 g
2. 24 g
3. Sodium sulfate
4.



5. 30 °C.
6. 240 g water
7. 60 °C
8. 0.23 mols
9. a. Unsaturated solution
   b. 390 g
   c. 119 g
   d. Supersaturated
10. a. The solubility of gases decreases as the temperature increases.
    b. 0.25 g
    c. The solubility of CO_2 decreases as the temperature increases.

16.3 Solution concentration

Practice problem 4

0.020 g

Practice problem 5

4.25 L

Practice problem 6

1.5×10^{-4} g $L^{-1} = 1.5 \times 10^{-1}$ ppm

Practice problem 7

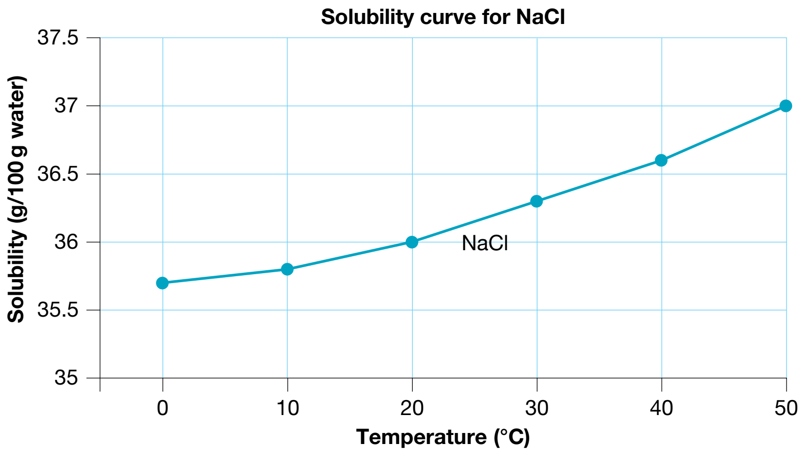
300 ppb

Practice problem 8

49% (m/m)

---

a*



Practice problem 9

0.4% (m/v)

Practice problem 10

The label is incorrect; it states a higher concentration, 32%(v/v), than found in the sample.

Practice problem 11

0.8 mol L^{-1}

Practice problem 12

30.3 g

Practice problem 13

562 mL

Practice problem 14

0.017 M

16.3 Solution concentration

16.3 Exercise

1. a. 1.70×10^4 mg L^{-1} b. 0.480 g L^{-1}
2. 0.0950 g
3. 4.00 mL
4. 3.75 μg g^{-1}
5. 0.352 g
6. 0.250 g
7. a. 25.6% (m/m) b. 0.50% (m/m)
   c. 0.0720% (m/m)
8. 1.25 g or 1 g (to 1 significant figure)
9. 2.6×10^2 g of solution.
10. 49%(m/m).
11. a. 20%(m/v) b. 0.0050%(m/v)
12. 5.6 g.
13. a. 20 mL. b. 0.10 mL.
14. 2.20%(v/v).
15. a. 37.5 mL b. 500 mL
16. 86.3 mL of alcohol in 750 mL of champagne.
17. a. 0.50 mol L^{-1} b. 4.29 mol L^{-1} c. 1.6 mol L^{-1}
18. a. 2.0 mol L^{-1} b. $c(\text{Na}^+) = 2.0\,\text{mol L}^{-1}$
   $c(\text{OH}^-) = 2.0\,\text{mol L}^{-1}$
19. 0.015 mol L^{-1}
20. 0.417 mol L^{-1}

16.4 Review

16.4 Exercises

16.4 Exercise 1: Multiple choice questions

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

16.4 Exercise 2: Short answer questions

1. a. 4×10^{-4} g L^{-1}.
   b. 0.4 ppm.
2. a. i. 48 °C ii. 58 °C iii. 24 °C
   b. 65 °C
   c. i. 36 g/100 g ii. 99 g/100 g iii. 55 g/100 g
3. An unsaturated solution contains less than the maximum mass of solute for a given volume of solution at a given temperature. A saturated solution contains the maximum mass of solute for a given volume of solution at a given temperature. A supersaturated solution contains more than the maximum mass of solute for a given volume of solution at a given temperature.
4. a. 11%(m/m) = 1×10^1%(m/m)
   b. 1.02%(m/m).
5. 3.4 g
6. 9.4 g
7. 2.9%(v/v).
8. a. 400 mL of ethanol b. 140 mL
9. 50 g of salt.
10. a. 18 g b. 49 g
    c. 2.8×10^3 g d. 1.3×10^2 g
11. a. 0.66 M b. 3.22 M c. 6.7 M
    d. 61.0 M e. 1.81 M f. 2.34 M
12. a. 0.50 M b. 0.340 M c. 0.127 M
13. a. 0.0030 mol b. 0.0045 mol
    c. 0.0092 mol d. 0.5 mol
14. a. 39 g b. 105 g c. 128 g
    d. 13.9 g e. 1.38 g f. 12.3 g
15. 16.4 g
16. a. 0.298 M b. 0.15 M c. 0.266 M d. 0.667 M
17. a. 1.3 M
    b. $c(\text{NH}_4^+)$= 1.3 M and $c\left(\text{NO}_3^-\right) = 1.3$M
18. a. 2.14 M
    b. $c(\text{Al}^{3+}) = 2.14$ M and $c(\text{Cl}^-)$
    = 6.42 M
19. 0.025 M
20. 0.083 M

16.4 Exercise 3: Exam practice questions

1. a. V, W
   b. T, X, Z
   c. U, Y
2. a. i. 3×10^1 g L^{-1} ii. 3×10^4 mg L^{-1}
   b. i. 0.005 g L^{-1} ii. 5 mg L^{-1}
   c. i. 2.50 g L^{-1} ii. 2.50×10^3 g L^{-1}
3. a. Between 0.10 and 2.0 ppb
   b. 1.0×10^9 L
   c. The energy cost required to extract 1 kg of gold from 1.0 billion litres of sea water would be well in excess of the value of the gold.

17 Analysis for salts in water

17.2 Chemical analysis

17.2 Exercise

1. a. Qualitative
   b. Qualitative
   c. Quantitative
   d. Qualitative

  - **e.** Both qualitative (colour change) and quantitative (pH values)
  - **f.** Qualitative
  - **g.** Qualitative
  - **h.** Qualitative
- **2. a.** It would be necessary to qualitatively establish the presence of a possible source of the oil waste before carrying out a more involved and expensive quantitative analysis.
  - **b.** The polluting oil would have to be analysed to determine what compounds were present in the mixture and in what quantities. The results of this analysis could then be compared with the analysis of any residual waste oil found on a ship.

17.3 The sources of salts in water

17.3 Exercise

1. A salt is a chemical substance which may be obtained from the reaction between an acid and a base. Other reactions produce salts. Salt is also sometimes used to specifically refer to sodium chloride or table salt.
2. Sodium chloride is the most common salt found in water.
3. Sample answer: Sodium, potassium, calcium, magnesium, chloride, sulfate, nitrate, iron or aluminium.
4. Despite $CaCO_3$ being almost insoluble in water, rainwater saturated with CO_2 is slightly acidic and forms calcium hydrogen carbonate, $Ca(HCO_3)_2$, which readily disassociates into its ions.
5. Heavy metals in drinking water can be from inappropriate discharges from industry, from corroded pipes and from pollution.
6. Both mercury and lead are toxic to humans and animals. They accumulate in the body, which can then lead to defects in children or other health problems in adults. They can also accumulate in food, such as fish grown in contaminated water, which can then become toxic when ingested by humans.
7. Methyl mercury is an example of an organo-metallic substance which is formed when a heavy metal combines with an organic substance.
8. Salts are ionic compounds; when they dissolve in water, they dissociate to produce mobile ions. This means that they can conduct electricity to an extent that depends on how many ions are present, but it will not distinguish between different types of ions.
9. Electrical conductivity is measured in EC units or $\frac{\mu S}{cm}$.
10. Na_2CO_3 will measure a higher electrical conductivity. Both solutions have the same concentration, but Na_2CO_3 has more ions than $NaNO_3$.

17.4 Analytical methods

Practice problem 1

Moles of water = 1.33

Practice problem 2

$m(PbI_2) = 3.976$ g

Practice problem 3

$m(Al_2O_3)= 9.44$ g

Practice problem 4

$c(H_2SO_4) = 600$ M

Practice problem 5

Percentage of water in the sample = 5.0%

Practice problem 6

%As in the pesticide = 9.95%
Note that a different treatment of significant figures (where you keep all information in your calculator until the final answer is achieved before rounding to the correct number of significant figures) can produce a small variation in your answer. For example: %As in the pesticide = 9.99%

Practice problem 7

Concentration of zinc = 1.252 g L^{-1}.

Practice problem 8

Absorbance = 0.24 nm.

Practice problem 9

Absorbance = 0.23 nm.

Practice problem 10

Concentration of the undiluted samples is = 97.2 ppm.

17.4 Exercise

1. a. 0.5 mol **b.** 0.5 mol **c.** 3.5 mol
2. a. 0.5 mol **b.** 0.2 mol **c.** 0.1 mol **d.** 0.2 mol
e. 0.05 mol **f.** 0.50 mol **g.** 8 mol
3. 6.3 g
4. a. 2.9 g **b.** 6.2 g **c.** 66 g
d. 1.97 g **e.** 15 kg
5. a. The barium carbonate is a solid.
b. 4.837 g
6. a. 6.04 g (3 significant figures)
b. All of the Na_2S will react.
7. a. All the HCl had reacted — the reaction was complete.
b. 0.909 M
c. Sources of error in this experiment would be in weighing the strip of magnesium and in measurement of the volume of hydrochloric acid.
d. Safety precautions include: Wear a safety coat, safety glasses and gloves, and work in a well-ventilated fume cupboard or a well-ventilated laboratory.
8. 96.7%

17.5 Exercise

1. Response should explain the chemical, economic, social and ethical factors and advantages and disadvantages. Detailed responses can be found in the worked solutions in your digital formats.

17.6 Review

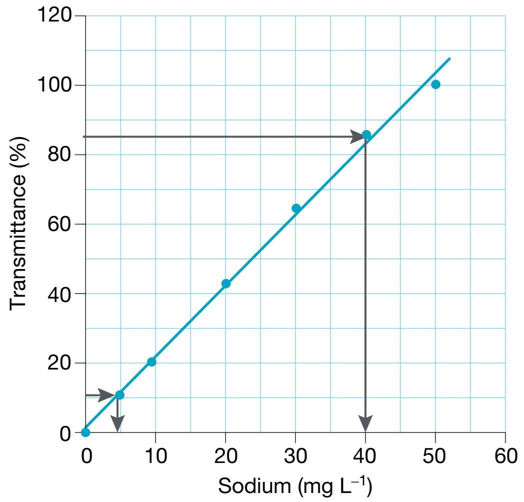
17.6 Exercises

17.6 Exercise 1: Multiple choice questions

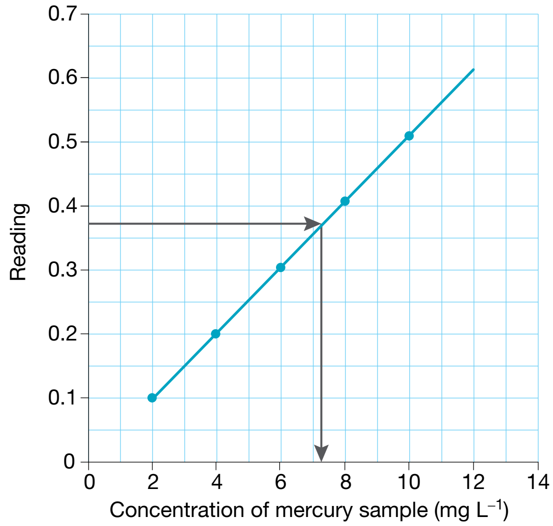
1. D.
2. B.
3. D.
4. D.
5. C.
6. D.
7. D.
8. A.
9. B.
10. D.

17.6 Exercise 2: Short answer questions

1. **a.** Qualitative **b.** Quantitative
   c. Quantitative **d.** Quantitative
2. **a.** Carbon dioxide is identified by the limewater test.
   b. Starch turns iodine dark blue.
   c. Oxygen gas is identified by the glowing splint test.
   d. Protein in a food sample turns a mixture of $CuSO_4$ and NaOH purple.
   e. Hydrogen gas ‘pops’ when a small sample is ignited.
   f. Water turns blue litmus paper pink.
3. **a.** $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow 2NaNO_3(aq) + BaSO_4(s)$
   Precipitate is barium sulfate.
   b. 5.972 g
4. **a.** 0.15 mol **b.** 13 g
5. **a.** 9.49 g **b.** 13.1 g
6. 0.029 g
7. **a.** 159 g **b.** 5.2 g
8. **a.** 0.921 kg
   b. $2NaOH(s) + CO_2(g) \rightarrow Na_2CO_3(aq) + H_2O(l)$
   c. 550 g
   d. When planning space travel, experts focus on minimising the mass of material that needs to be lifted into space. Lithium hydroxide would be preferred because a lower mass of lithium hydroxide would be needed to remove the same mass of carbon dioxide.
   e. $2OH^-(s) + CO_2(g) \rightarrow CO_3^{2-}(aq) + H_2O(l)$ for both reactions mentioned.
9. **a.** 46.8 tonnes
   b. The term ‘burning’ usually means the combustion of a substance with the oxygen in air.
10. 0.794 g
11. **a.** Ions dissolved in water are required for it to conduct electricity. The more ions from dissolved salts, the greater the conductivity. As salts form ions when dissolved in water, electrical conductivity can be used to measure salt levels in river water.
    b. Before the early 1980s, the trend was increasing. Since the early 1980s, the trend has been decreasing.
    c. These results represent total ion content, ions from salts other than NaCl are probably present also.
    d. Conductivity is related to the total amount of all ions present. It does not depend on which ions are present and individual types of ions cannot be identified through electrical conductivity measurements.
12. **a.**



    b. Concentration of sample 1 ≈ 41 mg L^{-1}.
    Concentration of sample 2 ≈ 5 mg L^{-1}.
13. **a.**



    b. The level of mercury is ≈ 7.2 mg L^{-1}.
    c. AAS
14. **a.** 13 ppm. **b.** 1.3% **c.** Orange
15. **a.** Lead **b.** Silver **c.** Barium **d.** Hydroxide
16. Precipitates are chosen so that their solubilities are very low. Keeping the volume of water used during the experiment, including that used to wash the precipitate, to a minimum is one way of reducing the chance that some of the precipitate may dissolve. It is possible, however, that some of the precipitate may dissolve and, hence, the percentage determined in any experiment may be a little lower than it should be. This would be a systematic error.
17. C, E, H, A, F, D

17.6 Exercise 3: Exam practice questions

1. **a.** Carbon dioxide
   b. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
   c. 83.2%
   d. Response should state one assumption. Detailed responses can be found in the worked solutions in your digital formats.
2. **a.** 1.14 g
   b. 79.7%

- **c.** Response should state one assumption. Detailed responses can be found in the worked solutions in your digital formats.

3. a. $NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
- **b.** 1.82 g
- **c.** 0.182 gL^{-1}
- **d.** It is far less likely that any of the AgCl precipitate will dissolve; i.e. it is far more likely that an accurate result will be obtained.
- **e.** Response should state one assumption. Detailed responses can be found in the worked solutions in your digital formats.

18 Analysis for organic compounds in water

18.2 Organic compounds in the environment

18.2 Exercise

1. Sample answer: fuels, plastics, polymers
2. The solubility of an organic substance is influenced by whether the substance is polar or not.
3. **a.** A broad class of carbon-based compounds
   - **b.** Carbon and hydrogen
   - **c.** Nitrogen, oxygen and possibly phosphorous and sulfur
4. **a.** They have to exist in the environment long enough to come into contact with the target insect pest.
   - **b.** The metabolite of methoxychlor is slightly more polar than DDT or methoxychlor.
5. $C_3H_8NPO_5$
6. Soluble organic compounds can dissolve directly into water while insoluble organic compounds can be adsorbed into soil, which can then be deposited in watercourses.
7. **a.** Sample answer: Oil spills may occur at sea if an oil tanker runs aground, or due to a malfunction or accident on an offshore oil rig.
   - **b.** Sample answer: An oil spill may occur on land due to a road accident involving a truck carrying oil, a factory accident involving the rupture of an oil storage tank or an accident at an oil refinery.

18.3 High-performance liquid chromatography (HPLC)

Practice problem 1

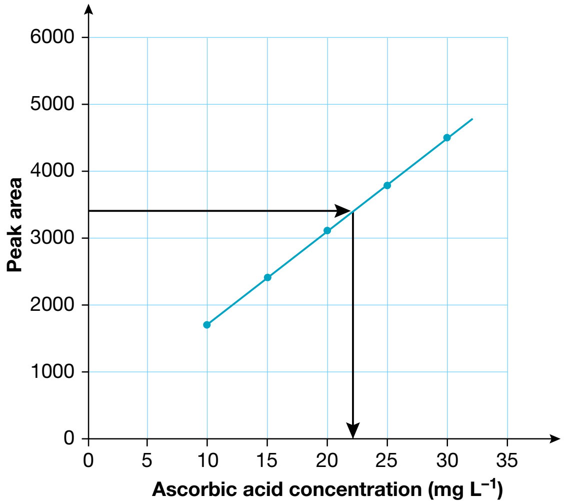
The peak areas 15, 24, 28 are all indicators of additional chemicals used in the preparation of that oil.

Practice problem 2

Expected peak height would be 40.5.

18.3 Exercise

1. **a.** The stationary phase provides a large surface area onto which components of a mixture can adsorb.
   - **b.** The mobile phase containing the mixture being analysed passes over the stationary phase and breaks any interactions that have been established by the components of the mixture and the stationary phase.
   - **c.** A graph with the peaks corresponding to the concentration of the substance being analysed.
   - **d.** Retention time is the time taken for a component of a mixture to pass through the column.
2. A calibration curve is a trend line — a line of best fit.
3. **a.**



   - **b.** ≈ 22 mg L^{-1}
   - **c.** ≈ 220 mg L^{-1}

18.4 Review

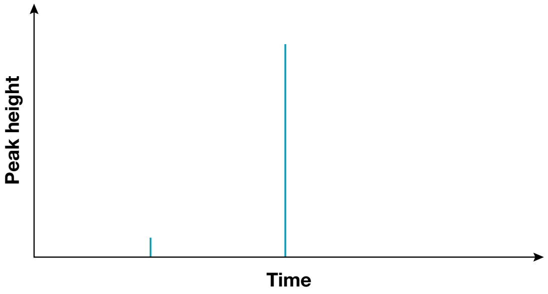
18.4 Exercises

18.4 Exercise 1: Multiple choice questions

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

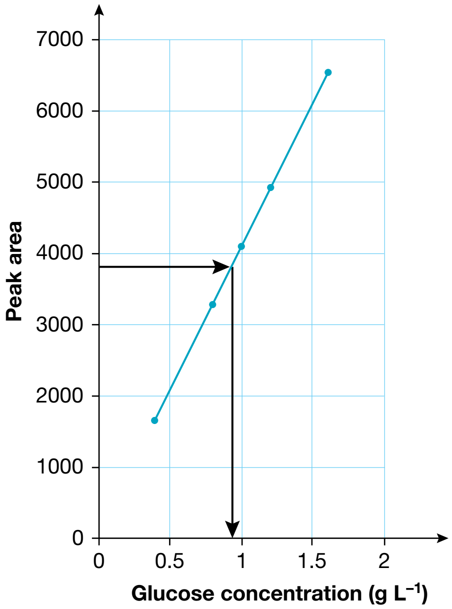
18.4 Exercise 2: Short answer questions

1. The solubility of organic chemicals in water is determined by the polarity of their molecules.
2. It is not possible to decide the solubility of an organic chemical based on the elemental make-up alone.
3. **a.** Blue, pink, yellow, black
   - **b.** The different components would be further separated from each other.
4. **a.**



   - **b.** There will only be one peak if there is only one compound.
     This would be the situation if the compound was pure.
5. 4.2 mL

6. a.



b. ≈ 0.9 g L^{-1}

c. 5 mM.

d. Yes

7. a. ≈ 16 mg L^{-1}.

b. ≈ 0.22 mg.

c. Yes

d. 2.2 mg/100 g; this is slightly lower than the claimed concentration.

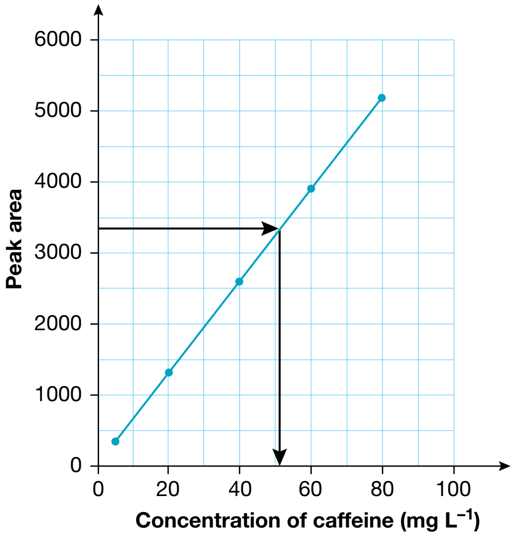
8. a. The light crude oil is highly volatile. Burning may be employed.

b. Because heavy crude oil is not as volatile as light crude, the best response may be containment with floating surface barriers and then application of dispersants or accelerated biodegradation.

c. A large spill occurring at sea is more difficult to contain because the waves will disperse the oil. One clean up method is to use dispersants that break the oil into droplets that then can be broken down by biodegradation.

18.4 Exercise 3: Exam practice questions

1. a.



b. 52 mg L^{-1}

c. ≈ 5200 mg L^{-1}

d. 52 mg in the tablet

2. a. Diazinon is relatively soluble and has a high strength of adherence to particles in the soil.

b. Oxamyl persists in the environment for a relatively short period of time after being applied.

c. The permethrin is likely to move away from the field because it is likely that substantial amounts of soil to which the permethrin has adhered would be washed away.

d. A few problems would arise because little insecticide would be moved due to its relatively low solubility and it strong adherence to the soil, which would stay in place.

19 Analysis for acids and bases in water

19.2 Sources of acids and bases in water

19.2 Exercise

1. Sample answers: volcanic eruptions, local geology, and decomposition of plants and animals.
2. Sample answers: burning fossil fuels, waste from mine tailing sites, industries that utilise acid such as galvanising, leather preparation, detergent, pharmaceutical and herbicide manufacture.
3. For pipes to clog, basic conditions lead to the deposition of scale.
4. A buffer is a substance dissolved in water that resists to changes of pH when acids or bases are added.
5. Anywhere between 6 and 8
6. $CO_2\left(g\right) + H_2O\,(1) \rightarrow H_2CO_3\,(aq)$

   $H_2CO_3\,(aq) + H_2O\,(1) \rightleftharpoons H_3O^+\,(aq) + HCO_3^-\,(aq)$
7. A measure of how a body of water withstands acid pollution.

19.3 Solution stoichiometry (volume–volume stoichiometry)

Practice problem 1

100 mL

Practice problem 2

100 mol L^{-1}

19.3 Exercise

1. a. $Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$

   b. 0.177 mol L^{-1}
2. 30.0 mL
3. 60.0 mL
4. 90.0 mL
5. 0.560 mol L^{-1}
6. 45 mL
7. 1.6 mol L^{-1}
8. 0.12 mol L^{-1}
9. 0.075 L or 75 mL
10. 1.18 mol L^{-1}

19.4 Acid–base titrations

Practice problem 3

13.25 g

Practice problem 4

0.0499 mol L^{-1}

Practice problem 5

0.340 mol L^{-1}

Practice problem 6

6.67 mol L^{-1}

19.4 Exercise

1. Responses should include all steps for making a standard solution. Detailed responses can be found in the worked solutions in your digital formats.
2. 0.095 01 mol L^{-1}
3. a. Standard solutions have accurately known concentrations.
   b. 0.055 13 mol L^{-1}
4. a. $2HNO_3(aq) + Na_2CO_3(aq) \rightarrow 2NaNO_3(aq) + CO_2(g) + H_2O(l)$
   b. 0.099 50 mol L^{-1}
5. The equivalence point is where the reactants have reacted and reachtline any trends and patterns in your ded their correct mole ratio. The end point in a titration is when the indicator changes colour.
6. Methyl orange, thymol blue, bromphenol blue, congo red or bromcresol green.
7. a. Litmus or bromothymol blue
   b. Litmus or bromothymol blue
   c. Phenolphthalein
   d. Methyl orange
8. 0.015 L
9. a. 0.0018 mol L^{-1}
   b. 64 ppm
10. 75 mg mol
    As 1 L was the volume analysed, the total alkalinity was 75 mg L^{-1}

19.5 Review

19.5 Exercises

19.5 Exercise 1: Multiple choice questions

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

19.5 Exercise 2: Short answer questions

1. a. Rainwater is made acidic when it reacts with the dissolved carbon dioxide in the atmosphere, make the weak acid, carbonic acid (H_2CO_3). Rainwater's acidity is enhanced by natural acids or acids that are generated by human activity dissolve in it.
   b. For the pH to change, rainwater must come into contact with salts and organic materials that are found in the soil.
   c. Acid rain.
   d. The burning of fossil fuels containing sulfur releases sulfur dioxide into the atmosphere. This, in turn, can be further oxidised to sulfur trioxide, which reacts with rainwater to form sulfuric acid.
2. a. Metallic sulfides such as pyrite (FeS_2)
   b. Calcium hydroxide, limestone, potassium hydroxide and potash
3. a. 21 mL b. 33 mL c. 3.1 mL
4. a. 200 mL b. 325 mL
   c. 1.60×10^3 mL (1.60 L)
5. 100 mL
6. a. 0.184 mol L^{-1} b. 0.105 mol L^{-1}
7. a. $NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
   b. 90 mL
8. 4.0 mol L^{-1}
9. 528 mL
10. 149 mL
11. 16.2 mL
12. a. 0.00410 mol L^{-1} b. 0.164 g L^{-1}

19.5 Exercise 3: Exam practice questions

1. a. 2.65 g b. 0.106 mol L^{-1}
2. a. $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
   b. 0.001 85 mol
   c. 0.000 926 mol
   d. 0.0370 mol L^{-1}
   e. 0.370 mol L^{-1}
   f. 0.0102 mol L^{-1}
   g. 35.3 L

20 Practical investigation

20.2 Key science skills in chemistry

Practice problem 1

To write a research question, the independent variable and dependent variable first should be determined. The research question should be testable and outline the investigation. The aim should outline the purpose of the investigation and the hypothesis needs to link the independent and dependent variables. Detailed responses can be found in the worked solutions in your digital formats.

20.2 Exercise

1. In order to be a testable question, the question must link the independent and dependent variable, be able to be investigated through scientific method and be practicable. Detailed responses can be found in the worked solutions in your digital formats.
2. A logbook is an important supplementary tool to gain an understanding to all aspects of an investigation.
   Detailed responses can be found in the worked solutions in your digital formats.
3. The dependent variable is measured by the investigator and is influenced by the independent variable. The independent variable is manipulated and changed by the investigator.
4. To make sure that the dependent variables measured are only influenced by the independent variable and not unfairly influenced by other factors.

Detailed responses can be found in the worked solutions in your digital formats.

5. a. Two independent variables: the amount in mol of the gas and the temperature.
   Dependent variable: volume of the gas.
   b. Controlled variables: the type of gas used, the temperature of the gas and the measuring equipment used.
6. a. The concentration of fertiliser.
   b. The mass of the shoots of the plants.
   c. Some examples of controlled variables may include type of seed, type of soil/potting mix, amount of soil/potting mix, size of the pots, temperature and humidity the pots are exposed to, type of fertiliser solution, volume of fertiliser solution, and time of day the fertiliser solution was added to each pot.
   d. The pot that received 0 g L^{-1} of fertiliser was the control in this experiment, because it was not exposed to the independent variable.
7. See table at foot of the page.*
8. B.
9. B, A, E, F, C, D.
10. A hypothesis must be testable by observation or experimentation in order to validate or invalidate it.
    Detailed responses can be found in the worked solutions in your digital formats.
11. B.

20.3 Concepts specific to an investigation

Practice problem 2

a. 282 000 is 2.82×10^5 km.
b. 1.235×10^6 m
c. 3×10^{-18} (or 3.0×10^{-18})

20.3 Exercise

1. Key terms can be defined either within a report (included as part of a sentence) or in a glossary at the end, with the key words bolded throughout and defined later.
2. Representations are useful in scientific reporting because they allow us to show complex and abstract concepts in an easy to understand form, enhancing the clarity in scientific communication.
3. a. 8×10^{-6} (or 8.0×10^{-6})
   b. 1.3×10^{-3}
   c. 3.844×10^8
4. a. 0.000721
   b. 573 900
   c. 15.52
   d. 7.112
   e. 471
   f. 0.000 002 577
   g. 0.000 091 79

20.4 Scientific research methodologies and techniques

Practice problem 3

a. Student 1
b. Student 1
c. Most precise data was Student 2. Most accurate data was Student 2.

20.4 Exercise

1. Reliability is linked with the repeatability of an investigation and whether or not someone could follow the investigation's protocol and get similar results. It relies on results being accurate and precise and can be affected by errors. Validity, however, is more about the credibility of data and whether or not it shows what it was intended to. This is often more affected by bias and variables not being correctly controlled.
2. Yes — an experiment needs to be replicated in order to be seen as reliable.
3. Results would not be considered reliable if repetitions of an investigation lead to vastly different answers.
   Detailed responses can be found in the worked solutions in your digital formats.
4. Sample answer: when variables have been controlled, the results address the hypothesis and the purpose of the investigation and bias has not been used to manipulate or alter results.
5. Sample answer:
   - not correctly testing the variable outlined in the hypothesis
   - misinterpreting measurements
   - not controlling variables correctly
   - testing more variables than those outlined in the hypothesis.
6. A valid experiment that accurately addresses the variables and controls factors must also be reliable, because others would get the same results. A method can be reliable (get the same results every time) but not address the hypothesis and the question, meaning it is not valid.
   Detailed responses can be found in the worked solutions in your digital formats.
7. Precision is a measure of how close various measurements are to each other. Accuracy is how close to the expected and known quantity a result is.
   Detailed responses can be found in the worked solutions in your digital formats.

7*

Topic	**Independent variables**		**Dependent variables**
	Numerical	**Categorical**	
Conductivity of	Voltage exposed to, size of metal, thickness of metal	Type of metal, structure of metal used (i.e. wire, sheet, block)	Resistance, voltage drop, current, brightness of a light globe
Concentration of salt in water	Amount of salt added, amount of water added	Cloudiness of solution, location water obtained (i.e. ocean versus tap)	Concentration of salt, electrical conductivity, refraction to light, reaction to silver nitrate
Boiling point of different soft drinks	Amount of sugar in soft drink, amount of soft drink	Type of soft drink, brand of soft drink	Temperature to boil, length of time taken to completely evaporate

8. A strength of quantitative data is it provides a clearer understanding of trends and patterns in results. A weakness is that it can be harder and more time-consuming to collect. A strength of qualitative data is that it is often easy and quick to collect. A weakness is that it can be subjective (for example, one person might identify a colour as blue, and another might state it is purple).

20.5 Ethics, health and safety guidelines

20.5 Exercise

1. a. Sample answer: 2.0 mol L^{-1} hydrochloric acid is a mid-range concentration, so can cause more of an impact than those of a lower concentration. Higher concentrations can irritate and burn eyes, lungs and skin. Appropriate safety precautions include the use of gloves, safety glasses and lab coats when using hydrochloric acid.
   b. Sample answer: Burettes are very fragile and prone to breakage, leading to cuts to students. Appropriate safety precautions include inspecting burettes for damage and to sweep up any broken glass immediately.
   c. Sample answer: Boiling water causes burns to exposed skin. Appropriate safety precautions include using equipment such as Hot Hand (hand protectors) around boiling water, allowing containers that have held boiling water to completely cool before touching and packing away, placing skin exposed to boiling water immediately under cold running water for at least 15 minutes and seeking medical assistance.
   d. Thermometers are made of glass and can break, leading to cuts. Some thermometers contain mercury, which is toxic. Appropriate precautions are to discard if damaged using a dustpan and brush and use carefully because they are fragile. If a mercury thermometer breaks, immediately cover with sulfur.
2. Sample answer: Ethics may be important when confidentiality or personal beliefs are involved
   Detailed responses can be found in the worked solutions in your digital formats.
3. Sample answer:
   - identifying hazards with equipment or chemicals
   - suggesting standard handling procedures
   - outlining the correct disposal of chemicals
   - outlining any first aid that may be required
   - providing information for the practical, including the location, time and date.
4. Sample answer: chemicals produced in an experiment (not just those used), further information about allergies, contacts for emergency hotlines, and clear information about how many students will be in each group and at each bench.
5. Answers will vary between schools and different laboratories.
   Detailed responses can be found in the worked solutions in your digital formats.
6. Detailed responses can be found in the worked solutions in your digital formats.
7. Sample answer: Students should wear safety glasses gloves and lab coats to avoid exposure to skin. The chemical should be kept away from any ignition sources and used in either a fume hood or a well-ventilated area.

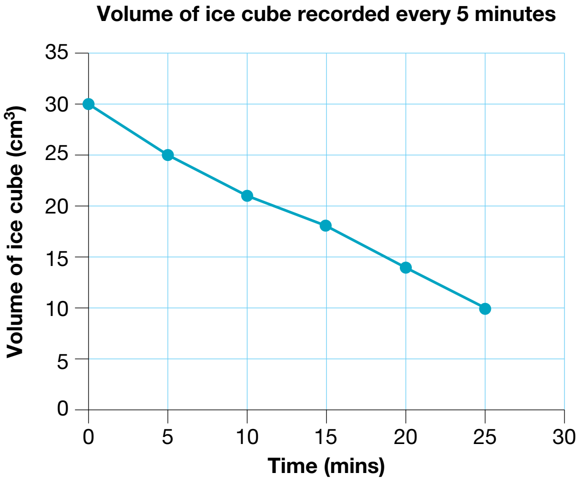
20.6 Methods of organising, analysing and evaluating primary data

Practice problem 4

a. 0.12 ± 0.005 g.
b. 0.195 ± 0.0005 g.

20.6 Exercise

1. Sample answer: Random error is a chance variation in measurements, and usually affects the precision of data. Systematic error usually affects the accuracy of an experiment, and is often due to equipment errors.
   Detailed responses can be found in the worked solutions in your digital formats.
2. Taking multiple results and averaging them out or by using more precise equipment.
   Detailed responses can be found in the worked solutions in your digital formats.
3. a. A bar graph would be most appropriate, because one variable is qualitative (household liquid) and the other is quantitative (pH).
   b. A line graph would be most appropriate because both variables (pH and temperature) are quantitative and a line graph would allow for trends to be seen. A scatterplot with a line of best fit may also be appropriate.
   c. A line graph would be most appropriate because a variable (temperature) is being observed over time.
   d. A histogram would be most appropriate because frequency and intervals are being used.
4. The following line graph would be an appropriate choice.



The graph shows that, as time passes, the volume of the ice cube decreases. On average, the ice cube decreases by around 4 cm^3 a minute, except between 10 and 15 minutes, where the rate of size reduction is slower, and between 0 to 5 minutes, when the rate of size reduction is slightly faster. This may be due to minor changes in room temperature, a different surface temperature or the changing surface area of the ice cube.
5. An outlier should be mentioned and discussed, with an outline of possible reasons it occurred. However, it usually is excluded when calculating averages or trying to add a line of best fit.
6. 11.42 mL or 11.43 mL

7. A burette reading is taken from the top down. This reading should be 7.32 mL.

8. a. 9.6±0.25 cm b. 8.5±0.125 cm c. 11.9±0.05 cm

9. a. Measuring cylinder
   b. Volumetric flask
   c. Conical flask
   d. Conical flask or a beaker
   Detailed responses can be found in the worked solutions in your digital formats.

20.7 Models and theories to understand observed phenomena

20.7 Exercise

1. Models can be useful because they:
   - allow us to visualise objects and concepts we can't see (for example, atoms)
   - make abstract concepts tangible
   - provide explanatory frameworks for a variety of chemical concepts, allowing for more in-depth concepts to be explored.
2. Some limitations of models include:
   - an oversimplification of complex situations, because models cannot show every detail
   - a misrepresentation of the concept in the real world can occur when conditions are not as carefully controlled
   - the ever-changing nature of models and research means that models aren't definite.
3. A model is a representation of a phenomena, whereas a theory is a well-supported explanation of a phenomena.
4. Theories allow us to understand phenomena because they are based on investigations and observations, providing us with a structured idea that can be applied to different concepts.
   Detailed responses can be found in the worked solutions in your digital formats.

20.8 Nature of evidence and key findings of investigations

20.8 Exercise

1. Hypotheses are supported based on the evidence we have at the time and how this is interpreted. However, as new knowledge becomes available the evidence we obtain changes. Therefore, we cannot definitively prove a hypothesis, but rather support it with what evidence we have at that point in time.
2. Sample answer: Quantitative data that was obtained when investigating a single variable, data that was not influenced by bias and data that has been obtained through a reliable method and repeated and obtained by other investigators.
3. Key findings communicated in your discussion should include information about trends and patterns in your data, the relationship to chemical concepts and theory, and a clear answer to the question of your investigation.
4. To show the information is being used to show and support chemical theory and phenomena. This gives an investigation more validity.

20.9 Options and issues related to water quality

20.9 Exercise

1. Water is vital in many aspects of life. It is important to test water to make sure it is free of contaminants and to make sure that any other chemicals present do not exceed dangerous amounts.
2. a. Mercury, lead and cadmium
   b. Bacteria such as E. coli and V. cholerae, algae, protozoa and phytoplankton
   c. Insecticides and pesticides, oil, dioxins and PCBs
3. The technique chosen would create a significant difference — not just for cost and time, but also for the concentrations of salt being detected. Some techniques only determine metal ions, whereas others are also able to detect anions. Some techniques, such as UV-visible spectrometry, are very effective, but can only be used on coloured solutions or those that absorb light in the ultraviolet region.

20.10 Conventions of scientific report writing and scientific poster presentation

Practice problem 5

a. 0.7823 kg
b. 2.13×10^8 pL

Practice problem 6

0.463 g mL^{-1}.

20.10 Exercise

1. a. Introduction: summarise key background concepts, explain any other research that has been done in this field, outline key terms, explain the purpose of your investigation and outline the hypothesis being explored.
   b. Discussion: outline any trends and patterns in your data, describe any outliers and how they were treated, evaluate your results and link them to theory, discuss any errors, uncertainties and limitations, describe the precision, accuracy, reliability and validity of your investigation. and suggest improvements and future investigations.
   c. Conclusion: sum up your investigation, and provide a clear answer to your question and a link back to your aim and your hypothesis.
2. a. 1.984 L b. 1231 m c. 0.053153 s
3. a. 123 000 mg b. 167.285 mmol
4. a. 6 b. 3
   c. 1 d. 7
5. a. 13 b. 3.7
   c. 76.075 d. 0.5
   e. 213 f. 18.7
6. a. 23 g mL^{-1}. b. 28.78 g.
   c. 3.2 mol. d. 35.6 s.
7. *In-text:* (Taylor, Stubbs & Stokes, 2020)
   Reference list: Taylor, N, Stubbs A, Stokes, R. (2020). *Jacaranda Chemistry 1 VCE Units 1 & 2.* 2nd edition. Milton: John Wiley & Sons.

20.11 Review

20.11 Exercises

20.11 Exercise 1: Multiple choice questions

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

20.11 Exercise 2: Short answer questions

1. **a.** Sample answer: Recorded pH and the concentration of H^+ ions
   b. Sample answer: A digital pH reader, universal indicator, potentiometric titrations or an acid–base titration
   c. Sample answer: If the pool is acidic or basic, or the colour an indicator turns on exposure to the pool.
   d. Sample answer: Indicators (such as litmus, phenolphthalein or bromophenol blue)
2. **a.** Independent variables are manipulated by the investigator. Dependent variables are measured and are the result of changes to the independent variable.
   b. Qualitative data is categorical data whereas quantitative data is numerical.
   c. A control group does not have the independent variable applied to it and is a point of comparison. Experimental groups have been impacted by the independent variable.
   d. Primary sources of data are from the initial source of data collection, often through direct investigation. Secondary data is a summary of analysed primary data.
   e. Uncertainty is around the limit to the precision of a piece of equipment. Error involves factors affecting the accuracy of the experiment and how close results are to their true value.
   Detailed responses can be found in the worked solutions in your digital formats.
3. **a.** 3 significant figures
   b. 5.4×10^{-2} mol L^{-1}
4. **a.** 12.5%; interpolation0.4 The nature of evidence an
   b. 14.1%; extrapolation
   c. It would be difficult to give a reliable estimate of the ethanol level.
   d. Sample answer: As the ethanol concentration increases, the peak area increases in a linear pattern when using gas chromatography. A calibration curve allows for the ethanol concentration of an alcoholic beverage to be determined based on the peak area measured in the readings.
   Detailed responses can be found in the worked solutions in your digital formats.
5. **a.** The graph should have a clear heading, scale and labelled axis.
   Detailed responses can be found in the worked solutions in your digital formats.
   b. Sample answer: As the concentration of cordial increases, the time it takes to freeze also increases. When no cordial is present, it takes 25 minutes for the cordial to freeze, but as this increases to 50%, it takes 65 minutes to freeze. While there is a positive correlation, this is not completely linear, with the rate levelling off as more cordial is added.
   c. As the concentration of cordial is increased, the time taken to freeze also increases.

20.11 Exercise 3: Exam practice questions

1. **a.** Sample answer: Safety measures have been undertaken (safety gloves and glasses) but are not sufficient. There is no mention if she tied her hair back or used metal tongs or Hot Hands when handling the test tube at 50 °C.
   b. The mass of sodium sulfate that dissolved in the water.
   c. Sample answer:
   Any two of the following:
   - amount of water
   - source of water
   - size of test tube
   - same batch of sodium sulfate
   - same thermometer used to record temperature
   - same conditions (for example, if one test tube is stirred or mixed, the other test tube should also be stirred or mixed in the same way).

   d. Sample answer: Measuring cylinder, scales and a thermometer (to determine the temperature).
   Some factors that may affect the accuracy of the equipment are described in the sample response available in the worked solutions in your digital formats.
   e. Sample responses can be found in thew worked solutions in your digital formats.
2. **a.** Sample answer: To examine if different brands of bottled water have different pH.
   b. The hypothesis must clearly link the independent variable and dependent variable, and must be tentative and testable. Detailed responses can be found in the worked solutions in your digital formats.
   c. Sample answer:
   Qualitative data: the colour the indicator turns or if the substance is acidic or basic.
   Quantitative data: pH, the concentration of H^+ ions or the number of bubbles.
   d. Sample answer:
   Factors that may lead to different results by students include:
   - the interpretation of the colour of universal indicator
   - the technique used to determine pH (i.e. litmus paper provides less information than universal indicator)
   - the equipment used (e.g. if glassware was contaminated)
   - the sample of water (students may all have tested the same brand of water, but perhaps one student uses a bottle that was opened a week ago, another uses a cold fresh bottle, and the third uses a closed bottle that is a year old).

   e. Detailed responses can be found in the worked solutions in your digital formats.

3. a. See figure at foot of the page.*

   b. Some trends shown in the data include:
      - a negative correlation/downwards trend, in which pH decreases as temperature increases
      - a drop of only 0.2 in pH between 0 and 10 °C and 10 and 20 °C, but a greater drop of 0.3 in pH between 20 and 30 °C, 30 and 40 °C and 40 and 50 °C
      - though the decline is quite small, the pH drops by 1.3 over 50 °C.

   c. You could not accurately determine the pH at 90 °C. While the graph shows a linear relationship, 90 °C is quite outside the data range.

4. a. Sample answer:
      - Faulty thermometer (systematic error): make sure all thermometers have been calibrated or test the temperature using three different thermometers and get an average.
      - Temperature being incorrectly recorded (random error): repeat the experiment multiple times and take an average or get another student to check the measurement.

      Detailed responses can be found in the worked solutions in your digital formats.

   b. 97 °C (to two significant figures)

   c. The known boiling point of water in 100 °C. Accuracy requires data to be closest to the known value. Alan's data had a maximum deviation of 2.5 °C from the actual value, whereas Will had a maximum deviation of 4.5 °C. Therefore, Alan is more accurate than Will.

   d. 0.000987 g L^{-1} (9.87×10^{-4} g L^{-1})

5. a. The student hasn't shown the dependent variable in their hypothesis, so it is not testable.

      Detailed responses can be found in the worked solutions in your digital formats.

   b. The experiment being conducted over three days leads to a loss of some controlled variables. This means that it is not a fair test, and other factors could have resulted in the change of temperature of the water.

   c. 0.25 °C. This means that a level of uncertainty exists with the data, and the results obtained by each individual are not as accurate as they could be.

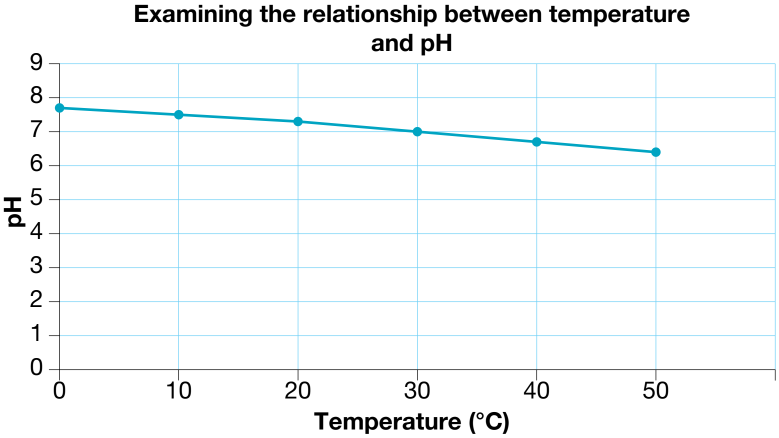
   d. Sample answers:
      - the inability to exactly interpret boiling point (because it is subjective)
      - the inability to completely control room temperature
      - the lack of accuracy in the recorded results
      - the source of the boiling water not being consistent because it was done over multiple days
      - the lack of repetition in the method
      - the starting temperature of the water not being consistent.

   e. Line graph with the three types of variables superimposed on the same graph.

   f. Sample conclusion:

      It was found that aluminium provides the greatest insulation for boiling water, leading to the slowest decrease in temperature (a 25 °C drop over 25 minutes). This is followed by wool (55.5 °C drop over 25 minutes) and finally by paper (58.5 °C drop over 25 minutes). The hypothesis in this investigation was supported, because aluminium worked best in insulating water.

3a*



GLOSSARY

%(m/m) percentage mass per mass

%(m/v) percentage mass per volume

%(v/v) percentage volume per volume

accuracy how close an experimental measurement is to a known value

acid substance that can donate a proton to a base

acidic oxide oxide of a non-metal

acid rain rain made acidic by mixing with atmospheric pollution resulting a pH less than 5

addition polymerisation the process in which monomers with at least one double bond react together to form a polymer by addition reactions

addition reaction a reaction in which one molecule bonds covalently with another molecule without losing any other atoms

adsorbent porous substance with a high surface area that adsorbs analytes by intermolecular forces

adsorption process by which molecules adhere, or ‘stick’, to the surface of other particles; commonly involves liquid or gas molecules sticking to the surface of solid particles

aim statement outlining the purpose of an investigation, linking the dependent and independent variables

alcohol compound containing at least one hydroxy (—OH) group

aliquot volume delivered by a pipette

alkali soluble base

alkalinity a measure of the ability of water to withstand sudden pH change caused by acid input

alkane a simple, saturated binary compound of carbon and hydrogen atoms with single bonds between the carbon atoms; the general formula for all members of this homologous series is C_nH_{2n+2}

alkene a hydrocarbon that contains one carbon–carbon double bond

alkyne a hydrocarbon that contains one carbon–carbon triple bond

allotrope different forms of an element

alloy a substance with metallic properties that consists of two or more elements, at least one of which is a metal

alloying mixing of a metal with one or more other elements (usually metals)

amphiprotic substance that can accept or donate protons

anhydrous without water in the chemical structure

anion a negatively charged atom or group of atoms

annealed metal that has been heated in an inert gas to a moderate temperature, its temperature maintained for several hours, and then cooled slowly

aqueous solution with water as the solvent

aqueous solutions mixtures in which substances are dissolved in water

atom a neutral particle with a nucleus; the smallest sample of an element

atomic absorption spectroscopy analytical method for detecting metal species

atomic emission spectrum a spectrum emitted as distinct bands of light of diagnostic frequencies by elements or compounds

atomic number the number of protons in the nucleus of an atom of a particular element

average titre titre determined by averaging the results of several titrations

Avogadro’s constant, N_A the number of elementary particles (atoms) in exactly 12 g of carbon-12. This number is equal to 6.02×10^{23}.

ball-and-stick model representation of a molecule in which the atoms are shown as balls and the bonds as sticks

ball bearing model representation of a metal using ball bearings to show the arrangement of cations into grains

bar graph graph in which data is represented by a series of bars; usually used when one variable is quantitative and the other is qualitative

base substance that can accept a proton from an acid

bias the intentional or unintentional influence on a research investigation

binary describes compounds made up of only two elements

biodegradable able to break down via natural processes over the medium term

biomagnification process by which a compound (such as a pollutant or pesticide) increases its concentration in the tissues of organisms as it travels up the food chain

boiling temperature temperature at which a liquid boils and turns to vapour

bond dipole separation of charge in a polar covalent bond

bonding electrons the pairs of electrons involved in forming a covalent bond

buffers solution that resists a change in pH upon the addition of a small amount of acid or base

carbon nanotube structure in the shape of a tube composed of carbon atoms

carboxylic acid compound containing at least one carboxy (—COOH) group

catalytic cracking the use of a catalyst to break bonds in long molecules to produce shorter molecules

categorical data data that has labels or names rather than a range of numerical quantities; also known as qualitative data

cathodic protection method of protecting a metal from corrosion by connecting it to a power source and another metal that is less active. The electrons feed onto the metal being protected, making it a cathode so that it cannot be oxidised.

cation a positively charged atom or group of atoms

causation when one factor or variables directly influences the results of another factor or variable

central atom the atom in a molecule with the most bonding electrons

chemical bond arrangement of electrons between two atoms that generates a force, causing the atoms to be bound to each other

chromatogram a visible record (commonly a graph) displaying the results of HPLC, showing the detector signal against time

computer-generated model representation of a molecule that is produced by a computer

concentrated has a large quantity of solute in a known volume of solution

concentration the amount of solute that is dissolved in a known volume of solvent

conclusion section at the end of a scientific report that relates back to the question, sums up key findings and states whether the hypothesis was supported or rejected

concordant individual titres within a certain limit

concordant results results of repeated titrations that are identical (within ± 0.10 mL of each other) and consecutive

condensation polymerisation polymerisation in which two monomers combine and a smaller molecule is eliminated

condensed formula formula derived from the structural formula that shows the atoms present in a structure carbon by carbon and is written on a single line; also called a semi-structural formula

conjugate acid the product that a base forms when it has accepted a proton from an acid

conjugate base the product that an acid forms when it has donated a proton to a base

conjugate redox pair two species that differ only by a certain number of electrons

contaminant the presence of unwanted chemicals, or chemicals present in a higher concentration than wanted that makes water unsuitable for its intended use

control a group that is not affected by the independent variables and is used as a baseline for comparison

control group group that is not affected by the independent variables, and is used as a baseline for comparison

controlled variable variable that is kept constant across different experimental groups

copolymer polymer formed from the polymerisation of two monomers

correlation measure of a relationship between two or more variables

corrosion oxidation of metals by materials in their environment

covalent bonding bonding between non-metal atoms that involves electron sharing

covalent layer lattice a substance that displays covalent bonding in three dimensions without the formation of discrete molecules

covalent molecular compound a molecular compound in which atoms of different elements share electrons with each other

covalent molecular element element made up of identical atoms held together by covalent bonds

cross-linking the bonding between two polymers

crude oil naturally occurring mixture of compounds that are mainly hydrocarbons; also known as petroleum

crystalline lattice structure of particles that are held together in an ordered, three-dimensional arrangement

cyclic compound compound containing a ring structure, such as benzene

delocalised describes electrons that are not bound to any one atom but are free to move throughout a lattice

density physical property found by dividing the mass of an object by its volume

dependent variable variable that is influenced by the independent variable; the variable that is measured

desalination the process of removing salt and other unwanted substances from sea water and waste water to produce water for human use

desorption the opposite of adsorption, where molecules are released from the surface of other particles

diatomic molecule substance containing two atoms only

difunctional refers to a molecule that contains two functional groups

dilute has a small quantity of solute in a known volume of solution

dipole unequal sharing of electrons between atoms in a molecule

dipole–dipole interactions weak bonding caused by the positive end of one dipole attracting the negative end of another dipole

diprotic can donate two protons

discrete separate, distinct; not in an infinite array or lattice

discussion detailed area of a scientific report in which results are discussed, analysed and evaluated, relationships to concepts are made, errors, limitations and uncertainties are assessed and suggestions for future improvements are outlined

dispersion force the bond between adjacent molecules formed by instantaneous dipoles; this weak non-directional bonding is also known as van der Waals force

displacement reaction chemical reaction in which a more reactive element displaces a less reactive element from its compound

dissociation the process by which ions separate when an ionic compound dissolves in a solvent

distillation the process used to separate a mixture of substances due to differences in boiling point

double (reversible) arrows arrows used in an equation for a reaction where the forward reaction is occurring at the same time as the reverse reaction

double-blind type of trial in which neither the participant nor the researcher is aware of which participants are in a control or experimental group

double bond strong bond between two atoms formed by two pairs of electrons that are shared by the two nuclei

dry corrosion corrosion of a metal in the air when water is not a reactant

ductile able to be drawn out into a thin wire

duplicate sample repeat of the original sample, taken under the same conditions, at the same time and the with same sampling equipment, in order to estimate the sample variability

elastomer polymer that can be stretched and return to its original shape

electrochemical protection a method of protecting a metal from corrosion by connecting it to a power source and another metal that is less active. The electrons feed onto the metal being protected, making it a cathode so that it cannot be oxidised.

electrolyte a liquid that can conduct electricity

electron configuration the arrangement of electrons in the shells of an atom

electron dot diagram representation where the atom's nucleus and all innershell electrons are replaced by its element symbol and the outershell electrons are represented by dots around the symbol in a square arrangement

electronegativity the electron-attracting power of an atom

electrostatic force of attraction between positive and negative charges

electrostatic attraction strong force of attraction between positive and negative ions that produces an ionic bond

element chemical species consisting of atoms of a single type

eluent solvent used to remove analytes from the adsorbent in a HPLC column

elution in HPLC when each component reaches the end of the column and exits from it

empirical formula formula that shows the simplest numerical ratio in which atoms are combined

endothermic chemical process or reaction that absorbs energy

end point experimentally determined equivalence point at which the indicator just changes colour or the pH curve becomes vertical

equivalence point point at which two reactants have reacted in their correct mole proportions in a titration

error difference between a measurement taken and the true value that is expected; errors lead to a reduction in the accuracy of the investigation

ester organic compound formed from a condensation reaction between an alcohol and a carboxylic acid

ethics acceptable and moral conduct determining what is 'right' and 'wrong'

eutrophication excess nutrients in rivers and lakes that promote growth and decay of plant life to the detriment of other species

excess reactant reactant left over when the limiting reagent has been used up

excited state raised to a higher than ground-state energy level

exothermic chemical process or reaction that releases energy through light or heat

experimental group test group that is exposed to the independent variable

filtrate liquid that has passed through a filter

fractional distillation method of separating the components of a liquid mixture that depends on the ease of vaporisation of the components

full equation equation that includes all of the reactants and products that are present

fullerenes nanoscale allotropes of carbon, characterised by having some type of hollow shape

functional group group of atoms attached to or part of a hydrocarbon chain that influence the physical and chemical properties of the molecule

galvanised the coating of an iron-containing metal with zinc; the zinc forms a physical and chemical barrier to protect the iron because zinc is more reactive than iron

gravimetric analysis analysis by mass of precipitate

ground state the least excited energy level of an atom

groups the vertical columns of elements in the periodic table

half-equation chemical equation showing either oxidation or reduction in a redox reaction

haloalkane compound containing a halogen and an alkane

heavy metal metal with high density or of high relative atomic weight and that is a risk to health or environment

high-density polyethene (HDPE) thermoplastic polymer with very few side branches produced from the monomer ethene

histogram graph in which data is sorted in intervals and frequency is examined; used when both pieces are data are quantitative

homogeneous uniform composition throughout

homologous series series of organic compounds with the same structure but in which the formula of one molecule differs from the next by a $—CH_2$ group

human error mistakes made by individuals conducting an investigation, such as mistakes in entering data or the use of the wrong type of chemical in a reaction

hydrated containing water molecules within the structure

hydrates adds water

hydride binary compound containing hydrogen

hydrocarbon compound containing only carbon and hydrogen

hydrogen bonding the bond between a hydrogen atom covalently bonded to an atom of F, O or N and another molecule that also contains an atom of H, F, O or N

hydrolysis reaction involving the addition of water

hydronium ion H_3O^+, the product of a reaction of a proton with water

hypothesis tentative, testable and falsifiable statement for an observed phenomenon; acts as a prediction for the investigation

immiscible when two substances are unable to form a homologous solution

independent variable variable that is changed or manipulated by an investigator

indicator substances that undergo an observable change in an acidic or basic solution

insoluble will not dissolve

instrumental colorimeter instrument used to determine the concentration of coloured compounds in solution through comparison of a test sample against samples of known concentration through production of a calibrations curve (absorbance versus concentration)

intermolecular bond bond between molecules

interstitial alloy an alloy in which the smaller atoms fit into the spaces between the larger atoms

intramolecular bonding bond involving the internal structure of a molecule

ion a charged atom

ion–dipole attraction attraction between an ion and a polar molecule

ionic bond bond between cations and anions in an ionic compound

ionic compound compound containing cations and anions in an ionic lattice structure or an ionic liquid, depending on ion size

ionic concentrations concentrations of the anions and cations resulting from the dissolution of an ionic lattice; designated by square brackets

ionic equation equation without the spectator ions

ionisation reaction in which a substance reacts with water to form ions

ionisation energy the energy required to remove an electron from a gaseous atom or ion

isoprene volatile liquid hydrocarbon obtained from petroleum, the molecule of which forms the basic structural unit of natural and synthetic rubbers

isotopes forms of an element with the same number of protons but different numbers of neutrons in the nucleus

isotopic symbol representation of an element as ${}^{A}_{Z}E$, where E is the symbol for the element, A is the mass number and Z is the atomic number

latent heat energy needed to change state

latent heat of fusion, (L_f) energy needed to change a fixed amount from the solid to the liquid state

latent heat of vaporization, (L_v) energy needed to change a fixed mass from the liquid to the gaseous state

Lewis structure structure that shows how outer shell electrons are arranged in a molecule; also known as electron dot structure

like dissolves like strongly polar and ionic solutes tend to dissolve in polar solvents, and non-polar solutes tend to dissolve in non-polar solvents

limitations factors that have affected the interpretation and/or collection of findings in a practical investigation

limiting reactant the reactant that is completely used up in a reaction

linear polymer polymer that does not have side chains

line graph graph in which points of data are joined by a connecting line; used when both pieces of data are quantitative (numerical)

line of best fit trend line added to a scatterplot to best express the data shown; these are straight lines, and are not required to pass through all points

logarithm quantity representing the power to which a fixed number (the base) must be raised to produce a given number

low-density polyethene (LDPE) thermoplastic polymer with a large number of side branches produced from the monomer ethene

lustrous having a shiny surface

malleable able to be hammered or pressed into shape without breaking or cracking

mass number the total number of protons and neutrons in the nucleus of a particular isotope of an element

mass spectrometer analytical instrument that determines the relative isotopic masses of the different isotopes of an element and abundance

melting temperature temperature at which state changes from solid to liquid

metal ductile, malleable, conducting element

metallic bonding positively charged metal cations arranged in a lattice with delocalised valence electrons being able to flow around them

metallic lattice structure crystal lattice made up of an array of cations

miscible able to be mixed and form a solution

mobile phase the mixture to be analysed in the chromatography process

model representation of ideas, phenomena or scientific processes; can be physical models, mathematical models or conceptual models

Mohs scale a scale used to measure the hardness of substances

molarity concentration measured in units of moles of solute per litre (mol l^{-1} or M) of solution

molar mass mass of one mole of a compound expressed in grams per mol

mole 6.02×10^{23} particles of a substance

molecular group of atoms bonded together covalently

molecular formula actual number of atoms of each element in a molecule of a substance

molecule group of atoms bonded together covalently

mole ratio represents the relative number of moles of each substance involved in the reaction

monomer molecule that links together to form a polymer

monoprotic can donate one proton

multiple bond bond formed when two atoms share two or more pairs of electrons

multi-stage flash distillation an energy-efficient process to desalinate water that occurs in multiple stages of decreasing pressure allowing boiling to occur at successively lower temperatures

nanofiltration a filtration method using membranes with nanometre-sized pores

nanomaterial a material that has at least one dimension from 1 to 100 nanometres

nanometre one billionth of a metre, 0.000000001 m (10^{-9} m)

nanotechnology technology at a dimension of less than 100 nanometres

network lattice a substance that displays covalent bonding in three dimensions without the formation of discrete molecules

neutralisation chemical reaction in which an acid and a base react to form a salt

noble coating protection of a metal from corrosion by attaching a thin metallic layer to it

non-biodegradable not able to break down in a natural environment

non-bonding electrons electrons that are not involved in bonding

non-metal non-ductile, non-malleable, non-conducting element

non-polar molecules that have an even distribution of charge

non-polar covalent bond bond formed between atoms with the same electronegativity

non-polar molecule molecule that does not have permanent dipoles or is symmetrical

numerical data data that involves numbers and can be measured or counted; also known as quantitative data

octet rule a generalisation that works for many (but not all) atoms, stating that atoms will donate or share electrons in order to achieve eight electrons in their outer shells

orbital three-dimensional wave describing a bound electron

organic chemistry the study of carbon-containing compounds and their properties

osmosis diffusion of a solvent through a membrane towards an area of higher concentration of dissolved substances

outlier result that is a long way from other results and seen as unusual

oxidant an electron acceptor

oxidation an increase in the oxidation number; a loss of one or more electrons

oxidation number a theoretical number that corresponds to the charge that atom would have if the compound was composed of ions

oxidising agent the element that accepts electrons

parts per million (ppm) number of a particular component present within one million objects

peer-reviewed work that has been reviewed for quality of research, experimental reproducibility, accuracy and validity

periods the seven horizontal rows in the periodic table

pesticide a chemical substance used to kill organisms that are harmful to cultivated crops or animals

PET thermoplastic polymer produced from the monomer ethylene terephthalate; typically used as synthetic fibres (polyester) and packaging (PET); commonly recycled

photon particle of light

pH scale used to measure acidity; pH can be calculated using $\text{pH} = -\log_{10}[h_3o^+]$

placebo substance or treatment that is not designed to affect an individual (used as a control)

plagiarise to copy or use information from a source without referencing it and pass it off as your own

plastic polymer that can be moulded when hot and retains its shape when cooled

plasticisers small molecules that are added to some polymers to improve their flexibility

polar molecules that have an uneven distribution of electrons due to differences in electronegativity of the atoms

polarity localised imbalances in electric charges within a molecule resulting in a negatively charged end and positively charged end

polar bond bond formed when two atoms that have different electronegativities share electrons unequally

polar covalent bond bond formed when two atoms that have different electronegativities share electrons

polar molecule a molecule which, due to its polar bonds *and* its asymmetric shape, has an overall imbalance in the distribution of its electrons

polyatomic ion charged ion composed of two or more atoms

polyester synthetic polymer in which the structural units are linked by ester bonds

polyethene polymer consisting of monomers of ethene

polymer large molecule formed by the joining together of many smaller molecules

polymerisation the formation of giant molecules by repeated monomers that are joined by covalent bonds

polymers molecule made up of thousands of units strung together into very long chains

polyprotic able to donate more than one proton to a base

pop test test for hydrogen gas where a small amount of the gas is collected and a lit match is placed nearby

precipitate solid with low solubility formed in a liquid

precipitation reaction reaction where two solutions are mixed and a solid is formed

precision how close multiple measurements of the same investigation are to each other

primary data direct or firsthand evidence about some phenomenon

primary source direct or firsthand evidence about some phenomenon

primary standard substance used in volumetric analysis that is of such high purity and stability that it can be used to prepare a solution of accurately known concentration

principal quantum number, n quantum number that indexes the energy and size of an atomic orbital

product chemical species obtained as the result of a chemical reaction

qualitative analysis use of experimental procedures to determine which elements are in a substance

qualitative data categorical data that examines the quality of something (e.g. colour or gender) rather than numerical values

quantitative analysis se of experimental procedures to determine the percentage composition of a compound or the percentage of a component in a mixture

quantitative data numerical data that examines the quantity of something (e.g. length or time)

quenched metal that has been plunged while hot into water to rapidly cool it

random error chance variation in measurements

randomized making sure that the assigning of individuals to an experimental or a control group is random and is not influenced by external means

reactivity series of metals an ordered list of how readily the metals react with oxygen, water, steam, dilute acids and salt solutions

redox describes a reaction that involves the transfer of one or more electrons between chemical species

reducing agent the element that donates electrons

reductant electron donor

reduction a decrease in the oxidation number; a gain of electrons

refractive index a measure of how much a substance refracts (bends) light when it enters it

relative atomic mass, A_r mean mass of the naturally occurring mixture of the isotopes of an element on the relative atomic mass scale, on which the masses of particles are compared with the mass of the carbon-12 isotope

relative isotopic mass mass of an individual isotope of an element on the relative atomic mass scale, on which the masses of particles are compared with the mass of the carbon-12 isotope

reliability whether another researcher could repeat an investigation by following the outlined method and obtain similar results

representative sample small sample with similar characteristics that accurately reflects the larger entity

residue substance remaining in filter paper after the process of filtration

results section in a scientific report in which all data obtained is recorded, usually in the form of tables and graphs

retention time time taken for a substance to travel through the chromatography column; R_t

reticulated water water provided through a network of pipes

reverse osmosis movement of a solvent under hydrostatic pressure through a membrane towards an area of lower concentration of dissolved substances; commonly used to purify water

risk assessment document that examines the different hazards in an investigation and suggests safety precautions

rubber an elastic polymer formed from the latex of tropical rubber trees and plants

rusting wet corrosion of iron

sacrificial anode a reactive metal used to prevent a less active metal from corroding

sacrificial coating protection of a metal from corrosion by attaching another more active metal to it; the more active metal reacts preferentially with corrosive elements and protects the less reactive metal

sacrificial protection protection of a metal from corrosion by attaching another more active metal to it. The more active metal reacts preferentially with corrosive elements and protects the less reactive metal.

salt ionic compound consisting of a metal ion and a non-metal ion, except oxides and hydroxides

saturated hydrocarbon compound composed of carbon and hydrogen and containing only single covalent bonds

saturated solution solution in which the maximum amount of solvent has been dissolved

scatterplots graph in which two quantitative variables are plotted as a series of dots

scientific method procedure that must be followed in scientific investigations that consists of questioning, researching, predicting, observing, experimenting and analysing; sometimes referred to as scientific process

scientific research methodology the principles of research based on the scientific method

secondary data comments on or summaries and interpretations of primary data

secondary standard a standard prepared for a specific volumetric analysis that has been standardised against a primary standard

self-ionisation reaction in water where water ionises very slightly to form H_3O^+ and $OH^–$

self-ionisation constant K_w; in all aqueous solutions this value is constant and results from the self-ionisation of water:

$$H_2O\,(l) + H_2O\,(l) \rightleftharpoons H_3O^+\,(aq) + OH^-\,(aq)$$

$$K_w = [H_3O^+]\,[OH^-] = 10^{-14}\ m^2$$

simple distillation the process to separate two liquids of different boiling points

single bond bond formed by one pair of electrons shared between two atoms

skeletal formula structural representation where a molecule is represented by vertices; carbons and hydrogens are assumed to be at each vertex and at the ends unless otherwise indicated

solubility grams of a substance that will dissolve in an amount of solvent at a particular temperature

solubility curves graphs of grams of a substance that will dissolve in 100 g of solvent at a particular temperature

soluble will dissolve

solute dissolved substance in a solution

solutions solute dissolved in a solvent

solvent substance in which the solute is dissolved

space-filing model three-dimensional representation of a molecule that shows the relative sizes of atoms within the molecule and the distances between them

species general term that could refer to atoms, molecules, molecular fragments or ions

specific heat capacity energy (measured in joules) needed to raise the temperature of 1 g of a pure substance by 1°C

spectator ion ions that take no part in a reaction but maintain electrical neutrality in a solution

standard solutions solution that has a precisely known concentration

stationary phase a material with a high surface area, used in the chromatography process

stock solution concentrated solution that will be diluted as required for later use

stoichiometry the amount and ratio of reactants and products in a chemical reaction

strong acid acid that fully ionises in water

strong base base that fully dissociates in water

structural formula a diagrammatic representation of a molecule showing every bond

structural isomers compounds that have the same molecular formula but different arrangements for the atoms in that formula

subatomic particles electrons, protons and neutrons

substitutional alloy an alloy in which the atoms of the different elements are about the same size, so they can replace each other in the metal crystals

substitution reaction reaction in which an atom (or group of atoms) is removed and replaced by a different atom (or group of atoms)

supersaturated solution solution in which a greater amount of solute is dissolved at a particular temperature than is predicted by a solubility curve temperature

surfactant a molecule that has a polar (hydrophilic) and a non-polar (hydrophobic) end; for example, a detergent molecule

synergy to have an effect that is greater than the individual effects combined

synthetic fibre a polymer that is stretched to make a fibre instead of being moulded

systematic errors errors that affect the accuracy of a measurement that cannot be improved by repeating an experiment; usually due to equipment or system errors

tempered a quenched metal that is heated again but to a lower temperature than the initial heating

theory well-supported explanation of a phenomena, based on facts that have been obtained through investigations, research and observations

thermal cracking the use of heat to break bonds in long molecules to produce shorter molecules

thermoplastic describes polymers that soften on heating; also known as thermosoftening

thermosetting describes polymers that do not soften on heating and char if heated strongly

titration process used to determine the concentration of a substance by using a pipette to deliver one substance and a burette to deliver another substance until they have reacted in exactly their mole ratio

titration curve results of a titration showing the pH change versus the volume of the titre added during titration

titre volume of liquid measured by a burette during titration

triple bond strong bond between two atoms formed by three pairs of electrons that are shared by the two nuclei

triprotic acid that can donate three protons

uncertainty limit to the precision of equipment; a range within which a measurement lies

unified atomic mass units; u standard unit of mass based on the mass of the C-12 isotope, where mass of an atom of C-12 equals 12 u; it is approximately equal to one nucleon (proton or neutron)

unsaturated hydrocarbon compound containing carbon and hydrogen but possessing one or more double or triple bonds

unsaturated solution solution containing less than the maximum amount of solute that can dissolve

UV–visible spectroscopy technique used to study compounds that absorb light in the UV–visible region

valence electrons electrons in the outermost shell (highest energy level) of an atom; largely determine chemical properties of an element and contribute to chemical bond formation

valence electrons outershell electrons

valence shell electron pair repulsion (VSEPR) theory a model with the main point that the structure around a given atom in a molecule is determined principally by minimising electron repulsions

valence structure a diagrammatic representation of the outershell electrons in a molecule; similar to an electron dot diagram but replaces each pair of electrons with a dash (–)

validity shows how much results measure what the investigation was intending to, and how well they show the claims the investigation makes

volumetric analysis measurement of volume of one solution of accurately known concentration to determine the concentration of a second solution using stoichiometry

vulcanisation the process in which sulfur is added to rubber and heated to cause cross-linking of polymer chains, increasing the strength of the rubber

water of crystallisation water molecules included in a crystal lattice structure

weak acid acid that partially ionises in water

weak base base that partially dissociates in water

wet corrosion corrosion of a metal by reaction with oxygen in the presence of water

work hardened a property of metal brought about by beating the metal when cold

INDEX

Note: Figures and tables are indicated by italic *f* and *t*, respectively, following the page reference.

PERIODIC TABLE OF THE ELEMENTS

Key

Period 1	1 Hydrogen **H** **1.0**	2 Helium **He** **4.0**	Atomic number Name Symbol Relative atomic mass

	Alkali metals	**Alkaline earth metals**	**Transition metals**						
	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	Group 9
Period 1	1 Hydrogen **H** **1.0**								
Period 2	3 Lithium **Li** **6.9**	4 Beryllium **Be** **9.0**							
Period 3	11 Sodium **Na** **23.0**	12 Magnesium **Mg** **24.3**							
Period 4	19 Potassium **K** **39.1**	20 Calcium **Ca** **40.1**	21 Scandium **Sc** **45.0**	22 Titanium **Ti** **47.9**	23 Vanadium **V** **50.9**	24 Chromium **Cr** **52.0**	25 Manganese **Mn** **54.9**	26 Iron **Fe** **55.8**	27 Cobalt **Co** **58.9**
Period 5	37 Rubidium **Rb** **85.5**	38 Strontium **Sr** **87.6**	39 Yttrium **Y** **88.9**	40 Zirconium **Zr** **91.2**	41 Niobium **Nb** **92.9**	42 Molybdenum **Mo** **96.0**	43 Technetium **Tc** **(98)**	44 Ruthenium **Ru** **101.1**	45 Rhodium **Rh** **102.9**
Period 6	55 Caesium **Cs** **132.9**	56 Barium **Ba** **137.3**	57–71 Lanthanoids	72 Hafnium **Hf** **178.5**	73 Tantalum **Ta** **180.9**	74 Tungsten **W** **183.8**	75 Rhenium **Re** **186.2**	76 Osmium **Os** **190.2**	77 Iridium **Ir** **192.2**
Period 7	87 Francium **Fr** **(223)**	88 Radium **Ra** **(226)**	89–103 Actinoids	104 Rutherfordium **Rf** **(261)**	105 Dubnium **Db** **(262)**	106 Seaborgium **Sg** **(266)**	107 Bohrium **Bh** **(264)**	108 Hassium **Hs** **(267)**	109 Meitnerium **Mt** **(268)**

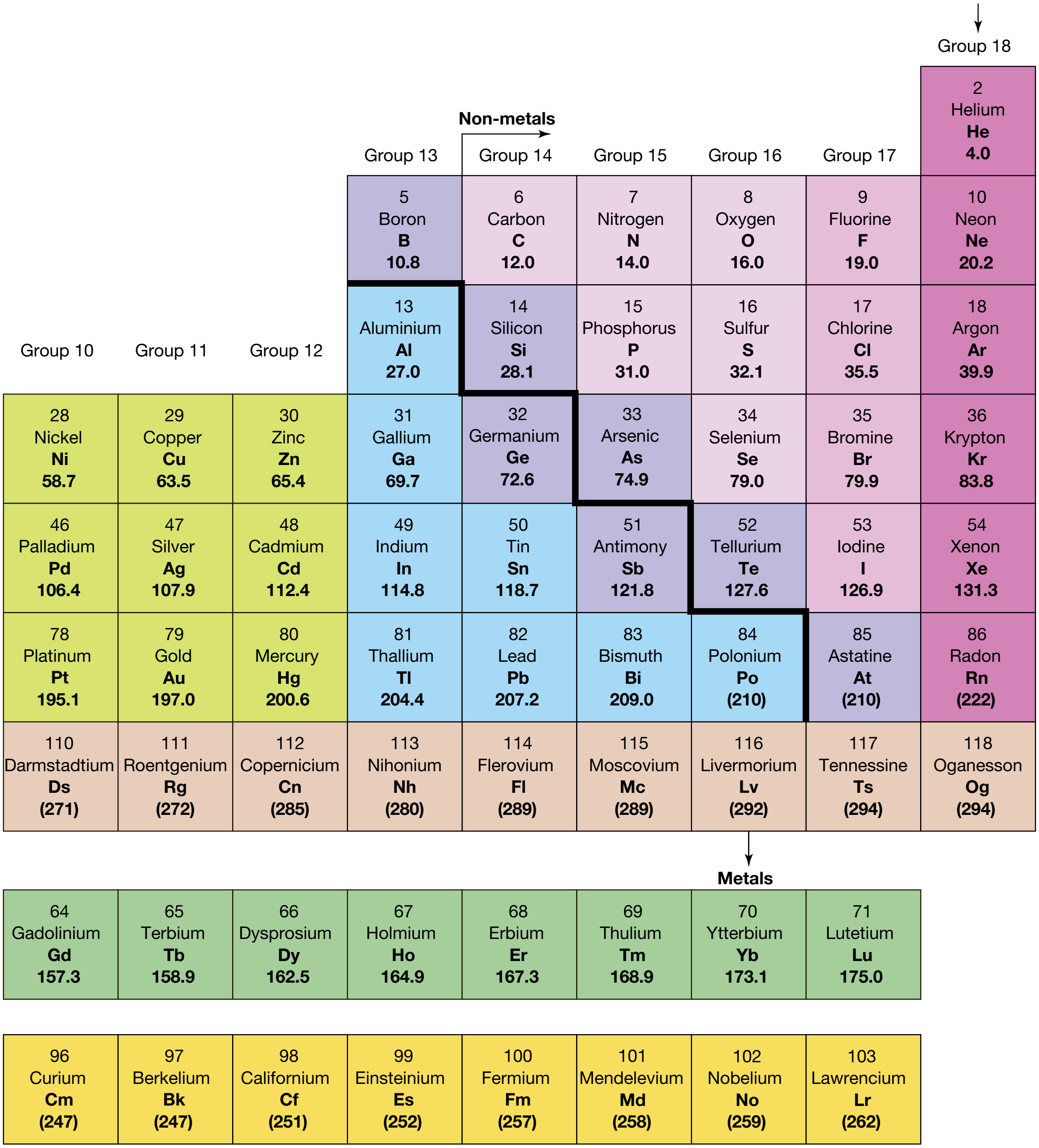
- 🟧 Alkali metal
- 🟨 Alkaline earth metal
- 🟩 Transition metal
- 🟩 Lathanoids
- 🟨 Actinoids
- 🟫 Unknown chemical properties
- 🟦 Post-transition metal
- 🟪 Metalloid
- 🟪 Reactive non-metal
- 🟪 Halide
- 🟪 Noble gas

Lanthanoids

57 Lanthanum **La** **138.9**	58 Cerium **Ce** **140.1**	59 Praseodymium **Pr** **140.9**	60 Neodymium **Nd** **144.2**	61 Promethium **Pm** **(145)**	62 Samarium **Sm** **150.4**	63 Europium **Eu** **152.0**

Actinoids

89 Actinium **Ac** **(227)**	90 Thorium **Th** **232.0**	91 Protactinium **Pa** **231.0**	92 Uranium **U** **238.0**	93 Neptunium **Np** **(237)**	94 Plutonium **Pu** **(244)**	95 Americium **Am** **(243)**



Group 10	Group 11	Group 12	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
								2 Helium **He** **4.0**
			5 Boron **B** **10.8**	6 Carbon **C** **12.0**	7 Nitrogen **N** **14.0**	8 Oxygen **O** **16.0**	9 Fluorine **F** **19.0**	10 Neon **Ne** **20.2**
			13 Aluminium **Al** **27.0**	14 Silicon **Si** **28.1**	15 Phosphorus **P** **31.0**	16 Sulfur **S** **32.1**	17 Chlorine **Cl** **35.5**	18 Argon **Ar** **39.9**
28 Nickel **Ni** **58.7**	29 Copper **Cu** **63.5**	30 Zinc **Zn** **65.4**	31 Gallium **Ga** **69.7**	32 Germanium **Ge** **72.6**	33 Arsenic **As** **74.9**	34 Selenium **Se** **79.0**	35 Bromine **Br** **79.9**	36 Krypton **Kr** **83.8**
46 Palladium **Pd** **106.4**	47 Silver **Ag** **107.9**	48 Cadmium **Cd** **112.4**	49 Indium **In** **114.8**	50 Tin **Sn** **118.7**	51 Antimony **Sb** **121.8**	52 Tellurium **Te** **127.6**	53 Iodine **I** **126.9**	54 Xenon **Xe** **131.3**
78 Platinum **Pt** **195.1**	79 Gold **Au** **197.0**	80 Mercury **Hg** **200.6**	81 Thallium **Tl** **204.4**	82 Lead **Pb** **207.2**	83 Bismuth **Bi** **209.0**	84 Polonium **Po** **(210)**	85 Astatine **At** **(210)**	86 Radon **Rn** **(222)**
110 Darmstadtium **Ds** **(271)**	111 Roentgenium **Rg** **(272)**	112 Copernicium **Cn** **(285)**	113 Nihonium **Nh** **(280)**	114 Flerovium **Fl** **(289)**	115 Moscovium **Mc** **(289)**	116 Livermorium **Lv** **(292)**	117 Tennessine **Ts** **(294)**	118 Oganesson **Og** **(294)**

64 Gadolinium **Gd** **157.3**	65 Terbium **Tb** **158.9**	66 Dysprosium **Dy** **162.5**	67 Holmium **Ho** **164.9**	68 Erbium **Er** **167.3**	69 Thulium **Tm** **168.9**	70 Ytterbium **Yb** **173.1**	71 Lutetium **Lu** **175.0**
96 Curium **Cm** **(247)**	97 Berkelium **Bk** **(247)**	98 Californium **Cf** **(251)**	99 Einsteinium **Es** **(252)**	100 Fermium **Fm** **(257)**	101 Mendelevium **Md** **(258)**	102 Nobelium **No** **(259)**	103 Lawrencium **Lr** **(262)**