

WORKBOOK

THIRD EDITION

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SACE STAGE 1

Australian Curriculum

Chemistry



Contents

ABOUT THIS BOOK	4
CHAPTER 1 TOPIC 1: MATERIALS AND THEIR ATOMS	5
1.1: Properties and uses of materials	6
1.2: Atomic structure	21
1.3: Quantities of atoms	40
1.4: The periodic table	51
Review Test 1	64
CHAPTER 2 TOPIC 2: COMBINATIONS OF ATOMS	71
2.1: Types of materials	72
2.2: Bonding between atoms	78
2.3: Quantities of molecules and ions	110
Review Test 2	115
CHAPTER 3 TOPIC 3: MOLECULES	121
3.1: Molecule polarity	122
3.2: Interactions between molecules	136
3.3: Hydrocarbons	145
3.4: Polymers	170
Review Test 3	181
CHAPTER 4 TOPIC 4: MIXTURES AND SOLUTIONS	187
4.1: Miscibility and solutions	188
4.2: Solutions of ionic substances	199
4.3: Quantities in reactions	208
4.4: Energy in reactions	232
Review Test 4	245
CHAPTER 5 TOPIC 5: ACIDS AND BASES	249
5.1: Acid-base concepts	250
5.2: Reactions of acids and bases	256
5.3: The pH scale	271
Review Test 5	283
CHAPTER 6 TOPIC 6: REDOX REACTIONS	289
6.1: Concepts of oxidation and reduction	290
6.2: Metal reactivity	307
6.3: Electrochemistry	316
Review Test 6	337
SOLUTIONS TO CHAPTER QUESTIONS AND REVIEW TESTS	343
PERIODIC TABLE OF ELEMENTS	409

About this book

The Third Edition of the popular SASTA Stage 1 Chemistry Workbook sets students on the path to success in their study of chemistry through its clear and engaging narrative, innovative use of art and photos, and superior contextual questions that enhance teaching and learning.

The Third Edition provides an unmatched comprehensive text fully mapped to the Stage 1 Chemistry Course in South Australia. The text is highly organised and emphasises essential chemistry concepts while keeping students engaged with learning outcomes that assess progress in understanding. The workbook has the concise, accessible, and engaging writing style of past editions while maintaining a clear emphasis on developing the reader's ability to apply their knowledge in new and familiar contexts rather than simply recalling it. Some major strengths of the text include its ability to develop critical thinking, problem-solving skills, comprehensive factual knowledge, and its aesthetically stunning artwork that represents complex topics clearly and succinctly.

The Third Edition provides hundreds of new Chapter Questions and six Review Tests that assess student learning in the key areas of Science Understanding, Science Inquiry Skills and Science as a Human Endeavour. In addition, the workbook includes a comprehensive set of solutions to all Chapter Questions and Review Tests. Each chapter is written with a contextual-based approach accompanied by frequent assessment provides the intellectual challenge needed to promote critical thinking and ensure academic success. The engaging and challenging questions have made this a leading textbook of choice for chemistry students.

SASTA is excited about the Third Edition of this quality workbook. We are confident that the clarity of the text in combination with its assessment tasks will support students in building the mental models needed to understand chemistry. In addition, the book is written and reviewed by experienced chemistry teachers who use their teaching knowledge as a guide in producing an up-to-date text that is beautifully illustrated and pedagogically sound, helping students to learn faster, study efficiently, and retain more knowledge of key concepts.

This book will be updated as the curriculum changes. SASTA and the author wish to thank its members who have written to us to identify errors and offer suggestions for improvements. Your feedback has been invaluable and is deeply appreciated. SASTA and the author continue to welcome constructive feedback about the content and design of this book. We also wish to thank the readers and schools who use our physics, chemistry and biology workbooks. Your support is greatly appreciated.

Both SASTA and the author wish you all the best with your studies in Stage 1 Chemistry, and we remind you that our Stage 2 Chemistry workbook is available for purchase if you wish to pursue Chemistry at Stage 2.

1.2: Atomic structure

All materials consist of atoms.

Atoms are commonly modelled as consisting of electrons orbiting a nucleus containing protons and neutrons.

Emission and absorption spectra of elements provide evidence that electrons are arranged in distinct energy levels and can be used to identify some elements in matter.

All materials, including elements, compounds, and mixtures, are composed of **atoms**. **Elements** are composed of only one type of atom, and **compounds** are composed of two or more chemical elements that combine in a fixed ratio.

Atomic Structure

Atoms have a central **nucleus** containing **protons** and **neutrons** surrounded by a cloud of one or more orbiting **electrons**. For example, **Figure 1.16** shows the positions of protons, neutrons and electrons in an atom with three protons, four neutrons and three electrons.

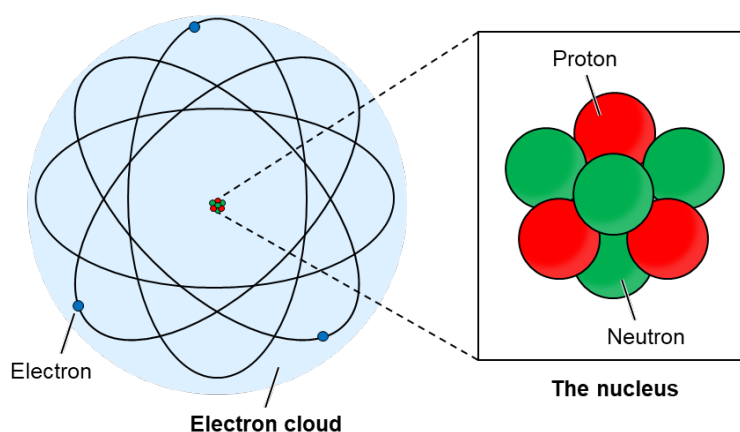


Figure 1.16: Atomic structure.

Protons, neutrons and electrons are defined as **subatomic particles** as they comprise the substructure of an atom. Some properties of the subatomic particles are provided in the table below.

Subatomic particle	Electric charge (e)*	Relative mass (u)*	Position in an atom
Proton	+1	1	Inside the nucleus
Neutron	0	1	Inside the nucleus
Electron	-1	Nearly 0	Outside the nucleus

Atoms are electrically neutral as each contains equal numbers of positively charged protons and negatively charged electrons. For example, sodium atoms contain 11 protons and 11 electrons.

* e is the elementary charge ($+1.6 \times 10^{-19}$ C)

u is an atomic mass unit equal to 1.66×10^{-27} kg.

Flame Tests

More than 1000 years ago, Arabic alchemists noted that the chemical elements emitted a specific colour of visible light when exposed to a flame. For example, **Figure 1.19** shows the light emitted by heated samples of lithium, sodium and potassium.



Figure 1.19: Flame tests for lithium (left), sodium (centre) and potassium (right).

Chemists used this property of matter to identify the chemical elements present in a material. However, as the years progressed and more elements were discovered, this test became less useful as some elements produced similar colours when heated. For example, **Figure 1.20** shows the green light emitted by copper (Cu) and barium (Ba) and the red light emitted by lithium (Li) and strontium (Sr).

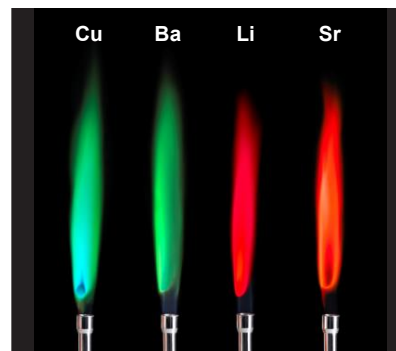


Figure 1.20: Similar flame colours

In 1859, German chemist Robert Bunsen described the limitations of flame tests to identify chemical elements to his colleague Gustav Kirchhoff. Several days later, Kirchhoff developed a **spectroscope**, a device that uses a prism to separate the wavelengths of light emitted by heated samples into a **spectrum**. When analysing the light spectra emitted by heated elements, Bunsen and Kirchhoff noted a series of sharp, coloured lines. Furthermore, he noted that each chemical element produced a unique spectrum of coloured lines, called a **line emission spectrum**.

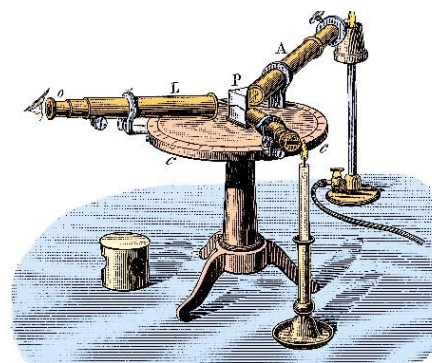


Figure 1.21: Kirchhoff spectroscope

Line emission spectra show the discrete wavelengths of visible light emitted by the atoms of different chemical elements. The emitted wavelengths measured in nanometres (nm) are unique to an element, and this property is used to identify a chemical element in a mixture.

Figure 1.22 shows the line emission spectra of hydrogen, helium and neon.

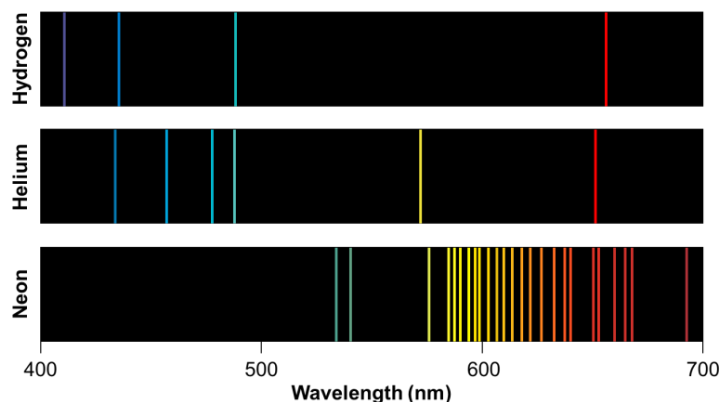


Figure 1.22: Line emission spectra

In contrast, an electron in a higher energy level transitions to a lower energy level by emitting light with energy equal to the difference between the initial and final energy levels (Figure 1.26).

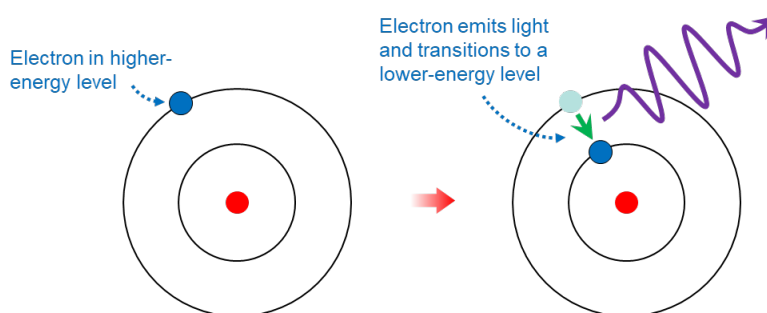


Figure 1.26: Emission of light by an atom.

Bohr realised that the coloured lines in the emission spectra of the

elements represented wavelengths of visible light emitted when electrons in atoms transition from higher to lower energy levels. Furthermore, Bohr determined the wavelength (λ) of a coloured line in the emission spectrum was related to the difference in energy (ΔE) between two energy levels in an atom. Figure 1.27 shows four electron transitions in a hydrogen atom from higher to lower energy levels. Each transition is associated with the emission of a wavelength of visible light that appears as a coloured line in hydrogen's line emission spectrum.

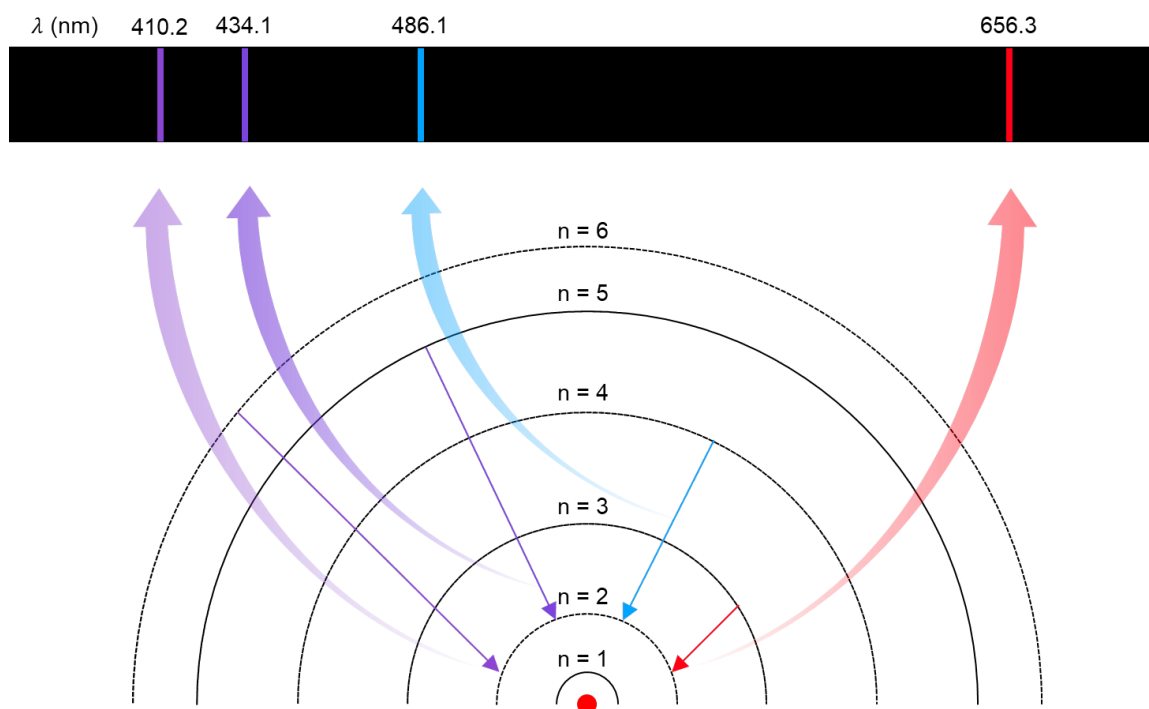


Figure 1.27: Electron transitions and the line emission spectrum of hydrogen.

Bohr's model accurately explained the coloured lines in hydrogen's line emission spectrum, and this discovery and his atomic model awarded him the 1922 Nobel Prize in Physics. Following the introduction of Bohr's atomic model, chemists began investigating the line emission spectra of other elements. It soon became apparent that an element's characteristic line emission spectrum was evidence that the atoms of different elements each have a unique set of energy levels not shared by atoms of other elements. This discovery allowed chemists to use spectroscopy to identify known elements in mixtures and to discover previously undetected elements.

Line Absorption Spectrum

When a continuous spectrum of visible light is transmitted through a gaseous or vaporised sample of an element or compound, discrete wavelengths of light are absorbed by atoms in the sample resulting in a series of dark lines in the continuous spectrum (Figure 1.28).

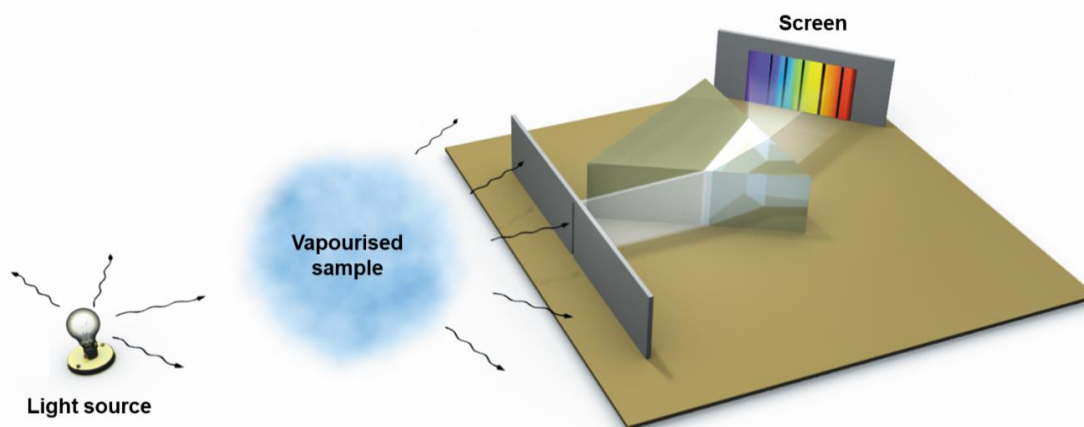


Figure 1.28: Continuous spectrum of visible light transmitted through a gas.

The dark lines are called **absorption lines**, and the resulting spectrum is called a **line absorption spectrum**. Each of the absorption lines corresponds to a wavelength of visible light that has been absorbed by the electrons of atoms in the gaseous sample as they transition from lower to higher energy levels.

Figure 1.29 shows the line absorption spectra of hydrogen, helium and neon.

Each element has a characteristic line emission and absorption spectrum due to the unique energy levels within its atoms. Furthermore, the position of the coloured lines in the emission spectrum of an element matches the positions of the dark lines in its absorption spectrum.

(Figure 1.30). This property enables chemists to identify a chemical element in a compound or mixture by comparing its absorption lines to the emission lines of the known elements. This technique is used to identify chemical elements in mineral, soil and water samples and the atmospheres of stars and planets.

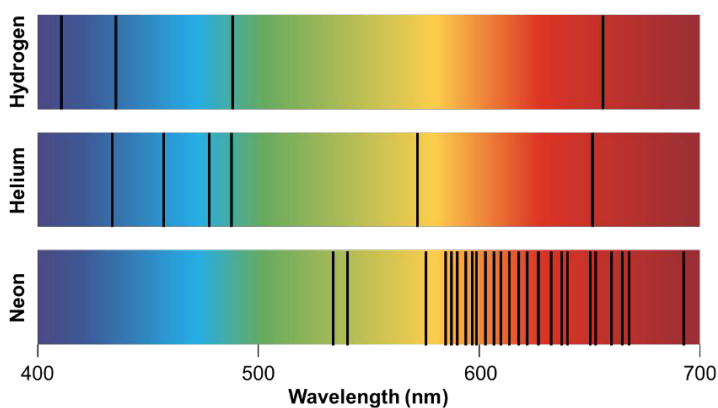


Figure 1.29: Line absorption spectra.

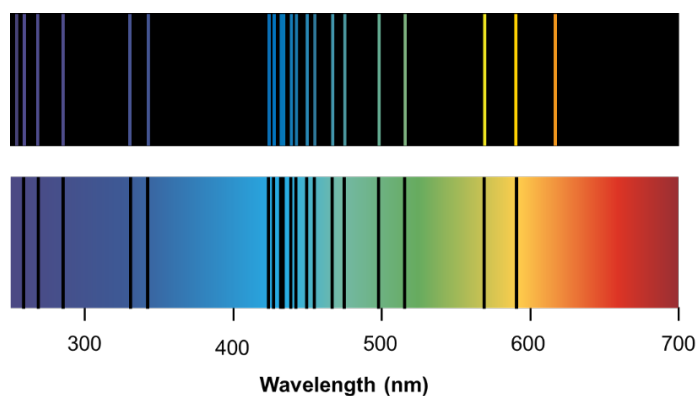


Figure 1.30: Emission and absorption spectra

Atomic numbers and mass numbers provide information about the numbers of subatomic particles in an atom.

Many elements consist of several different isotopes with different physical properties but the same chemical properties.

- Represent isotopes of an element using appropriate notation.

A chemical **element** is a material composed of one type of atom. There are 92 naturally occurring elements and 26 synthetic elements, each identified by a one or two-letter symbol (X). Furthermore, elements are also identified by their atomic, neutron and mass numbers. The **atomic number** (Z) represents the number of protons, the neutron number (N) represents the number of neutrons, and the **mass number** (A) is the sum of protons and neutrons. The usual notation for an element uses the symbol, with a pre-subscript equal to the atomic number Z and a pre-superscript equal to the mass number A , A_ZX . For example, ${}^{12}_6\text{C}$ denotes carbon-12 and ${}^{14}_6\text{C}$ denotes carbon-14. **Figure 1.31** lists values of A , Z , and N for several elements. The table shows some elements with the same atomic number but different neutron and mass numbers.

Atoms of a given element with different neutron and mass numbers are called **isotopes** of the element.

Examples of isotopes are carbon-12 and carbon-14.

Each has six protons, but carbon-12 has six neutrons, while carbon-14 has eight. Isotopes of an element have different physical properties, such as densities, melting and boiling points and diffusion rates, resulting from their different masses, but identical chemical properties as their atoms have the same electron arrangement. The chemical names of isotopes are commonly represented by stating the element's name followed by the mass number. The number of neutrons in an atom is determined by subtracting the atomic number from the mass number, as shown in the example below.

Example 1.01

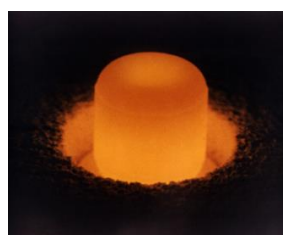
Plutonium-238 is an isotope used to fuel space probes as a sustained power source.

Calculate the number of neutrons in plutonium-238.

$$\begin{aligned} N &= A - Z \\ N &= 238 - 94 \\ N &= 144 \end{aligned}$$

Element	Symbol	Z	N	A
hydrogen	${}^1_1\text{H}$	1	0	1
helium	${}^4_2\text{He}$	2	2	4
lithium-6	${}^6_3\text{Li}$	3	3	6
lithium-7	${}^7_3\text{Li}$	3	4	7
beryllium	${}^9_4\text{Be}$	4	5	9
boron	${}^{11}_5\text{B}$	5	6	11
carbon-12	${}^{12}_6\text{C}$	6	6	12
carbon-14	${}^{14}_6\text{C}$	6	8	14
nitrogen	${}^{14}_7\text{N}$	7	7	14
oxygen	${}^{16}_8\text{O}$	8	8	16
aluminium	${}^{27}_{13}\text{Al}$	13	14	27
iron	${}^{56}_{26}\text{Fe}$	26	30	56
copper	${}^{65}_{29}\text{Cu}$	29	36	65
silver	${}^{108}_{47}\text{Ag}$	47	61	108
gold	${}^{197}_{79}\text{Au}$	79	118	197
lead	${}^{206}_{82}\text{Pb}$	82	124	206
uranium-235	${}^{235}_{92}\text{U}$	92	143	235
uranium-238	${}^{238}_{92}\text{U}$	92	146	238
oganesson	${}^{294}_{118}\text{Og}$	118	176	294

Figure 1.31: Compositions of several elements.



Question 14

The table below gives the names and properties of some isotopes used in nuclear medicine.

Isotope	X	Z	N	A
oxygen-15	$^{15}_8\text{O}$			
fluorine-18	$^{18}_9\text{F}$			
gallium-67	$^{67}_{31}\text{Ga}$			
rubidium-81	$^{81}_{37}\text{Rb}$			
technetium-99	$^{99}_{43}\text{Tc}$			
indium-111	$^{111}_{48}\text{In}$			
iodine-123	$^{123}_{53}\text{I}$			
thallium-201	$^{201}_{81}\text{Tl}$			

Complete the table by stating the atomic number (Z), neutron number (N) and mass number (A).

(8 marks) KA4

Question 15

The table below contains data for four different atoms, A, B, C and D.

Atom	N	A
A	10	22
B	12	22
C	16	28
D	18	34

(a) Identify the atom with the lowest atomic number.

_____ (1 mark) KA1

(b) Two of the atoms are isotopes of an element.

(1) Define an isotope.

_____ (1 mark) KA1

(2) Identify the two atoms in the table that are isotopes of the same element.

_____ (1 mark) KA1

(3) State why the two isotopes have the same chemical properties.

_____ (1 mark) KA1

(c) Identify the atoms in the table with the following electron configurations:

(1): 2,8: _____ (2): 2,8,6: _____ (2 marks) KA1

1.3: Quantities of atoms

The quantities of different substances can be conveniently compared using the mole unit.

The relative atomic mass of an element is determined from all the isotopes of that element.

The number of moles of atoms in a sample can be determined from the number of atoms present or the mass of the atoms.

- Undertake calculations using the relationship $n = \frac{m}{M}$ and its rearrangements.

Materials are composed of one or more types of **elementary particles**. For example, elements are composed of atoms, molecular compounds are composed of molecules, continuous compounds are composed of empirical formula units, and electrolytes are composed of ions. The number of elementary particles in a material is expressed using the **mole unit**.

The Mole

The **mole** (symbol, n) is a unit of measurement defined as an amount of a substance that contains as many elementary particles as there are atoms in 12 grams of pure carbon-12 (Figure 1.34).

The number of atoms in 12 grams of carbon-12 is approximately 6.02×10^{23} . Hence, one mole of any material contains 6.02×10^{23} elementary particles. For example, Figure 1.35 shows one-mole samples of copper, water and sodium chloride. The one-mole sample of copper, an element, contains 6.02×10^{23} copper atoms, the one-mole sample of water, a molecular compound, contains 6.02×10^{23} water molecules and the one-mole sample of sodium chloride, a continuous compound, contains 6.02×10^{23} empirical formula units of sodium chloride equivalent to 6.02×10^{23} sodium ions and 6.02×10^{23} chloride ions. The number of elementary particles (symbol N) in a sample is calculated by multiplying the molar quantity of the material by **Avogadro's number** (symbol N_A), $6.02 \times 10^{23} \text{ mol}^{-1}$.

Example 1.02

A mercury thermometer contains 0.0037 mol of mercury.

Calculate the number of mercury atoms in the thermometer.

N	=	$n \times N_A$
N	=	$0.0037 \times 6.02 \times 10^{23}$
N	=	1.1×10^{19} atoms



Figure 1.34: One mole of carbon-12



Figure 1.35: Molar quantities

Molar Mass

One mole of any material contains the same number of elementary particles as one mole of any other material. However, elementary particles have different masses, resulting in one-mole quantities of different materials having different masses. For example,

Figure 1.36 shows the masses of one-mole quantities of six different materials. The mass of one mole of a material is called the **molar mass** (symbol, M), and its units are grams per mole (g mol^{-1}). For example, the molar mass of sodium chloride is 58.5 g mol^{-1} , indicating one mole of sodium chloride weighs 58.5 g.

The chemical elements have well-defined molar masses, each written below the element's name on the periodic table. For example, **Figure 1.37** shows the molar masses of the elements carbon, nitrogen and oxygen. The molar masses of the elements are used when determining the molar mass of a compound. In general, the molar mass of a compound is determined by summing the molar masses of the atoms that comprise the compound's chemical formula.



Figure 1.36: One-mole quantities

6	7	8
C	N	O
Carbon	Nitrogen	Oxygen
12.01	14.01	16.00

Figure 1.37: Molar masses of elements.

Example 1.03

The molar mass of carbon dioxide, CO_2 , a molecular compound, is calculated by summing the molar masses of one carbon and two oxygen atoms.

$$\begin{aligned} M_{\text{CO}_2} &= M_{\text{C}} + (M_{\text{O}} \times 2) \\ M_{\text{CO}_2} &= 12.01 + (16.00 \times 2) \\ M_{\text{CO}_2} &= 44.01 \text{ g mol}^{-1} \end{aligned}$$

Example 1.04

The molar mass of calcium carbonate, CaCO_3 , a continuous compound, is calculated by summing the molar masses of one calcium, one carbon and three oxygen atoms.

$$\begin{aligned} M_{\text{CaCO}_3} &= M_{\text{Ca}} + M_{\text{C}} + (M_{\text{O}} \times 3) \\ M_{\text{CaCO}_3} &= 40.08 + 12.01 + (16.00 \times 3) \\ M_{\text{CaCO}_3} &= 100.09 \text{ g mol}^{-1} \end{aligned}$$

Question 18

The photograph shows one mole of sucrose, copper sulfate, sulfur, sodium chloride, mercuric oxide, and copper.



- (a) Define one mole of sucrose.

(1 mark) KA1

- (b) The molar mass of sulfur is 32.06 g mol^{-1} .

- (1) Define the molar mass of sulfur.

(1 mark) KA1

- (2) State the mass of sulfur shown in the photograph and give a reason.

(2 marks) KA2

- (c) Mercury oxide, HgO , is a red solid composed of mercury and oxide ions in a 1:1 ratio.

State the number of mercury ions in the sample of mercury oxide shown in the photograph and give a reason.

(2 marks) KA2

- (d) Copper sulfate, CuSO_4 , is the blue solid to the left of the photograph, and copper, Cu , is the red-brown metal to the right of the photograph.

- (1) State which sample, copper sulfate or copper, in the photograph has the most copper atoms and give a reason.

(2 marks) KA2

- (2) The copper sulfate sample is a hydrated compound with the formula $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$, where x represents the number of water molecules, H_2O , in one formula unit.

Calculate the value of x if the molar mass of $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$ is $249.69 \text{ g mol}^{-1}$.

(2 marks) KA4

Atomic Radius

An element's **atomic radius** measures the size of its atoms and is defined as the average distance from the centre of the nucleus to the boundary of the valence shell. **Figure 1.41** shows the atomic radii of helium and magnesium atoms.

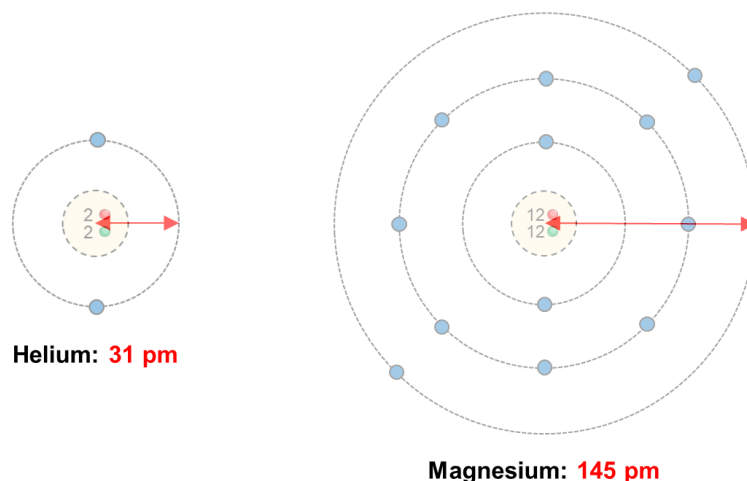


Figure 1.41: Atomic radii of helium and magnesium atoms measured in picometres ($\times 10^{-12}$ m)

Figure 1.42 shows trends in the atomic radii of the elements on the periodic table. The diagram shows that elements located in the bottom left of the table have large atomic radii, and those located at the top right have small atomic radii. Furthermore, the atomic radii of elements increase down a group and decrease across a period.

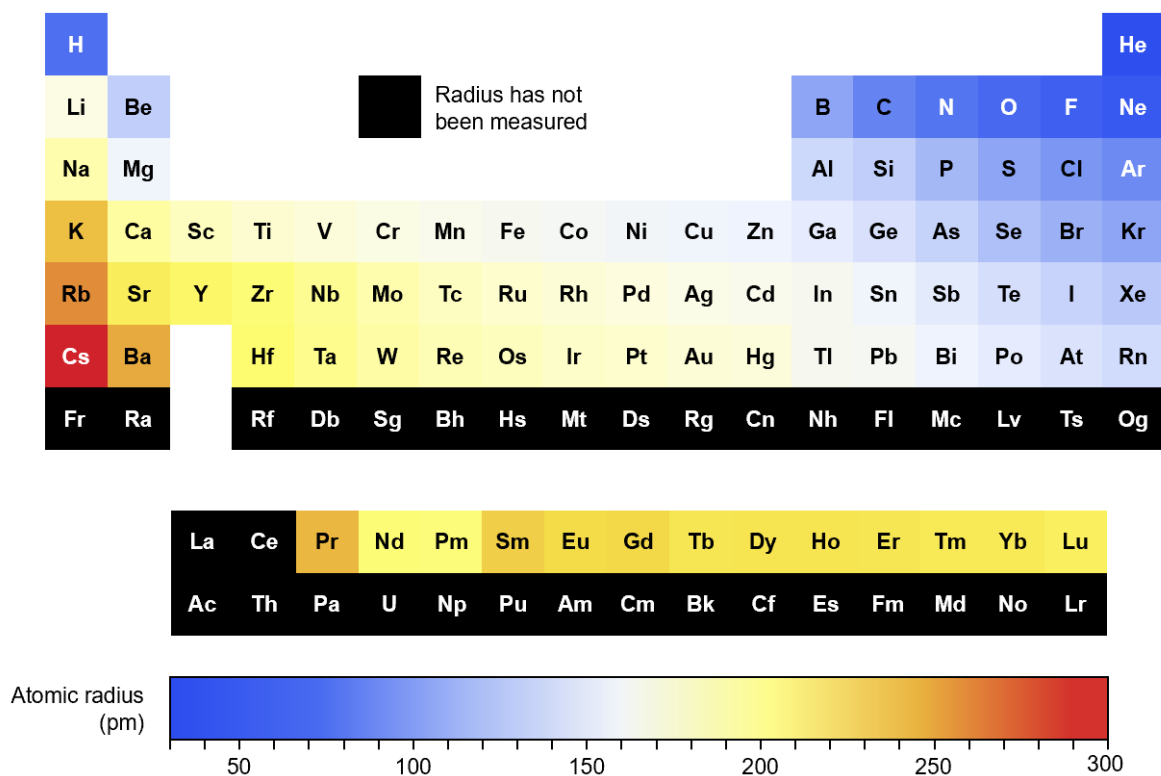


Figure 1.42: Trends in the atomic radii of the elements.

The atomic radii of the elements increase down a group. For example, **Figure 1.43** shows the trend in atomic radii down group 2 on the periodic table. The increase in atomic radius down a group is due to the increase in the number of electron shells in the atoms. For example, barium (Ba) has six electron shells and has a larger atomic radius than strontium (Sr) which has five electron shells.

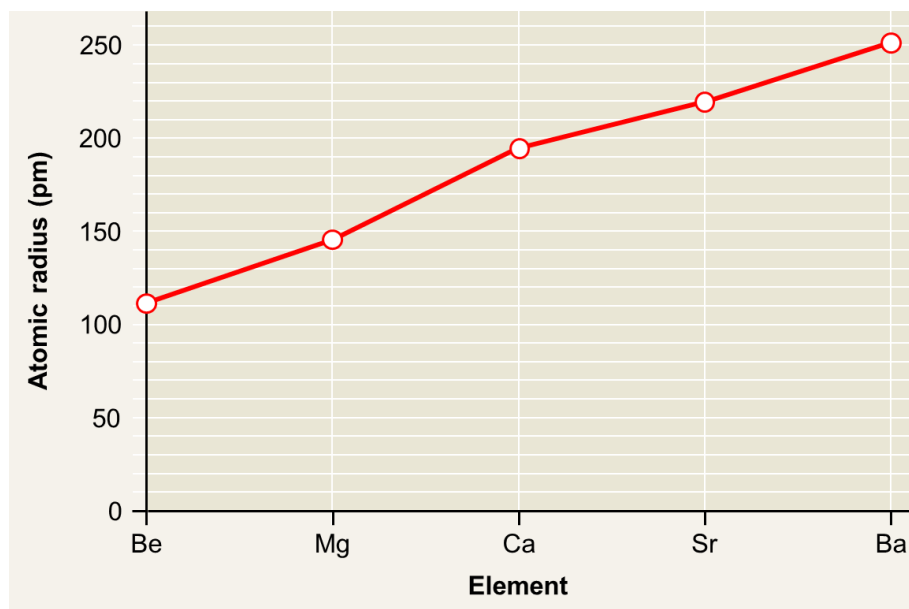


Figure 1.43: Change in atomic radii down group 2.

The atomic radii of the elements decrease from left to right across a period. For example, **Figure 1.44** shows the trend in atomic radii from left to right across period 3. The decrease in atomic radius is due to the increase in atomic number. As the atomic number increases, the protons attract the electrons with greater force, causing the electron shells to move towards the nucleus, reducing the atomic radius. For example, sodium (Na) has the lowest atomic number in period 3 and the largest atomic radius and argon (Ar) has the highest atomic number and smallest atomic radius.

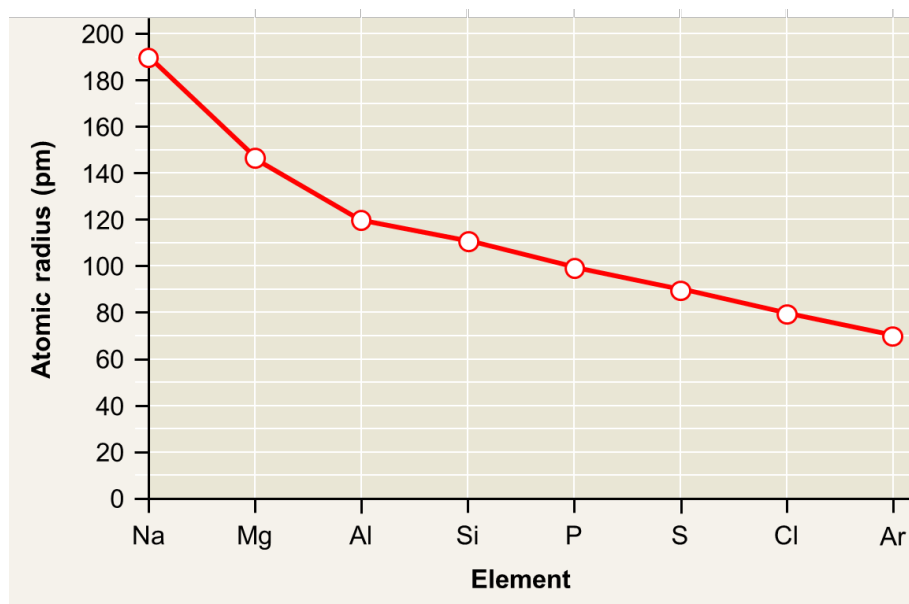


Figure 1.44: Change in atomic radii across period 3

2.1: Types of materials

Materials can be classified according to their structure and bonding into four types of substances. Melting points can be used to classify materials into molecular and non-molecular lattices. The electrical conductivity of non-molecular materials provides evidence for three types of primary bonding: metallic, ionic, and covalent.

- Classify materials as molecular, metallic, ionic, and covalent network, given relevant conductivity and melting point data.

Materials are composed of particles, including atoms, ions and molecules. The chemical structure of a material refers to the arrangement of its particles, and bonding refers to the types of attractive electrical forces between particles. Materials are classified according to their chemical structure and bonding into four substances: metallic, ionic, covalent network, and covalent molecular.

Metallic Substances

Metallic substances, including metallic elements and alloys, are materials composed of one or more metal atoms. Metals are typically solid at room temperature, with few exceptions. In the solid phase, metallic substances are composed of metal cations surrounded by a sea of delocalised electrons arranged into a regular repeated three-dimensional structure called a **lattice**. The metal cations and delocalised electrons are joined within the lattice by attractive electrical forces called **metallic bonds**. For example, **Figure 2.01** shows the structure and bonding within a pure metal.

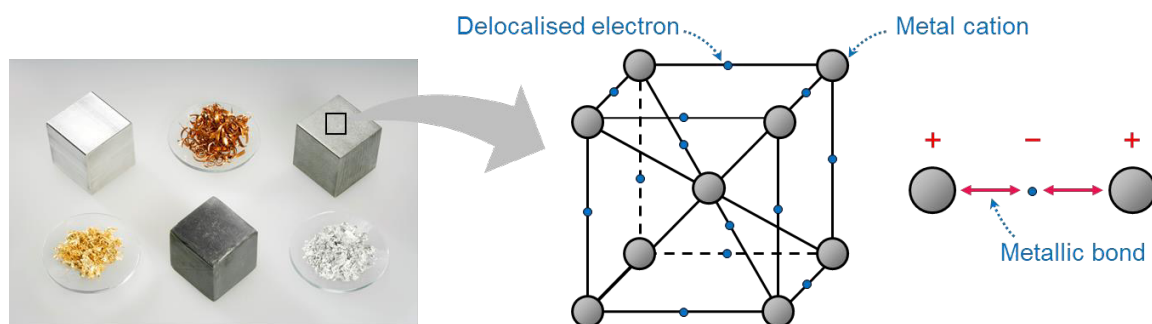


Figure 2.01: Structure and bonding in a solid metallic substance.

Metallic substances tend to have high melting points as a large quantity of heat must be transferred to overcome the strong metallic bonds, allowing the metal cations to separate and transition to the liquid phase of matter. In addition, metallic substances have good electrical conductivity as the delocalised electrons are free to move, producing an electric current when the material is connected to a source of electric potential difference, such as a battery. For example, **Figure 2.02** shows a simple electric circuit comprising a battery, light bulb, wires, and an iron nail. The light bulb glows as electrons move freely through the circuit.

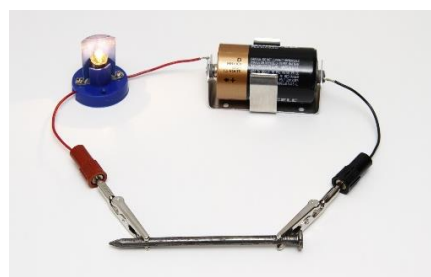


Figure 2.02: Metal conductivity

Ionic Substances

Ionic substances, including salts and electrolytes, are compounds composed of cations and anions combined in a fixed ratio. In the solid phase, the cations and anions are joined by attractive electrical forces called **ionic bonds** and arranged into a regular repeated three-dimensional lattice structure. For example, **Figure 2.03** shows the structure and bonding in a solid ionic compound.

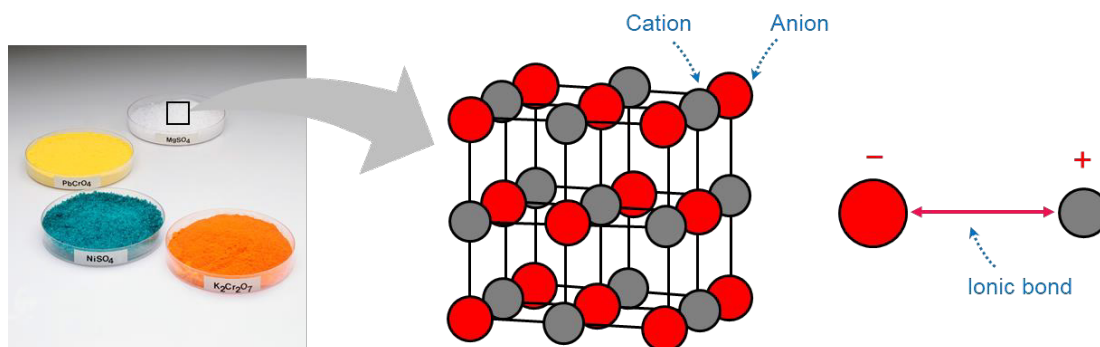


Figure 2.03: Structure and bonding in a solid ionic substance.

Ionic substances tend to have high melting points as a large quantity of heat must be transferred to overcome the strong ionic bonds, allowing the cations and anions to separate and transition to the liquid phase. When melted or dissolved in water, the ionic bonds weaken, and the cations and anions are free to move around. A dissolved or molten ionic substance is called an **electrolyte** as it contains separated ions. For example, **Figure 2.04** shows solid copper sulfate, CuSO_4 , an ionic compound composed of copper cations (Cu^{2+}) and sulfate anions (SO_4^{2-}). When solid copper sulfate is dissolved in water, the cations and anions separate, forming an electrolyte.



Figure 2.04: Electrolyte formation

Ionic substances are poor conductors in the solid phase and good conductors when melted or dissolved in water. In the solid phase, the cations and anions are strongly bonded and are not free to move when the salt is connected to an electric potential difference. However, when the ionic substance is melted or dissolved in water, the cations and anions separate and flow freely, producing an electric current when the solution is connected to a source of electric potential difference. For example, **Figure 2.05** shows an electric circuit in which two electrodes are placed in a beaker of solid sodium chloride, NaCl . In the solid phase, the sodium (Na^+) and chloride ions (Cl^-) are bonded strongly and cannot flow between the electrodes, inhibiting the flow of charge to the light bulb. However, when dissolved in water, the sodium and chloride ions separate and flow freely between the electrodes, completing the circuit and powering the light bulb.



Figure 2.05: Ionic substance conductivity

Covalent Network Substances

Covalent network substances, including silicon, germanium, silica, diamond, and graphite, are pure substances or compounds composed of one or more types of atoms combined in a fixed ratio. In the solid phase, the atoms are arranged into a regular repeated three-dimensional lattice structure and held together by attractive electrical forces called **covalent bonds**. For example, silica, SiO_2 is composed of silicon (Si) and oxygen (O) atoms combined in a 1:2 ratio. The atoms are joined by strong covalent bonds and arranged into a regular repeated lattice structure (Figure 2.06).

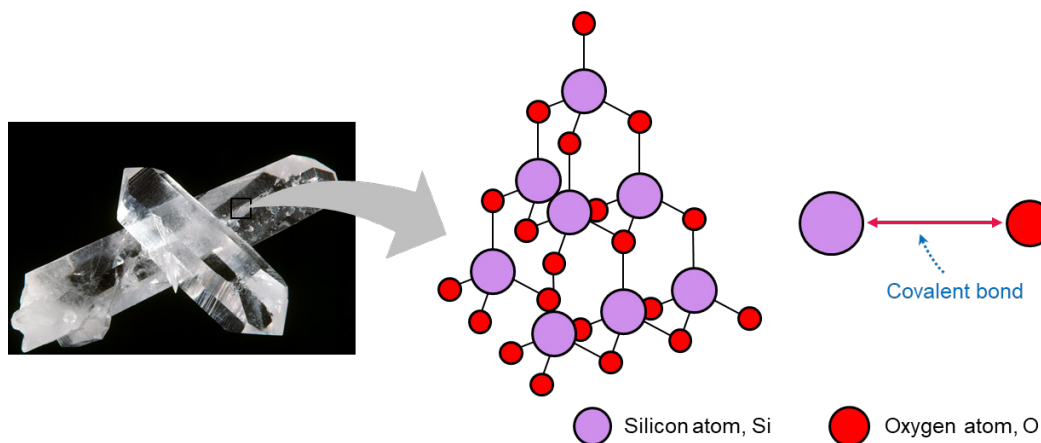


Figure 2.06: Structure and bonding in silica, a covalent network substance.

Covalent network substances tend to have very high melting points as a large quantity of heat must be transferred to overcome the strong covalent bonds, allowing the atoms to separate and transition to the liquid phase of matter. As a result, many covalent network substances, including boron nitride, silicon carbide, silica, diamond and graphite, have melting points greater than 2000°C .

Covalent network substances are poor electrical conductors with some exceptions. In all phases, the atoms comprising covalent network substances are uncharged, and their flow does not constitute an electric current. However, some covalent network substances, including silicon, and germanium are weak electrical conductors, called **semiconductors**, and others, such as graphite and graphene, are excellent conductors. Semiconductors contain electrically charged particles, including electrons and ions that flow as tiny currents when the material is connected to a source of electric potential difference. Similarly, graphite and graphene contain delocalised electrons in their structures that flow freely as an electric current when these materials are connected to a source of electric potential difference. For example, Figure 2.07 shows an electric circuit containing a battery, wires, a light bulb and a graphite rod. When the wires are connected to the battery, the delocalised electrons in the wires and graphite rod flow as a current through the circuit and power the light bulb.

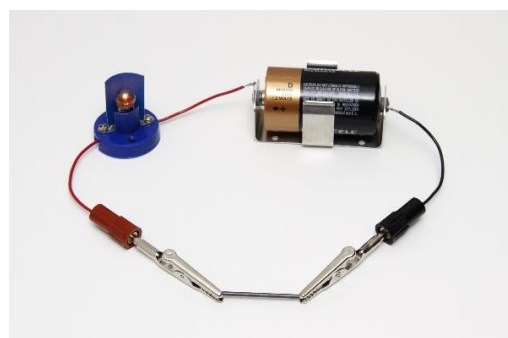


Figure 2.07: Graphite conductivity

Covalent Molecular Substances

Covalent molecular substances are compounds composed of **molecules**, electrically neutral groups of two or more atoms held together by attractive electrical forces called **covalent bonds**. In the solid phase, the molecules have a regular and orderly structure. However, the molecules are separated in the liquid and gas phases, and there is no orderly arrangement. In all phases, the molecules exert weak attractive electrical forces on one another called **secondary interactions**. For example, oxygen, O_2 is a covalent molecular compound composed of small oxygen molecules, each comprised of two oxygen atoms joined by covalent chemical bonds. The oxygen molecules are in the gas phase at room temperature, but at -183°C , they condense and form a pale blue liquid shown in **Figure 2.08**.

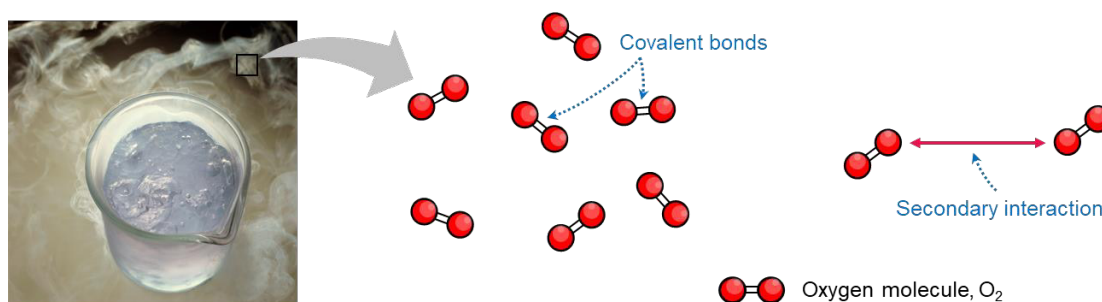


Figure 2.08: Structure and bonding in oxygen, a covalent molecular substance.

Covalent molecular substances may contain small or large molecules. Substances containing small molecules typically have low melting points as very little heat must be transferred to overcome the weak secondary interactions, allowing molecules to separate and transition to the liquid phase. In contrast, the secondary interactions are stronger between larger molecules, increasing the melting point. However, the melting points of covalent molecular substances are much lower than those of metallic, ionic and covalent network substances as the particles in these substances are joined by primary bonds. In contrast, weak secondary interactions join the molecules in covalent molecular substances.

For example, **Figure 2.09** shows samples of sugar and table salt in an oven. Sugar is a covalent molecular substance that melts at a much lower temperature than table salt, an ionic substance.



Figure 2.09: Sugar and salt in oven

Most covalent molecular substances have poor electrical conductivity in all phases as the molecules are uncharged, and their flow does not constitute an electrical current. For example, **Figure 2.10** shows an electric circuit containing two electrodes, wires, and a light bulb connected to a power supply (not shown). The water molecules must flow and carry a charge between the electrodes to power the light bulb. However, because water molecules are uncharged, their flow does not constitute a current, and hence, the circuit is incomplete, and charge does not flow into the light bulb.



Figure 2.10: Water conductivity

Question 34

The photograph below shows samples of sodium, a metallic substance, chlorine, a covalent molecular substance, and sodium chloride, an ionic substance.



- (a) Describe the structure and bonding in each substance.

(6 marks) KA2

- (b) The melting points and electrical conductivities of the three substances are shown below.

Substance	Melting point	Electrical conductivity
A	Low	Poor
B	High	Good
C	High	Poor

- (1) Identify the letter corresponding to sodium and give a reason.

(2 marks) KA2

- (2) Identify the letter corresponding to chlorine and give a reason.

(2 marks) KA2

Model for Metallic Bonding

The structure of a solid metal element or alloy is described using the model for metallic bonding.

Solid metals have a continuous structure composed of metal cations surrounded by a sea of delocalised electrons. The cations are arranged into a regular repeated three-dimensional lattice and are joined to the delocalised electrons by strong electrical forces of attraction called metallic bonds.

Figure 2.16 illustrates several lattice structures present in different metals. The lattices have different arrangements, but each contains metal cations surrounded by a sea of delocalised valence electrons (not pictured).

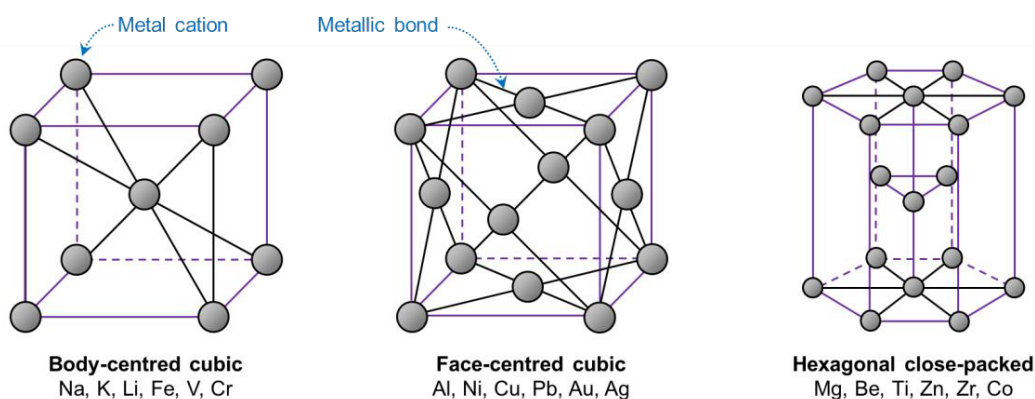


Figure 2.16: Lattice structure of different metals.

The physical properties of most metallic substances, including high melting and boiling points, and high electrical conductivity, are explained using the model for metallic bonding. Most metallic substances have high melting and boiling points as a large quantity of heat must be transferred to overcome the strong metallic bonds, allowing the cations to separate and transition to the liquid or gaseous phase. The high electrical conductivity is due to delocalised electrons that move freely as a current when the substance is connected to a source of electric potential difference.

In addition, metallic substances have high thermal conductivity, a property that enables rapid conductive heat transfer. When heated, the delocalised electrons absorb heat and transfer it rapidly throughout the metal structure, as shown in Figure 2.17.

Furthermore, metallic substances are also highly **malleable** and **ductile**, allowing these materials to be flattened under compression or stretched under tension. These properties are explained using the model for metallic bonding. When compressive or tensile forces are applied, the rows of cations and delocalised electrons move and slide past one another within the lattice. Still, no metallic bonds are broken, preventing the material from breaking. For example, Figure 2.18 shows a lead sample flattened under compressive forces.



Figure 2.17: Heated metal wire



Figure 2.18: Malleable lead

A covalent bond may be polar or non-polar.

- Use electronegativity values, or the position of atoms in the periodic table, to predict and explain the polarity of a covalent bond.
- Indicate the polarity of a covalent bond using the appropriate convention.

Atoms with higher electronegativities can complete their octets by sharing one or more electron pairs to form covalent bonds. The electron pairs may be shared evenly or unevenly between two atoms depending on the **electronegativities** of the two elements. In the context of covalent bonding, electronegativity is defined as the tendency of an element to attract a shared electron pair to itself when covalently bonded to another atom. When the bonding atoms have the same or similar electronegativities, they share the electron pairs evenly such that there is an even distribution of electric charge across the covalent bond. However, when the difference in electronegativity is sufficient, the electrons are attracted to one of the bonding atoms more than the other, leading to an uneven distribution of charge across the length of the covalent bond.

Polar Covalent Bond

When the difference in electronegativity of the bonding atoms is sufficiently large, the shared electron pairs are attracted towards the element with higher electronegativity, leading to an uneven distribution of electric charge across the length of the covalent bond. In such cases, the electron pair or pairs move towards the atom with the higher electronegativity and away from the element with lower electronegativity, producing a polarisation of electric charge across the covalent bond. The electron pair or pairs are positioned closer to the atom with higher electronegativity, giving it a partial negative charge, and away from the atom with lower electronegativity, giving it a partial positive charge. The separation of electric charge is called **polarisation**, and covalent bonds with polarisation are called **polar covalent bonds** or **polar bonds**. For example, hydrogen fluoride, HF is a covalent molecular compound with molecules composed of one hydrogen and one fluorine atom that share an electron pair, forming a single covalent bond (Figure 2.23). Fluorine has a much higher electronegativity than hydrogen, resulting in the shared electron pair being attracted towards the fluorine atom. The attraction of the electron pair causes fluorine to acquire a partial negative charge and hydrogen a partial positive charge. The uneven sharing of the electron pair causes polarisation across the hydrogen-fluorine bond. The polarity of the bond is indicated using the **delta notation**, a convention in which δ^+ represents the partially positive end of the bond and δ^- represents the partially negative end.

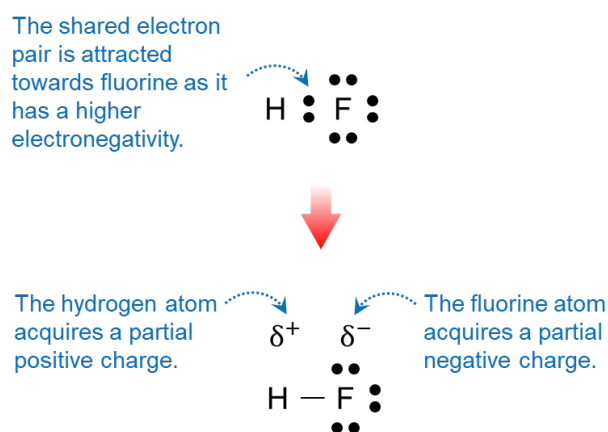


Figure 2.23: Polarity of H-F bond.

Non-Polar Covalent Bond

When the difference in electronegativity is sufficiently small or zero, the electron pairs are shared evenly, leading to an even distribution of electric charge across the length of the covalent bond. In such cases, there is no significant polarisation of charge across the covalent bond between the two atoms, and neither acquires a partial electric charge. Covalent bonds between two elements with no significant electronegativity difference are called **non-polar covalent bonds** or **non-polar bonds**. For example, chlorine, Cl_2 , is a covalent molecular compound composed of chlorine molecules. The atoms in a chlorine molecule share one electron pair, forming a single covalent bond (Figure 2.24). The two chlorine atoms have the same electronegativity and attract the electrons equally. Consequently, there is no charge polarisation, and the bond is non-polar.

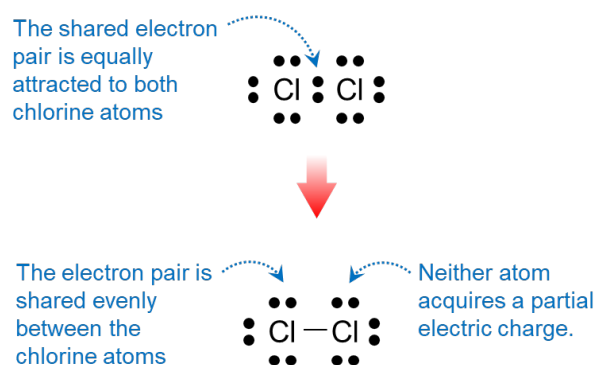


Figure 2.24: Polarity of Cl-Cl bond.

Covalent bonds between atoms of the same element are always non-polar as the atoms have the same electronegativity and attract shared electron pairs evenly. However, non-polar covalent bonds are also seen between atoms of similar electronegativities. For example, methane, CH_4 is a covalent molecular compound composed of methane molecules, each composed of a carbon atom covalently bonded to four hydrogen atoms. Each hydrogen atom shares one electron pair with carbon to fill its valence shell (Figure 2.25). The electron pairs are shared evenly as the difference in electronegativity between carbon and hydrogen atoms is very small, leading to no polarisation of charge across the bonds.

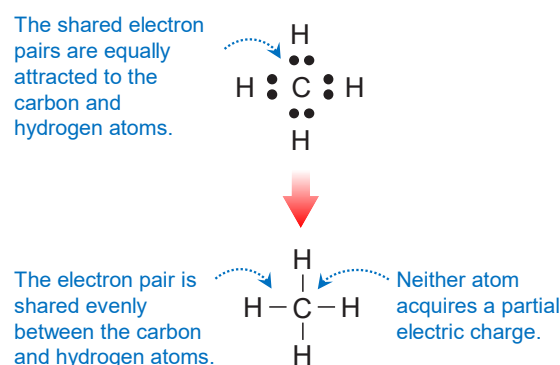


Figure 2.25: Polarity of C-H bond.

It is also common for a covalent molecular compound to feature polar and non-polar covalent bonds. For example, methanol, CH_3OH is a covalent molecular compound with polar and non-polar covalent bonding (Figure 2.26). The electron pairs between the carbon and hydrogen atoms are shared evenly as the electronegativity difference is negligible. However, the electronegativity difference between oxygen and hydrogen is significantly large, with oxygen attracting the electron pair more strongly than hydrogen.

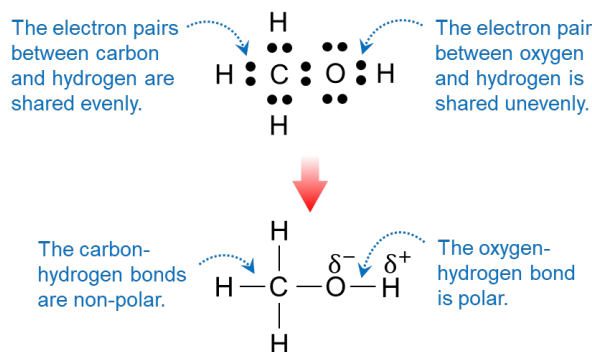


Figure 2.26: Polarity of bonds in CH_3OH

Covalent bonding is found in molecular and non-molecular (continuous) substances.

A molecule can be represented by a molecular formula.

A continuous covalent substance is represented by an empirical formula.

The physical properties of continuous covalent substances can be explained using the model for covalent bonding.

- Explain the melting point, hardness, and electrical conductivity of continuous covalent substances.

Covalent bonds are present in covalent molecular and covalent network substances. As described in **Chapter 2.1**, covalent molecular substances are compounds composed of molecules, each containing a discrete number of atoms joined by covalent bonds. Some covalent molecular compounds such as water, H_2O and carbon dioxide, and CO_2 are small molecules containing a few atoms. In contrast, others, such as polymers, are composed of giant molecules containing thousands of atoms. Covalent molecular compounds are represented using a **molecular formula** stating the number of atoms of each element present in one molecule of the covalent molecular compound. For example, the molecular formula of hydrogen peroxide is H_2O_2 denoting that one molecule of hydrogen peroxide contains two hydrogen atoms and two oxygen atoms. **Figure 2.27** lists the molecular formulae of some common covalent molecular substances.

Substance	Molecular formula
water	H_2O
carbon dioxide	CO_2
ammonia	NH_3
hydrogen sulfide	H_2S
sulfur dioxide	SO_2
methane	CH_4
nitric oxide	NO
hydrogen peroxide	H_2O_2
ethanol	$\text{C}_2\text{H}_5\text{OH}$
carbon monoxide	CO

Figure 2.27: Molecular formulae

In contrast, covalent network substances, also called **continuous covalent substances**, have a continuous structure composed of a large number of atoms joined by strong covalent bonds. The structure of a covalent network solid is described using the model for covalent bonding.

Covalent network substances have a continuous structure composed of covalently bonded atoms arranged into a regular repeated three-dimensional lattice.

Covalent network substances are common when non-metal atoms with a high valency, such as oxygen or nitrogen, share electrons with boron, carbon, and silicon as these elements have a high valency and form multiple covalent bonds, enabling the formation of a giant lattice structure. A covalent network compound is represented using an **empirical formula** stating the simplest ratio of the atoms in a covalent network compound. **Figure 2.28** lists the empirical formulae of some common covalent network substances. The physical properties of covalent network substances are explained using the model for covalent bonding. As described in

Substance	Empirical formula
boron nitride	BN
boron carbide	B_4C
silicon carbide	SiC
silicon dioxide	SiO_2
graphite	C
diamond	C
rhodium diboride	ReB_2

Figure 2.28: Empirical formulae

Chapter 2.1, covalent network solids have very high melting points and poor electrical conductivity, with some exceptions. Covalent network solids have high melting points as a large quantity of heat must be transferred to overcome the strong covalent bonds and separate the atoms, allowing them to transition to the liquid phase. **Figure 2.29** shows the melting points of some covalent network substances. It should be noted that some covalent network substances, including graphite and diamond, rarely melt under standard conditions.

Substance	Melting Point (°C)
boron nitride	3273
boron carbide	2350
silicon carbide	2730
silicon dioxide	1710
graphite	>3500
diamond	>3500
rhenium diboride	2400

Figure 2.29: Melting points

In addition to a high melting point, covalent network substances also have high hardness. **Hardness** measures the resistance of a substance to deformation when a force is applied. Most covalent network solids have high hardness as the strong covalent bonds restrict the movement of the atoms, preventing them from shifting when a force is applied. **Figure 2.30** shows the hardness of some covalent network substances measured using **Mohs hardness scale**.

Substance	Mohs Hardness
boron nitride	10
boron carbide	9
silicon carbide	9
silicon dioxide	7
graphite	1
diamond	10
rhenium diboride	10

Figure 2.30: Hardness scale

This scale measures hardness ranging from 0 (not hard) to 10 (very hard). **Figure 2.30** shows that most covalent network solids are hard substances except for graphite. Graphite is a covalent network substance composed of carbon atoms arranged into layers of repeating hexagonal rings. The layers are joined by weak secondary interactions, as shown in **Figure 2.31**. When sufficient force is applied, the layers slide past each other, making graphite a slippery material that is highly useful as a lubricant.

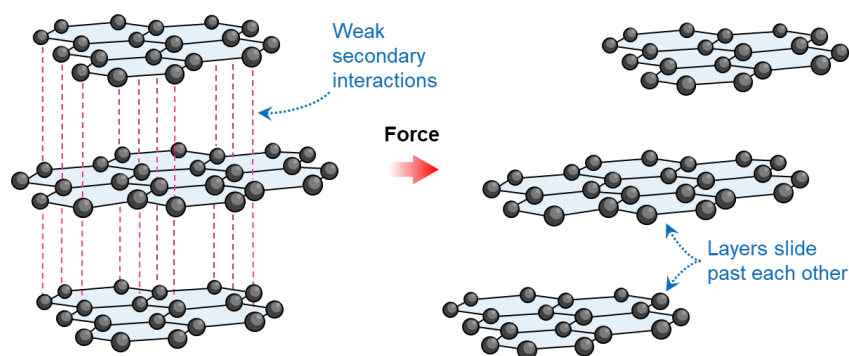


Figure 2.31: Movement of graphite layers when a force is applied

As mentioned in **Chapter 2.1**, most covalent network solids have poor electrical conductivity as there are no charged particles in the lattice to flow as an electric current when the substance is connected to a source of electric potential difference. Graphite is an exception to this trend. The carbon atoms in graphite share only three of their four valence electrons with other carbon atoms, resulting in most having a delocalised valence electron. The delocalised electrons move freely as an electric current when graphite is connected to a source of electric potential difference. The high electrical conductivity of graphite makes it a useful material for use in electrodes, batteries, and solar panels.

Question 52

The photograph shows a ceramic disc brake on a sports car.



- (a) The disc is made from silicon carbide, a covalent network solid.

The brakes are heated to a high temperature when braking.

State the property of silicon carbide that makes it suitable for use in ceramic disc brakes and explain this property using the model for covalent bonding.

(2 marks) KA2

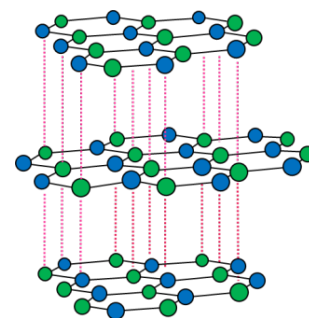
- (b) Silicon carbide is a hard material used in the protective ceramic plates in bulletproof vests.

Explain the hardness of silicon carbide using the model for covalent bonding.

(2 marks) KA2

Question 53

The diagram shows the chemical structure of boron nitride, a covalent network solid with a similar structure to graphite.



- (a) Unlike graphite, boron nitride is a poor electrical conductor.

Explain why boron nitride has poor electrical conductivity using the model for covalent bonding.

(2 marks) KA2

- (b) Boron nitride is used as a lubricant that reduces friction and wear in machines.

Explain, with reference to structure and bonding, why boron nitride is a useful lubricant.

(2 marks) KA2

Dispersion Forces

The electron clouds of molecules contain electrons in random motion. As the electrons move, they alter the distribution of electric charge within the electron cloud, leading to temporary polarisation. Regions of the electron cloud with a greater density of electrons are partially negative, δ^- and regions with fewer electrons are partially positive, δ^+ . The temporary polarisation of electron clouds leads to the formation of weak secondary interactions called **dispersion forces**. For example, iodine, I_2 is a molecular substance composed of iodine molecules. The electrons move randomly within the electron clouds of iodine molecules and will often become unevenly distributed, leading to temporary polarisation, as illustrated in **Figure 3.08**. Dispersion forces are formed between the temporary dipoles on the interacting molecules.

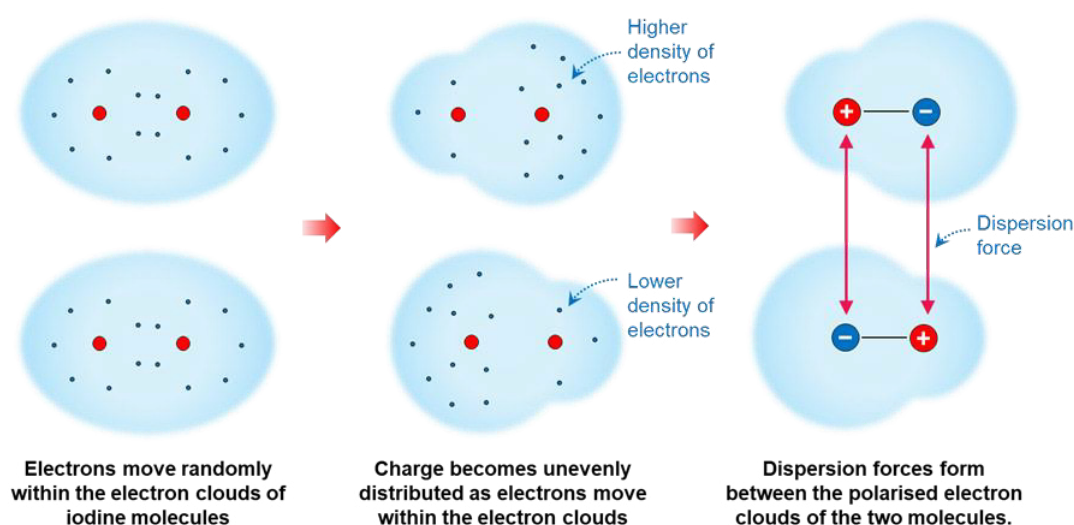


Figure 3.08: Dispersion forces between molecules.

Other molecules also polarise electron clouds. For example, when molecule B, with a temporary or permanent dipole, approaches the electron cloud of molecule A, the electron cloud of molecule B repels that of molecule A, polarising the electron cloud of molecule A, as in **Figure 3.09**.

The strength of the dispersion forces between molecules is dependent on the polarisability of their electron clouds. Dispersion forces are stronger between molecules with polarised electron clouds and weaker between molecules with less polarisable electron clouds. The polarisability of an electron cloud is dependent on many factors. Firstly, polarisability is dependent on molecular size. Larger molecules have more polarisable electron clouds as the electrons

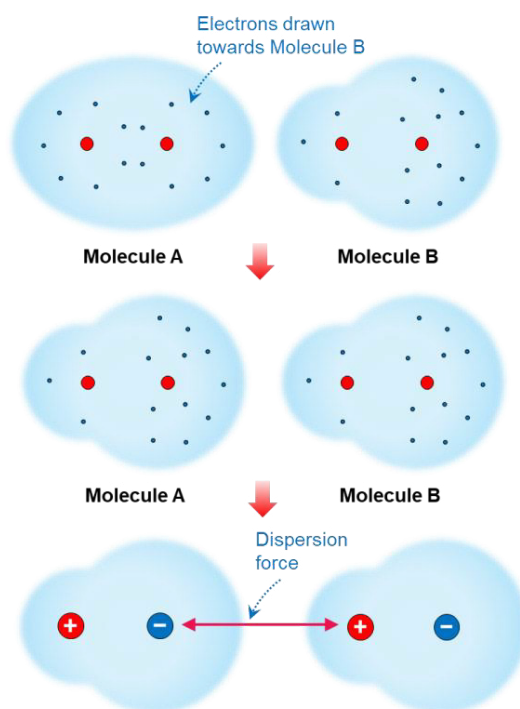


Figure 3.09: Polarisation of electron cloud.

Dipole-dipole interactions exist between polar molecules, and their strength depends on the polarity and size of the molecules.

Hydrogen bonding is a particularly strong form of dipole-dipole interaction that exists between molecules.

- Predict the relative strengths of interactions between molecules, given relevant information.
- Draw diagrams showing partial charges and hydrogen bonding between HF, H₂O, and NH₃ molecules.
- Explain the boiling points of HF, H₂O, and NH₃ in terms of hydrogen bonding between the molecules.

Dispersion forces are secondary interactions between the electron clouds of polar and non-polar molecules caused by the temporary polarisation of the electron cloud. In contrast, dipole-dipole interactions and hydrogen bonds are secondary interactions between polar molecules only.

Dipole-dipole interaction

A **dipole-dipole interaction**, also called a dipole-dipole force, is a secondary interaction between the positively-charged region (δ^+) of the electron cloud in one polar molecule and the negatively-charged region (δ^-) in the electron cloud of a second polar molecule. For example, in hydrogen bromide (HBr), the region of the electron cloud surrounding bromine is negatively charged, and the region surrounding hydrogen is positively charged. The positively charged region in the electron cloud of one HBr molecule attracts the negatively charged region in an adjacent molecule, forming a dipole-dipole interaction (Figure 3.10). Similarly, in methanal (CH₂O), the electron cloud is negatively charged near oxygen and positively charged near carbon, leading to the formation of dipole-dipole interactions between methanal molecules (Figure 3.10).

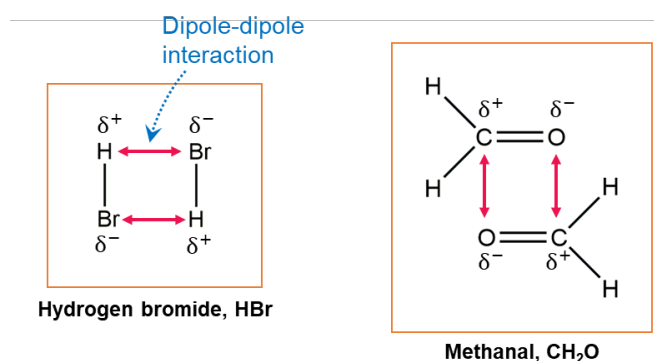


Figure 3.10: Dipole-dipole interactions between polar molecules.

Dipole-dipole interactions are stronger than dispersion forces due to the permanent polarisation of the electron clouds. Furthermore, the strength of dipole-dipole interactions between molecules is dependent on the magnitudes of the partial charges on the interacting atoms. In general, dipole-dipole forces are stronger between two dipoles with greater partial charges.

Hydrogen Bonding

A **hydrogen bond** is a type of dipole-dipole interaction between a partially positively charged (δ^+) hydrogen atom in one molecule and a partially negatively charged (δ^-) oxygen, nitrogen, or fluorine atom in a second molecule. Hydrogen bonds are stronger than other dipole-dipole interactions due to the larger difference in electronegativity between hydrogen (2.1) and either nitrogen (3.0), oxygen (3.5), or fluorine (4.0). The significant difference in electronegativity produces greater polarisation within the electron cloud, increasing the partial charges and the strength of the secondary interactions between the interacting molecules. **Figure 3.11** shows hydrogen bonding in hydrogen fluoride (HF), water (H_2O) and ammonia (NH_3).

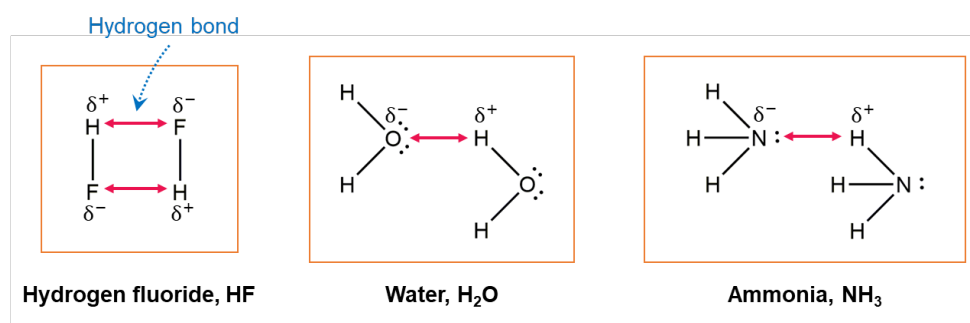


Figure 3.11: Hydrogen bonding between molecules in HF, H_2O and NH_3

The boiling points of covalent molecular compounds containing hydrogen bonds are higher than other molecules of similar size and shape as a larger quantity of heat must be transferred to overcome the hydrogen bonds and allow the molecules to separate and transition from liquid to gas.

Example 3.14

Figure 3.12 shows the boiling points of several molecular compounds called hydrides.

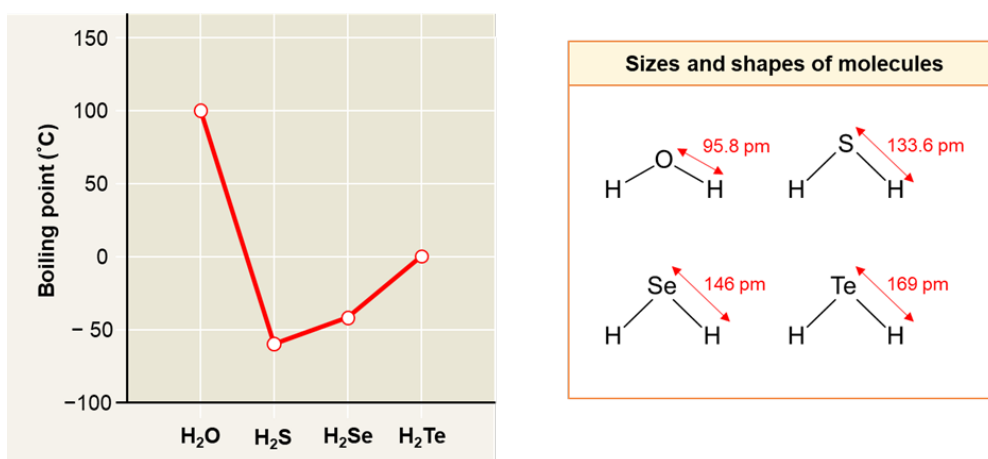


Figure 3.12: Boiling points of different hydrides.

The diagram shows that the boiling point of water (H_2O) is greater than the other hydrides, even though the molecules are all small and V-shaped. The variation in boiling point is due to the difference in the strength of the secondary interactions between molecules. Hydrogen bonds in water are stronger than the dipole-dipole interactions between molecules in the other three hydrides.

Example 4.01

Methanol, CH_3OH , is a polar solvent that is miscible in water. Therefore, when methanol is mixed with water, the secondary interactions between molecules in both materials are overcome and the molecules mix and form a solution, as illustrated and shown in **Figure 4.01**.

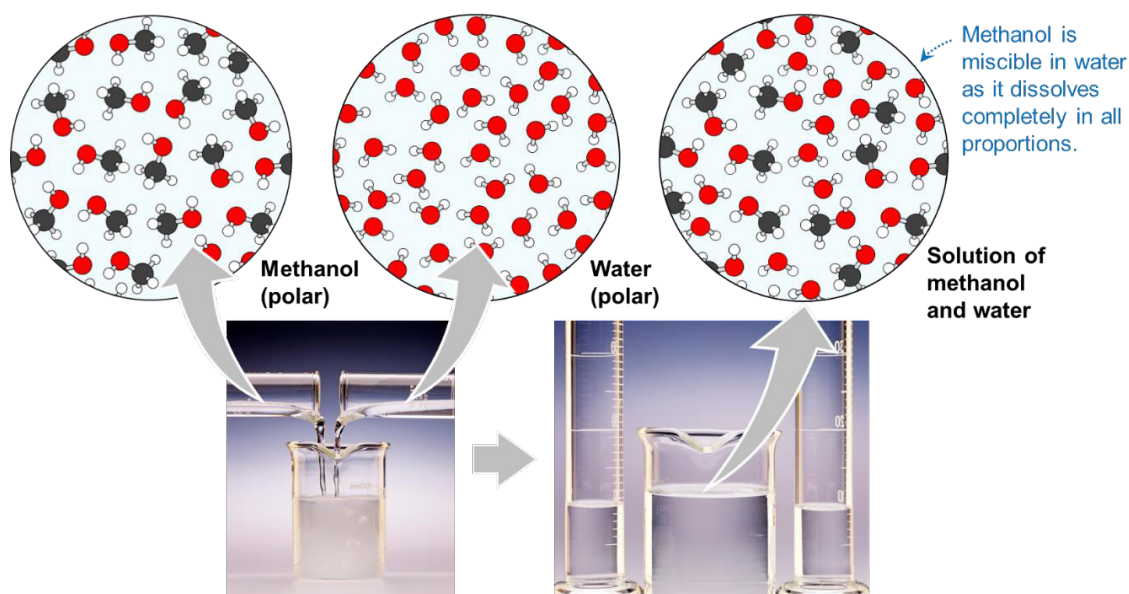


Figure 4.01: Methanol mixed with water.

Example 4.02

Hexane, C_6H_{14} , is a non-polar solvent that is immiscible in water. A network of dispersion forces joins the hexane molecules, and a network of hydrogen bonding joins the water molecules. Therefore, when mixed, the molecules remain joined, and the substances form two separate layers (**Figure 4.02**) as the water molecules cannot separate the hexane molecules and mix with them thoroughly, forming a solution.

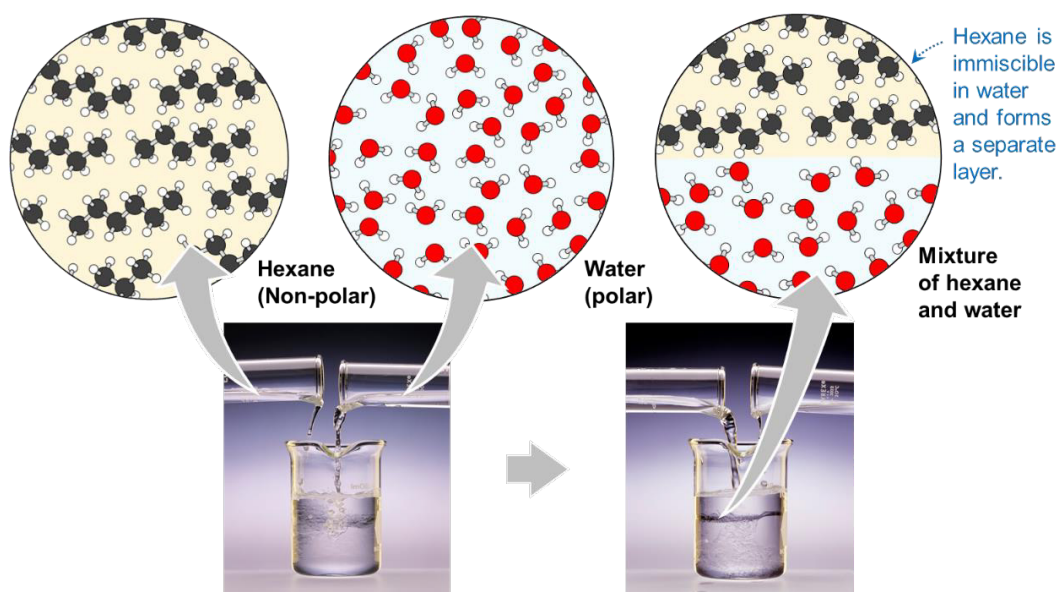


Figure 4.02: Hexane mixed with water.

4.2: Solutions of ionic substances

Many ionic substances are soluble in water. This is particularly so for ammonium and alkali metal salts.

Equations can be written to represent the dissociation and hydration of ions that occurs when ionic substances dissolve in water.

- Describe the formation of ion-dipole interactions when ionic substances dissolve in water.
- Write equations for the dissolving of ionic substances in water.

Many ionic compounds are soluble in water, especially those containing ammonium, lithium, sodium and potassium cations, or nitrate anions. Soluble ionic compounds dissolve in water to form a solution in two stages: dissociation and hydration.

Dissociation

Dissociation is the process by which the cations and anions are separated from the lattice when the compound is placed in water. The first stage of dissociation involves the binding of water molecules to the cations and anions, a process facilitated by the formation of **ion-dipole interactions**, secondary interactions between ions and water molecules. The oxygen atom of water has a partial negative charge (δ^-), and the hydrogen atoms have a partial positive charge (δ^+), facilitating the binding of water molecules to the ions in the lattice. **Figure 4.10** illustrates the ion-dipole interactions between the ions and water molecules when sodium chloride, NaCl is dissolved in water.

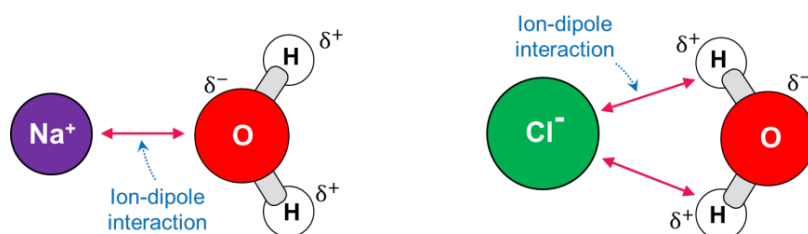


Figure 4.10: Ion-dipole interactions between ions and water molecules.

The second stage of dissociation involves water molecules removing ions from the lattice. The water molecules overcome the ionic bonds and separate the ions from the lattice if sufficient energy is available. **Figure 4.11** illustrates the dissociation of sodium and chloride ions in sodium chloride.

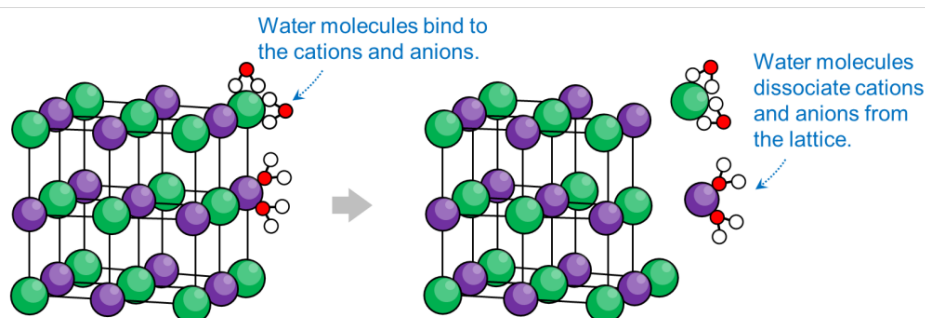


Figure 4.11: Dissociation of ions in sodium chloride.

Hydration

Hydration is the process in which water molecules surround the dissociated cations and anions and form a **hydration shell** that prevents the ions from recombining (Figure 4.12).

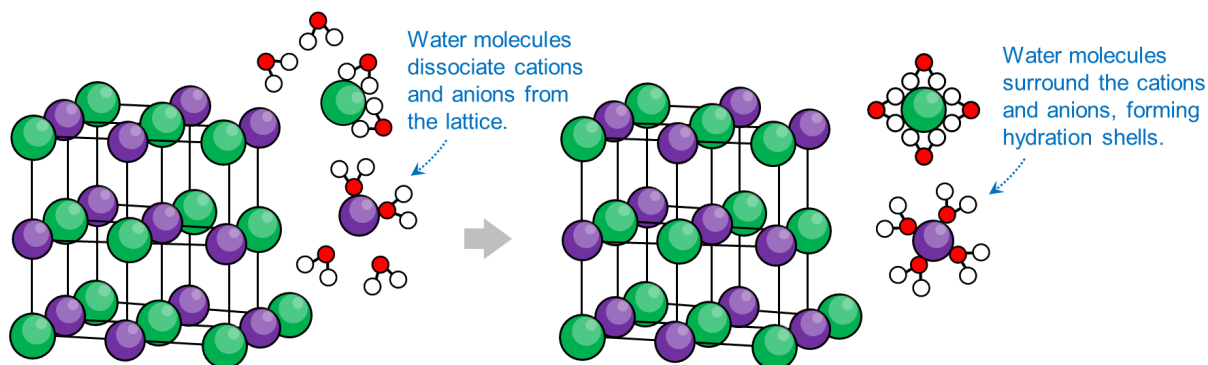


Figure 4.12: Hydration of ions following dissociation.

The processes of dissociation and hydration continue until all cations and anions have been separated from the lattice. Once the ions have been dissociated and hydrated, the ionic compound is said to be in a dissolved or aqueous state (symbol: aq) in which water molecules surround all cations and anions (Figure 4.13).

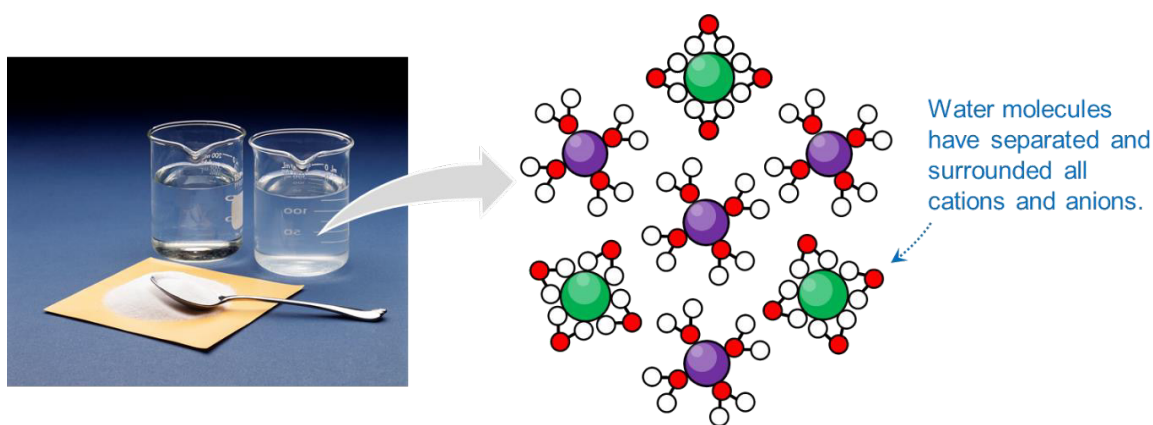
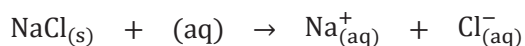


Figure 4.13: Dissolving of an ionic compound.

Dissociation Equations

The dissolving of an ionic compound in water is described in the form of a **dissociation equation**. The empirical formula of the ionic compound is written as a reactant along with the subscript (s) to denote that the material is initially solid. The other reactant is water which is represented using the symbol (aq) rather than H₂O. The products of the reaction are the hydrated cations and anions that are represented using the symbol of the ions and the subscript (aq) to denote the ions are in an aqueous or dissolved state. For example, the dissociation of sodium chloride, NaCl is described below.



Soap Scum

The effectiveness of soap is greatly reduced when the material is used in **hard water**. Hard water is the term used to describe water that has a high concentration of calcium, Ca^{2+} and magnesium ions, Mg^{2+} . When used in hard water, a precipitation reaction occurs in which soap anions combine with calcium and magnesium ions to form an insoluble precipitate called **soap scum** (Figure 4.15).

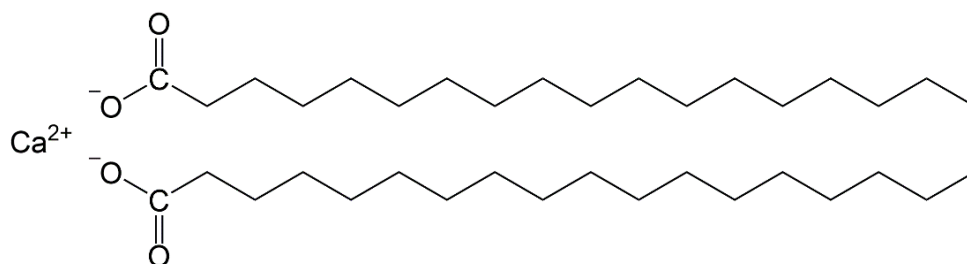
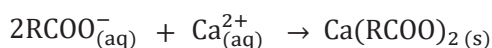


Figure 4.15: Insoluble soap scum.

The formation of soap scum is described in the net ionic equation below. The hydrocarbon component of the soap anion is denoted using the symbol R.



Soap scum is an ineffective cleaning agent as the ionic charge on the hydrophilic component of the molecule has been neutralised, preventing the material from forming micelles in water. Instead, the soap anions form an insoluble mass seen as a white precipitate in water and dries to form a white solid on surfaces, as shown in Figure 4.16.

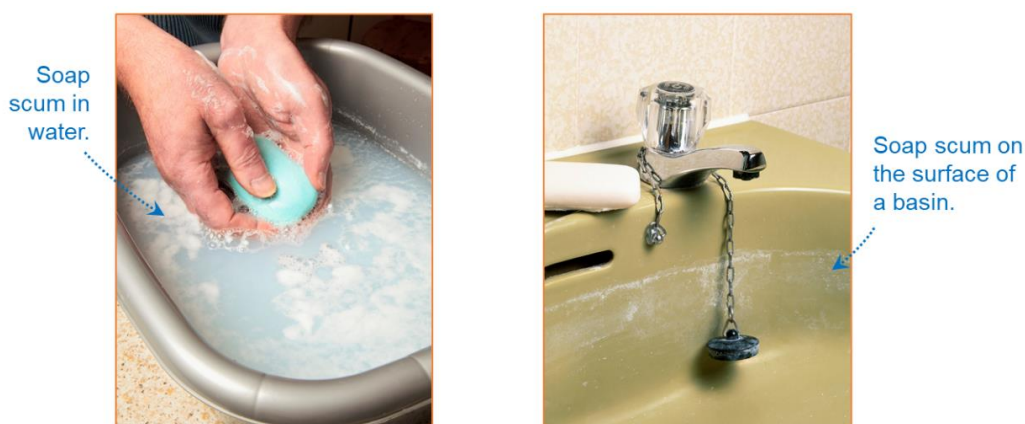


Figure 4.16: Soap scum in water (left) and on a surface (right).

The formation of soap scum is minimised using a **water softener**, a device containing two tanks. One tank contains millions of tiny resin beads that attract magnesium and calcium ions from hard water and exchange them with sodium ions, effectively softening the water. The second tank contains sodium chloride solution used to remove the magnesium and calcium ions from the beads and replenish the sodium ions.

The concentration of a solution can be described in terms of mass concentration (mass of solute per unit volume, ρ) or as a molar concentration (moles of solute per unit volume, c).

- Undertake calculations using the relationship $\rho = \frac{m}{V}$, and its rearrangements.
- Undertake calculations using the relationship $c = \frac{n}{V}$, and its rearrangements.
- Undertake conversions between mass concentrations and molar concentrations.

Concentration is a measure of the quantity of solute per unit volume of solution. Solution concentrations are described in terms of **mass concentration** (ρ) and **molar concentration** (c).

Mass concentration

The mass concentration of a solution is the mass of solute per unit volume of solution.

Formula	$\rho = \frac{m}{V}$	
Symbol	Variable	Unit
ρ	Mass concentration	g L^{-1}
m	Mass of solute	g
V	Volume of solution	L

Example 4.13

A coffee drink consists of 57.5 grams of sucrose dissolved in 600 mL of solution.

Determine the mass concentration of sucrose in the coffee drink.

$$\begin{aligned}\rho &= \frac{m}{V} \\ \rho &= \frac{57.5}{0.600} \\ \rho &= 95.8 \text{ g L}^{-1}\end{aligned}$$

Example 4.14

A bottle of balsamic vinegar contains ethanoic acid in a concentration of 60.0 g L^{-1} .

Determine the mass of ethanoic acid in 250 mL of balsamic vinegar.

$$\begin{aligned}m &= \rho V \\ m &= 60.0 \times 0.25 \\ m &= 15.0 \text{ g}\end{aligned}$$

Question 107

A chemical reaction requires the five solutions, which are identified in the table below.

Substance	c (mol L ⁻¹)	V (mL)	M (g mol ⁻¹)
potassium bromate	0.500	6.00	167.00
potassium bromide		1.00	119.00
sulfuric acid	6.00	0.600	98.08
malonic acid	0.500		104.06
ferroin	0.0250	1.00	

- (a) 60.0 mg of potassium bromide was measured out for the reaction.

Calculate the molar concentration of the potassium bromide solution.

(3 marks) KA4

- (b) Calculate the number of moles of potassium bromate present in the solution.

(2 marks) KA4

- (c) Calculate the mass concentration of sulfuric acid in g L⁻¹.

(2 marks) KA4

- (d) Calculate the volume of malonic acid solution if 1.25 mmol is present.

(2 marks) KA4

- (e) 15.0 mg of ferroin was used in the reaction.

Use the information above to determine the molar mass of ferroin.

(4 marks) KA4

Conversions

The mass concentrations in grams per litre, parts per million, parts per billion and percent weight per volume can be interconverted with the molar concentration. **Figure 4.17** summarises the conversions between the different mass concentrations and the molar concentration.

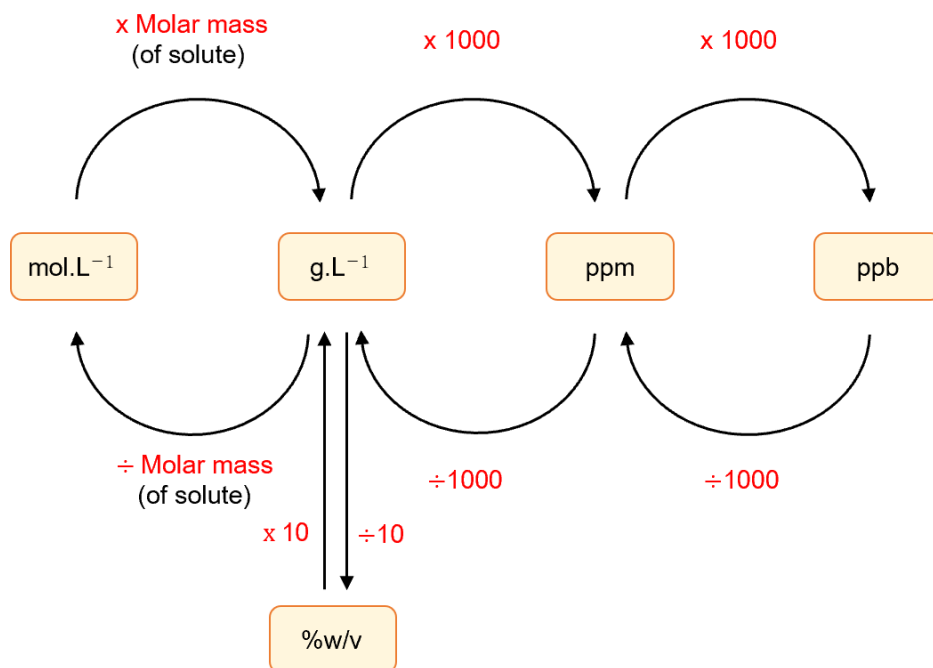


Figure 4.17: Conversions of mass concentrations and the molar concentration.

Example 4.23

The glucose concentration in the blood is 5.0 mmol L^{-1} .

Determine the mass concentration of glucose in g L^{-1} , ppm, and ppb if the molar mass of glucose is $180.156 \text{ g mol}^{-1}$.

$$\begin{aligned} \rho &= cM \\ \rho &= 5.0 \times 10^{-3} \times 180.156 \\ \rho &= 0.90 \text{ g L}^{-1} \end{aligned}$$

$$\begin{aligned} \rho &= 0.90 \times 1000 \\ \rho &= 9.0 \times 10^2 \text{ ppm} \end{aligned}$$

$$\begin{aligned} \rho &= 900 \times 1000 \\ \rho &= 9.0 \times 10^5 \text{ ppb} \end{aligned}$$

Dilution

Dilution is a method of decreasing the concentration of a solution by increasing the solvent volume while keeping the solute mass constant. **Figure 4.18** illustrates the process of dilution in which a 70 mL volume of solvent is added to the solution while keeping the solute mass constant. The diagram shows the number of solute particles is constant.

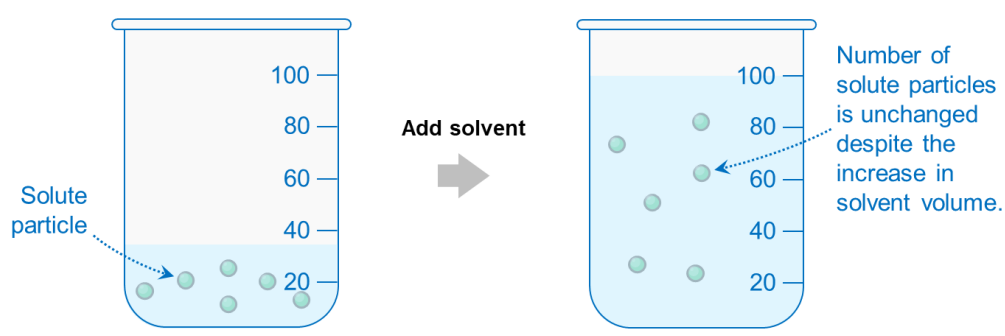


Figure 4.18: Dilution.

The number of moles of solute is unchanged by adding additional solvent. Hence, the product of the initial volume and concentration is equal to the product of the final volume and concentration, as shown in the formula below.

Formula	$n_1 = n_2 \therefore c_1V_1 = c_2V_2$
----------------	--

Symbol	Variable	Unit
c_1	Initial molar concentration	mol L^{-1}
V_1	Initial volume of solution	L
c_2	Final molar concentration	mol L^{-1}
V_2	Final volume of solution	L

Example 4.24

10 mL of 0.50 mol L^{-1} potassium dichromate solution was diluted to 250 mL with water. Calculate the final concentration of the potassium dichromate solution.

$$\begin{aligned}
 c_1V_1 &= c_2V_2 \\
 c_2 &= \frac{c_1V_1}{V_2} \\
 c_2 &= \frac{0.50 \times 10 \times 10^{-3}}{0.25} \\
 c_2 &= 0.02 \text{ mol L}^{-1}
 \end{aligned}$$



Standard Solution

A **standard solution** is a solution of precise concentration prepared by dissolving a known quantity of solute in an appropriate solvent. Standard solutions are primarily used to determine the concentrations of other substances through a technique called **titration** that is explored in Stage 2 Chemistry. Standard solutions may be prepared from solid, liquid, aqueous or gaseous solutes. The following steps are used when preparing a 1000 mL standard solution using solid copper sulfate as the solute and water as the solvent.

1. Weigh out the appropriate mass of solute required to produce the desired concentration.



2. Dissolve the solute using the solvent used to make the solution. In this example, copper sulfate is dissolved in pure water.



3. Transfer the solution to a **volumetric flask** with a volume equal to the desired volume of the standard solution. In this example, a 1000 mL volumetric flask is used as the desired volume of the standard solution is 1000 mL.



4. Add solvent until the volume approaches the neck of the volumetric flask.



5. Add solvent slowly as the solution approaches the graduation mark or calibration mark on the neck of the volumetric flask. This is done at eye level to ensure that the final volume of the solution is not greater or less than the desired volume.



6. The desired volume has been obtained when the meniscus of the solution rests on the graduation mark when viewed at eye level.



7. Once graduated, the volumetric flask may be stoppered and inverted once to ensure that the solute and solvent particles are distributed homogeneously throughout the solution.



Question 117

The photograph below shows the equipment used to prepare a 500.0 mL standard solution of calcium hydroxide, $\text{Ca}(\text{OH})_2$ in water.

- (a) Calculate the mass of calcium hydroxide required to prepare a 500.0 mL, $0.0225 \text{ mol L}^{-1}$ standard solution.



(4 marks) KA4

- (b) Name the piece of glassware used to prepare and store the standard solution.

(1 mark) KA2

- (c) Describe the steps involved in the preparation of the standard solution.

(6 marks) KA2

- (d) State the effect of the following on the final concentration of the standard solution.

(1) More water was added, and the solution's final volume was above the graduation mark.

(1 mark) KA2

(2) Some solid calcium hydroxide was not transferred to the flask.

(1 mark) KA2

Enthalpy changes for solution reactions can be determined experimentally.

- Explain the relationships $Q = mc\Delta T$ and $\Delta H = \frac{Q}{n}$ and undertake calculations involving their rearrangements.
- Experimentally determine enthalpies of solution.
- Identify a reaction as exothermic or endothermic, given a thermochemical equation or the value of its enthalpy change.

The molar enthalpy of solution can be determined experimentally using a technique called **calorimetry**. Calorimetry is a technique used to measure the quantity of heat absorbed or released in a chemical process. For example, in the context of dissolving an ionic compound in water, a solid sample is transferred to a known mass of water, causing the solid to dissolve and either absorb or release heat. The heat absorbed or released causes a change in the temperature (ΔT) of the solution, which is measured using a thermometer. The temperature change is used to calculate the quantity of heat using the formula below.

Formula	$Q = mc\Delta T$	
Symbol	Variable	Unit
Q	Quantity of heat	J
m	Mass of water	g
c	Specific heat capacity of water	$\text{J g}^{-1} \text{ }^\circ\text{C}^{-1}$
ΔT	Change in water temperature	$^\circ\text{C}$

The formula above determines the quantity of heat absorbed or released into the surrounding solution composed primarily of water. If the dissolving is endothermic, heat is transferred from the water to the reaction decreasing the water temperature. If the dissolving is exothermic, heat is transferred from the reaction, increasing the water temperature. The temperature change is dependent on the specific heat capacity of water which is the amount of energy required to increase the temperature of 1 g water by 1°C .

Example 4.28

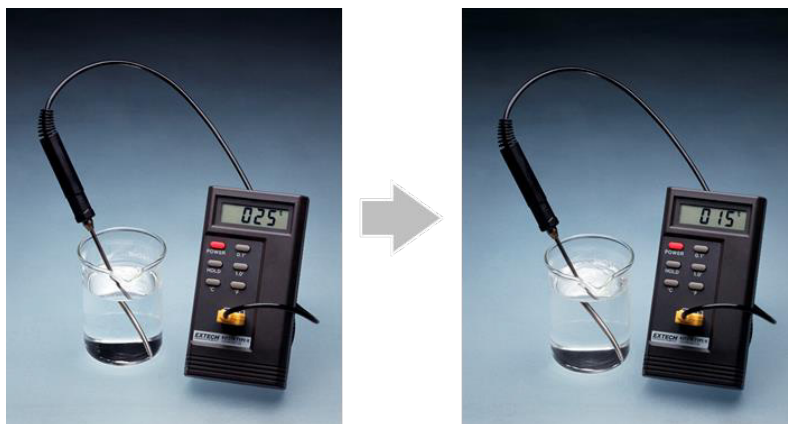
Sodium acetate was dissolved in 200 mL of water, increasing the temperature from 20.0 to 22.5°C . Calculate the quantity of heat released if the specific heat capacity of water is $4.18 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1}$.

$$\begin{aligned} Q &= mc\Delta T \\ Q &= 200 \times 4.18 \times (20.0 - 22.5) \\ Q &= -2090 \text{ J} \end{aligned}$$

The quantity of heat is negative, indicating that energy was released in this reaction.

Question 124

The photograph below shows the temperature change when potassium sulfate is added to water.



- (a) Describe and explain the change in water temperature in the reaction.

(2 marks) **KA2**

- (b) 44.2 g of potassium sulfate was dissolved in 150.0 g of water in the reaction.
Calculate the molar enthalpy of solution of potassium sulfate.

(3 marks) **KA4**

- (c) Write a thermochemical equation for the dissolving of potassium sulfate, K_2SO_4 in water.

(3 marks) **KA1**

Question 125

Calcium chloride was dissolved in 150.0 g of water and the temperature increased by 10.0°C.

- (a) The molar enthalpy of solution of calcium chloride, $CaCl_2$, is $-81.3 \text{ kJ mol}^{-1}$.

Write a thermochemical equation for the dissolving of ammonium nitrate in water.

(3 marks) **KA1**

- (b) Calculate the mass of calcium chloride that dissolved in the reaction.

(3 marks) **KA4**

Acid-base indicators are weak acids or bases where the acidic form is of a different colour from the basic form.

Acid-base indicators are chemical compounds or mixtures that change colour when reacted with acids and bases. The indicator colour in acid is different from the colour in a base. **Figure 5.03** shows the colours of several indicators in acidic and basic solutions.

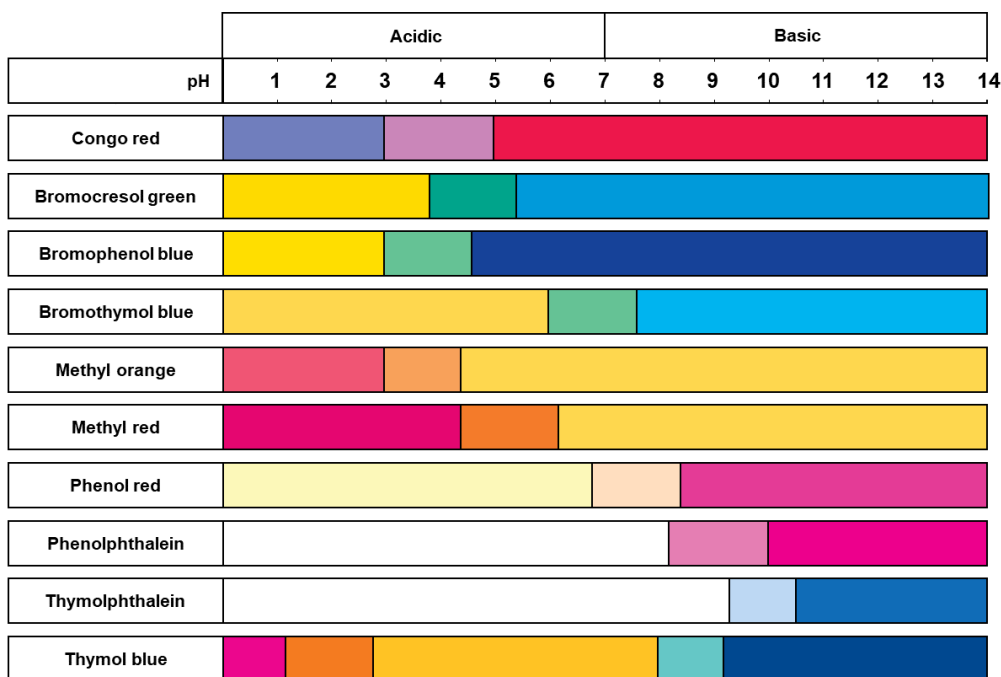


Figure 5.03: Acid-base indicator colour chart.

Acid-base indicators are usually organic compounds or mixtures of organic compounds composed of molecules classified as either weak acids or weak bases. The addition or removal of a proton causes a change in the indicator's molecular structure, which causes it to absorb and reflect different wavelengths of visible light. For example, **Figure 5.04** shows the change in the structure and colour of the indicator methyl orange in acidic and basic solutions.

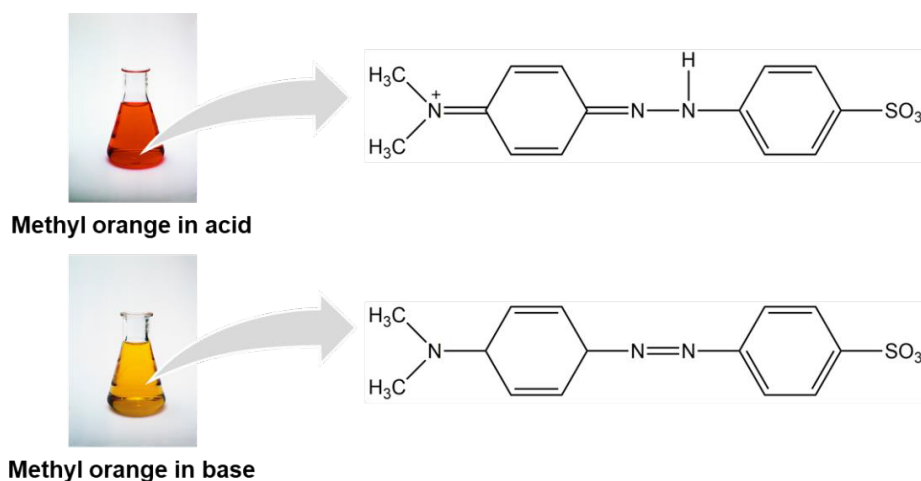


Figure 5.04: Molecular structure and colour of methyl orange in acid and base.

Acids can be classified as monoprotic or polyprotic, depending on the number of protons available for donation.

- Given the structural formula of an acid, classify it as monoprotic, diprotic, or triprotic.

Acids are compounds or ions containing one or more protons they transfer to bases in chemical reactions. Acids are classified as monoprotic or polyprotic, depending on the number of protons available for transfer.

Monoprotic Acids

Monoprotic acids have one proton available for transfer in a chemical reaction. **Figure 5.05** shows the structural formulae of some monoprotic acids with the proton available for transfer coloured red. Monoprotic acids can be identified from their molecular formulae. In most cases, monoprotic acids have only one hydrogen atom; the proton transferred in a reaction. However, in organic acids with multiple hydrogen atoms, the proton is typically part of a COOH group.

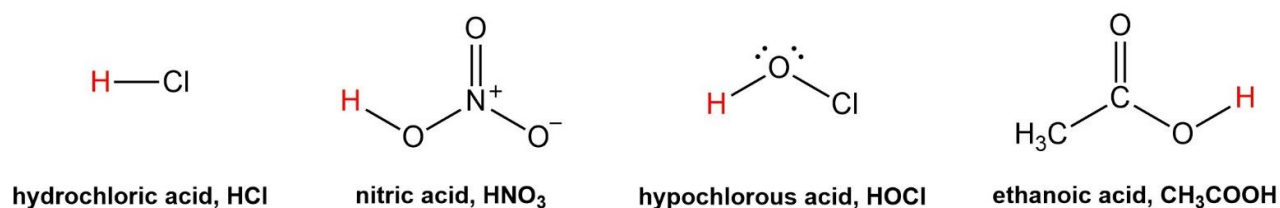
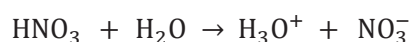


Figure 5.05: Structural formulae of several monoprotic acids.

Monoprotic acids ionise in one step when reacted with water. For example, the equation below describes the ionisation of nitric acid, HNO₃, in water.



Polyprotic Acids

Polyprotic acids have more than one proton available for donation in a chemical reaction. Polyprotic acids with two protons are classified as **diprotic**, and those with three are classified as **triprotic**.

Figure 5.06 shows the structural formulae of several polyprotic acids.

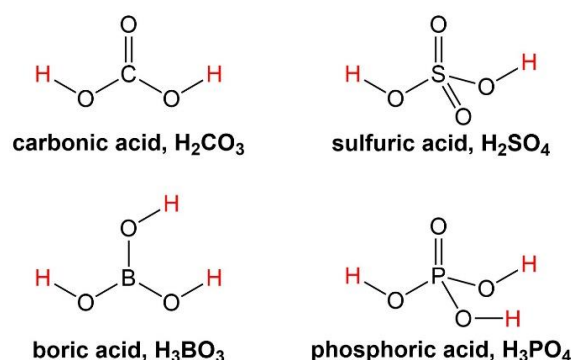


Figure 5.06: Structural formulae of several polyprotic acids.

Question 133

The photograph shows the reaction of water with a white solid named potassium oxide, K_2O .

- (a) Write a fully balanced chemical equation for the reaction.

_____ (2 marks) KA1

- (b) Phenolphthalein indicator solution was dissolved in the water.

Some information about phenolphthalein indicator is given below.

	Acidic							Basic						
pH	1	2	3	4	5	6	7	8	9	10	11	12	13	14
Phenolphthalein														

Describe and explain the colour change observed in the reaction.

(2 marks) KA2

Question 134

The photograph shows the reaction of water with phosphorus pentoxide, P_4O_{10} .

- (a) Write a fully balanced chemical equation for the reaction.

_____ (2 marks) KA1

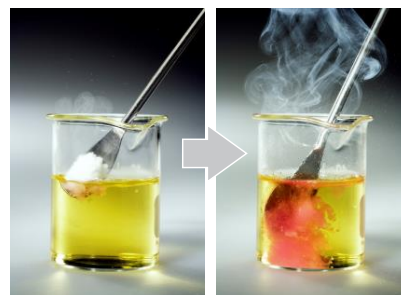
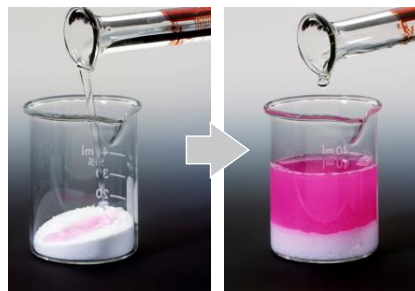
- (c) Methyl red indicator solution was dissolved in the water.

Some information about methyl red indicator is given below.

	Acidic							Basic						
pH	1	2	3	4	5	6	7	8	9	10	11	12	13	14
Methyl red														

Describe and explain the colour change observed in the reaction.

(2 marks) KA2

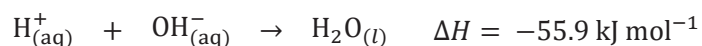


Similarities in the reactions of different acids with bases (metal oxides, hydroxides, and carbonates) allow products to be predicted from known reactants.

Neutralisation is an exothermic reaction.

- Identify the products obtained and write full and ionic equations for reactions between a given acid and a nominated metal oxide, hydroxide, carbonate, or hydrogencarbonate.
- Undertake stoichiometric calculations for reactions between acids and bases.

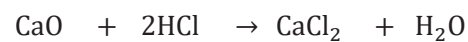
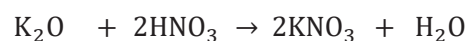
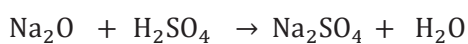
An acid-base reaction is termed **neutralisation** as the concentrations of acidic hydrogen ions, H^+ and alkaline anions, including O^{2-} , OH^- , CO_3^{2-} , and HCO_3^- are reduced to zero, effectively neutralising the acidity and alkalinity of the solution. Neutralisation is exothermic as energy is released when protons combine with hydroxide ions to form covalent bonds. The molar enthalpy of neutralisation, ΔH ranges from -55 to -64 kJ mol^{-1} depending on the concentrations of the acid and base.



Examples of neutralisation include the reactions of acids with metal oxides, metal hydroxides, metal carbonates, and metal hydrogencarbonates.

Reactions of Acids and Metal Oxides

Metal oxides contain oxide ions, O^{2-} , that accept protons from acids in a chemical reaction. The reaction of acids and metal oxides produces water and a salt, an ionic compound containing the metal oxide's cation and the acid's anion. Some reactions of acids with metal oxides are described in the fully balanced equations below.



The reactions of acids and metal oxides are also described using ionic equations. The acid is represented as H^+ in the ionic equation, and the metal oxide is represented by its empirical formula. As per the ionic equation, the two products are metal cations and water, as the acid's anions are excluded as spectator ions.

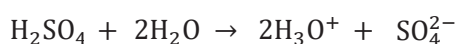


The strength of acids is explained by the degree of ionisation in an aqueous solution.

Acids are classified as weak and strong depending on their degree of ionisation in water.

Strong Acids

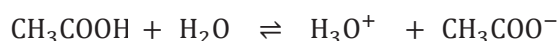
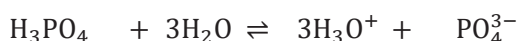
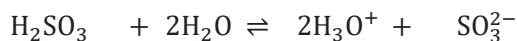
Strong acids ionise completely in water such that one mole of acid produces one mole of hydronium ions and one mole of the conjugate base. Examples of strong acids include hydrochloric acid, nitric acid and sulfuric acid.



A one-way arrow is used when writing equations to show the ionisation of a strong acid in water. The one-way arrow implies that the ionisation of the acid is complete.

Weak Acids

Weak acids partially ionise in water such that one mole of acid produces less than one mole of hydronium ions and conjugate base. Examples of weak acids include sulfurous acid, phosphoric acid, and carboxylic acids such as ethanoic acid.



A two-way arrow is used when writing equations to show the ionisation of a weak acid in water. The two-way arrow implies that the ionisation of the acid is incomplete. The table below shows some other examples of strong and weak acids.

Strong acids	Weak acids
hydrobromic acid, HBr	nitrous acid, HNO ₂
hydroiodic acid, HI	hydrofluoric acid, HF
perchloric acid, HClO ₄	hypochlorous acid, HOCl

The strength of an acid is tested by measuring the change in electrical conductivity when reacted with water. Strong acids produce a high concentration of ions, significantly increasing the solution's electrical conductivity. Conversely, weak acids produce a low concentration of ions, only slightly increasing the solution's electrical conductivity.

Question 142

Hydrochloric acid, HCl and hydrofluoric acid, HF are acids containing group 7 elements.

- (a) Hydrochloric acid is a strong acid, and hydrofluoric acid is a weak acid.

Explain the difference between strong and weak acids using the example of HCl and HF.

(2 marks) KA2

- (b) Write fully balanced equations for the ionisations of HCl and HF in water.

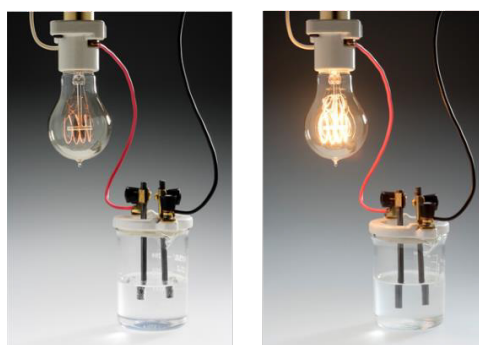
(2 marks) KA4

Question 143

An investigation was conducted in which two acids, A and B, were added to equal volumes of water.

The electrical conductivity was measured by connecting the solution to an electric circuit with a light bulb that glows when a current flows.

The photographs below show the results of the investigation.



Acid A

Acid B

- (a) Describe and explain the results of the investigation.

(2 marks) KA2

- (b) The chemical formula of acid A is H_3BO_3 , and the formula of acid B is HNO_3 .

Write fully balanced equations for the ionisations of acids A and B in water.

(2 marks) KA4

Calculating pH

The pH of an aqueous solution is calculated as the negative logarithm of the hydrogen ion concentration $[H^+]$ in moles per litre.

$$\text{pH} = -\log [H^+]$$

Hydrogen ions combine with water molecules in an aqueous solution, making pH equivalent to the negative logarithm of the hydronium ion concentration $[H_3O^+]$.

$$\text{pH} = -\log [H_3O^+]$$

Example 5.04

Pure water has a hydrogen ion concentration of $1.0 \times 10^{-7} \text{ mol L}^{-1}$.

Calculate the pH of pure water.

$$\begin{aligned} \text{pH} &= -\log [H^+] \\ \text{pH} &= -\log [1.0 \times 10^{-7}] \\ \text{pH} &= 7.0 \end{aligned}$$

Calculating $[H^+]$

The hydrogen ion or hydronium ion concentration of an aqueous solution is calculated using a rearrangement of the formula used to calculate pH.

$$[H^+] = 10^{-\text{pH}}$$

Example 5.05

The photograph shows an acidic solution with a pH of 2.86.



Calculate the solution's hydrogen ion concentration.

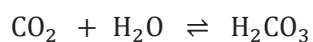
$$\begin{aligned} [H^+] &= 10^{-\text{pH}} \\ [H^+] &= 10^{-2.86} \\ [H^+] &= 1.38 \times 10^{-3} \text{ mol L}^{-1} \end{aligned}$$

CO₂ dissolves in rainwater to form carbonic acid, which is a weak acid, giving rainwater a pH of about 5.6.

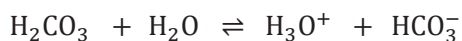
Oxides of sulfur and nitrogen in the atmosphere can produce rain with a pH below 5.6.

- Write equations for the reaction of CO₂ with water to produce hydrogen ions.
- Write equations for the reactions of oxides of sulfur and nitrogen with water that lead to acid rain.
- Examine the human activities that can cause acid rain to form and the strategies used to prevent this from happening.

The atmosphere contains carbon dioxide and water, which react to form carbonic acid.



Carbonic acid is a weak acid that partially ionises in clouds, releasing hydronium ions that lower the pH of rainwater to approximately 5.6.



Acid Rain

Acid rain is the term used to describe rain, snow, or fog with a pH lower than 5.6. Acid rain is primarily caused by sulfur oxide (SO_x) and nitrogen oxide (NO_x) emissions from human sources, including fossil fuel combustion for electricity production and vehicle transportation. In addition, sulfur and nitrogen oxides are emitted in metal production, which extracts metal ores and processes them by smelting, reduction or electrolysis into pure metals and alloys. **Figure 5.12** shows the primary sources of the emissions of sulfur and nitrogen oxides.

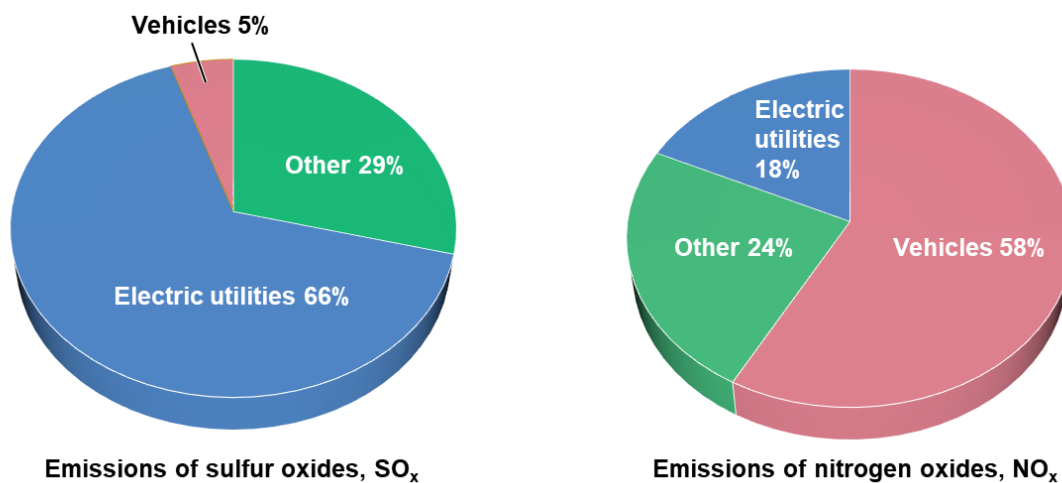


Figure 5.12: Primary sources of SO_x and NO_x emissions in the USA in 2017.

The oxides of sulfur and nitrogen react with water in the atmosphere forming acids that reduce the pH of rainfall to a level below 5.6.

Sulfur Oxides SO_x

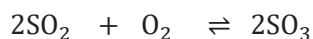
Sulfur oxides, including sulfur dioxide, SO₂, and sulfur trioxide, SO₃, are emitted into the atmosphere from human sources, primarily the combustion of fossil fuels and the smelting of metal ores. Fossil fuels including coal, petroleum and natural gas contain small amounts of sulfur released into the atmosphere when these fuels are combusted in power stations and motor vehicles to release heat for electrical energy production and transport.



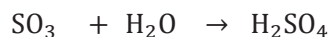
In addition, some mineral ores, including copper sulfides, iron sulfides and zinc sulfides, contain sulfur that is released into the atmosphere in the form of sulfur dioxide. When metal sulfides are processed. For example, the equation below shows the formation of sulfur dioxide when producing iron metal, Fe, from the mineral pyrite, FeS₂.



Sulfur dioxide may react with oxygen in the atmosphere forming sulfur trioxide.

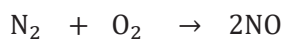


Both sulfur dioxide and sulfur trioxide react with water vapour in the atmosphere producing sulfurous and sulfuric acids that lower the pH of rainfall to a level below 5.6.

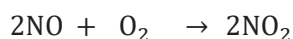


Nitrogen Oxides NO_x

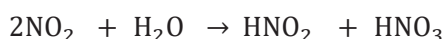
Nitrogen oxides, including nitric oxide, NO and nitrogen dioxide, NO₂, are emitted into the atmosphere primarily from motor vehicles and electrical power stations where fossil fuels are combusted to release heat. For example, Nitric oxide forms inside a vehicle's internal combustion engine and a power station's boilers from the reaction of nitrogen and oxygen in the air.



Nitric oxide reacts with oxygen in the atmosphere and produces nitrogen dioxide.



Nitrogen dioxide reacts with water vapour in the atmosphere producing nitric and nitrous acids that lower the pH of rainfall to a level below 5.6.



Preventing Acid Rain

The primary causes of acid rain globally are the emissions of sulfur and nitrogen oxides from human sources. **Figure 5.13** shows the changes in global emissions of sulfur and nitrogen oxides between 1930 and 1980.

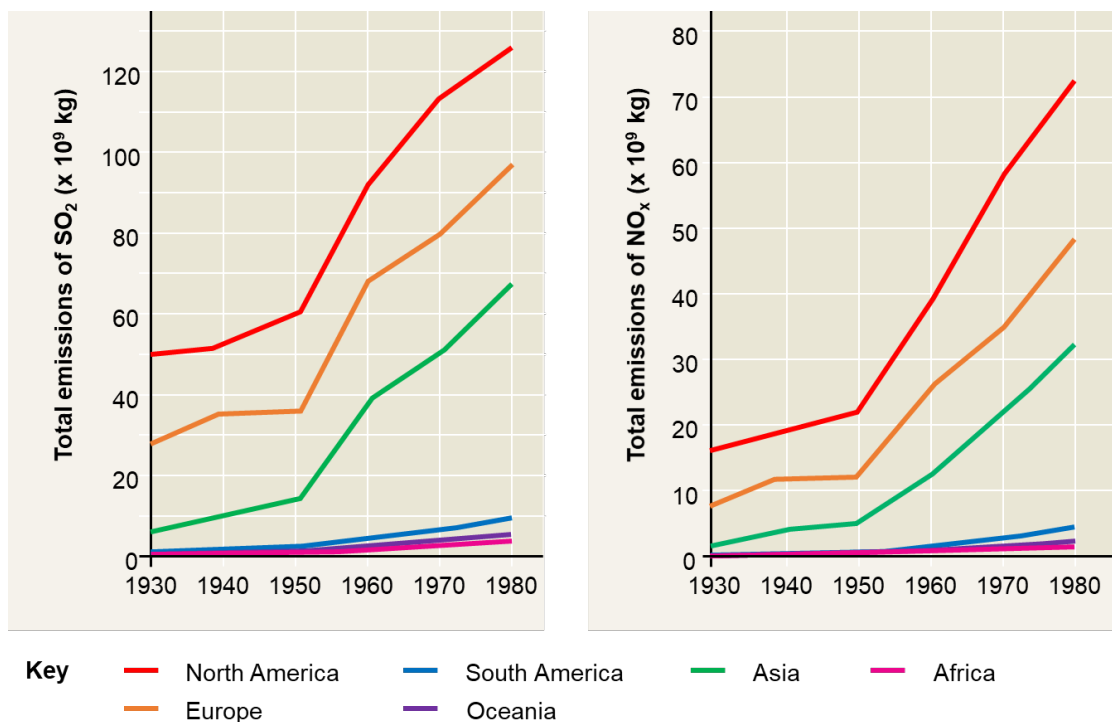


Figure 5.13: Global emissions of sulfur and nitrogen oxides between 1930 and 1980.

The increase in global emissions of sulfur and nitrogen oxides in **Figure 5.13** was primarily due to the increase in consumption of fossil fuels that accompanied the increase in the human population. However, more recently, there has been a decline in global emissions due to increased air quality monitoring, the implementation of new environmental laws and international regulations, and the development of technologies that minimise emissions from vehicles and power stations. For example, a catalytic converter (**Figure 5.14**) is a device connected to a motor vehicle's exhaust system. The catalytic converter contains metals including platinum and rhodium that catalyse the conversion of nitric oxide to nitrogen, reducing the emissions of nitrogen oxides from motor vehicles. The conversion is described in the equation below.

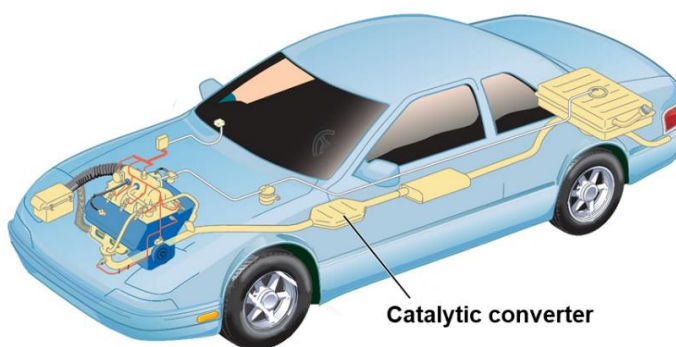
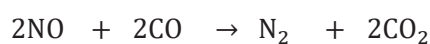


Figure 5.14: Catalytic converter



Acid rain changes the chemical composition of lakes, rivers, and soils, affecting plant and animal growth. Acids react with water in lakes, rivers and soils, releasing hydronium ions that lower the pH of these ecosystems and alter their biodiversity. For example, the decrease in pH in lakes interferes with cellular processes in fish, crayfish, and clams, reducing their population sizes over time.



In soils, hydrogen ions are exchanged with aluminium ions bound to soil particles. This cation exchange releases soluble aluminium cations into the soil solution, as shown in the equation below.

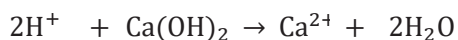


Aluminium ions are highly toxic to plants and animals in an aqueous form. Aluminium ions shorten plant roots, reducing their surface-area-to-volume ratio for nutrient, mineral and water absorption. Furthermore, aqueous aluminium ions are released from soils into lakes and rivers, damaging fish gills and impairing respiration. To combat the adverse effects of aluminium toxicity, scientists add bases to neutralise acids in lakes, rivers, and soils. **Figure 5.16** shows two common methods that neutralise the acidity of lakes and soils.



Figure 5.16: Neutralising acidity in lakes (left) and soils (right).

The image on the left in **Figure 5.16** shows a dosing column that adds calcium hydroxide to lake water which neutralises acidity according to the ionic equation below.



The image on the right in **Figure 5.16** shows the addition of powdered calcium carbonate to soils which neutralises acidity according to the ionic equation below.



Both methods successfully raise the pH of soils, lakes and rivers, leading to a more pH neutral environment within the tolerance limits of the living things in that environment.

Review Test 5

Question 1

Lactic acid, $\text{CH}_3\text{CH}(\text{OH})\text{COOH}$, is a weak acid produced in muscle cells during strenuous exercise.

- (a) State the property of lactic acid that allows it to behave as an acid.

_____ (1 mark) KA1

- (b) State why lactic acid is classified as a weak acid.

_____ (1 mark) KA1

- (c) Lactic acid ionises in muscle cells and blood producing the lactate ion, $\text{CH}_3\text{CH}(\text{OH})\text{COO}^-$.

- (1) Write an equation to show the ionisation of lactic acid in water.

_____ (1 mark) KA1

- (2) Classify lactic acid as monoprotic or polyprotic.

_____ (1 mark) KA1

- (d) Lactic acid production during exercise results in a burning sensation in active muscles, which causes pain and prevents the body from overworking.

In 2018, several horse trainers and stablehands were found guilty of doping horses with sodium hydrogencarbonate on race day.

- (1) Write an ionic equation for the reaction of lactic acid and sodium hydrogencarbonate.

_____ (1 mark) KA1

- (2) Describe and explain why horses were doped with sodium hydrogencarbonate.

_____ (2 marks) KA2

- (3) Scientists detect doping by measuring the total carbon dioxide (TCO_2) concentration in venous blood samples.

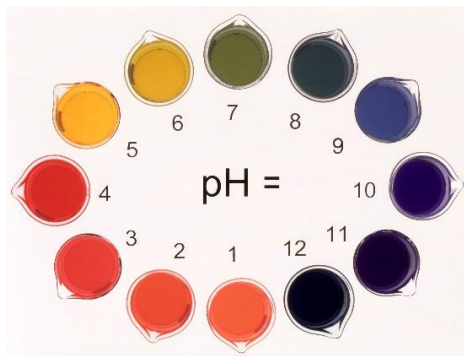
Suggest a limitation of this method of detection.

_____ (1 mark) KA3

Question 2

Universal indicator is a mixture producing a spectrum from red to violet as pH increases.

The photograph below shows the colour of universal indicator at different pH.



- (a) Magnesium oxide, MgO , was added to water containing universal indicator, and the pH increased from 7 to 10.

(1) Write an equation to show how magnesium oxide increases the pH of water.

_____ (2 marks) KA1

(2) State the colour change in the reaction.

_____ (1 mark) KA2

- (b) Phosphorus pentoxide, P_4O_{10} , was added to water containing universal indicator, and the colour changed from green to orange.

(1) Write an equation for the reaction.

_____ (2 marks) KA1

(2) Draw the structural formula of the reaction product that changes the colour of universal indicator from green to orange.



(2 marks) KA1

- (c) The pH 1 solution was prepared using nitric acid, HNO_3 and the pH 12 solution was prepared using potassium hydroxide, KOH .

The pH 1 solution was neutralised with the pH 12 solution.

(1) Write a fully balanced chemical equation for the neutralisation reaction.

_____ (1 mark) KA2

(2) State whether temperature increases or decreases in the reaction.

_____ (1 mark) KA2

Question 4

The Australian bushfires in 2020 and 2021 released large quantities of nitrogen oxides into the atmosphere that resulted in acidic rainfall.

- (a) Define acidic rainfall.

_____ (1 mark) **KA1**

- (b) Use fully balanced equations to describe the formation of nitrogen oxides and acidic rainfall.

_____ (6 marks) **KA1**

- (c) Extensive scientific research in the 1960s and 1970s in Europe and the USA revealed that nitrogen oxide emissions were posing significant risks to human health, and acid rain was damaging aquatic and terrestrial ecosystems.

In 1979, leaders from 32 countries in the Northern hemisphere signed the UNECE Convention on Air Pollution to solve this problem. The Convention outlined general principles of international cooperation for reducing air pollution and setting up an institutional framework for international collaboration between scientists and policymakers.

The Convention has contributed significantly to developing international environmental law and producing a framework for controlling and reducing damage to human health, reducing nitrogen oxide emissions by 40 to 80% since 1990.

The success of the Convention in reducing air pollution resulted from the development of a common knowledge base, including a scientific infrastructure aimed at monitoring and modelling programs by an international network of scientists of various disciplines. In addition, the Convention has provided a platform for scientists and policymakers to exchange information which has developed mutual trust and led to innovative approaches.

One key concept of science as a human endeavour is 'communication and collaboration'.

Discuss two examples in the text above that illustrates this key concept.

_____ (4 marks) **KA3**

6.1: Concepts of oxidation and reduction

A range of reactions, including reactions of metals, combustion, and electrochemical processes, can be considered redox reactions.

Oxidation and reduction can be defined in terms of combination with oxygen, transfer of electrons, or change in oxidation number.

- Identify oxidation and reduction in given equations.
- Write oxidation and reduction half-equations, in neutral and acidic conditions, given reactant and product species.
- Combine half-equations to write a chemical equation.
- Determine the oxidation states of atoms in elements and monatomic ions, and in compounds and polyatomic ions.

A **redox reaction** is a chemical reaction in which electrons are transferred from one reactant to another. **Redox** is a combination of reduction and oxidation, chemical reactions defined in terms of combination with oxygen, transfer of electrons, or change in oxidation number.

Combination with Oxygen

The concepts of oxidation and reduction were introduced by French scientist Antoine Lavoisier (**Figure 6.01**) in the late 18th century. Lavoisier was a passionate chemist who investigated a wide range of chemical reactions, including substances with oxygen. Lavoisier discovered that substances reacted with oxygen to form oxides and introduced the term **oxidation** to describe a gain of oxygen in a chemical reaction. For example, hydrogen is oxidised to water according to the equation below.

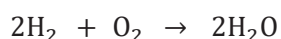


Figure 6.01: Antoine Lavoisier

Lavoisier had the highest precision measuring equipment and carefully analysed the masses of reactants and products produced in the reactions he investigated. Lavoisier noted the mass of an oxide or oxygen-containing compound decreased when heated to a high temperature and introduced the term **reduction** to describe this reduction in mass. For example, when mercury(II) oxide is heated above 500°C, as in **Figure 6.02**, the oxide decomposes, reducing mercury oxide to mercury and releasing oxygen that fills the balloon. The reduction of mercury(II) oxide to mercury is described according to the equation below.



Figure 6.02: Reduction of HgO

Oxidation can be defined as a gain of oxygen, and reduction can be defined as a loss of oxygen.

Electron Transfer

Oxidation and reduction are also defined in terms of electron transfer in a chemical reaction. In a redox reaction, one substance transfers electrons to one or more other substances. The substance that loses electrons is oxidised, and any substance that gains electrons is reduced.

Oxidation is a loss of electrons, and reduction is a gain of electrons.

A redox chemical reaction can be divided into two **half-reactions** that separately describe oxidation and reduction processes. Each half-reaction is described using an equation, called a **half-equation**, that identifies the number of electrons lost or gained by a substance in a reaction. The term half-equation is used in reference to the half equations describing half a redox reaction.

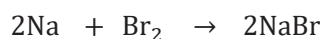
Example 6.01

Figure 6.03 shows the reaction of sodium and bromine.

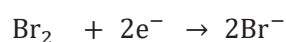


Figure 6.03: Reaction of sodium and bromine.

The reaction is described in the equation below.



Bromine is more electronegative than sodium and gains electrons in the reaction. Each bromine atom in Br_2 gains one electron, forming two bromide ions according to the half-equation below. The electrons appear on the left-hand side of the arrow in the half-equation, indicating that bromine molecules gain electrons and are reduced in the reaction.



Sodium has a lower electronegativity than bromine and loses electrons to bromine molecules in the reaction. Each sodium atom loses one electron, forming sodium cations according to the half-equation below. The electrons appear on the right-hand side of the arrow in the half-equation, indicating that sodium atoms lose electrons and are oxidised in the reaction.



Writing Half-Equations in Acidic Conditions

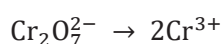
Some redox reactions will only occur in a low pH environment. In such reactions, the protons, H^+ in the acidic solution gain oxygen atoms from the substance being reduced in the reaction. For example, the half-equation below shows the reduction of permanganate ions (MnO_4^-) to manganese ions (Mn^{2+}) in acidic conditions. Protons gain oxygen atoms from permanganate ions, reducing them to manganese ions according to the half-equation below.



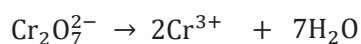
A method of writing half-equations in acidic conditions is described in [Example 6.04](#).

Example 6.04

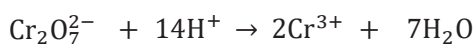
Ethanol, $\text{C}_2\text{H}_5\text{OH}$, is oxidised by an acidic solution containing dichromate ions, $\text{Cr}_2\text{O}_7^{2-}$. Dichromate ions are reduced to chromium(III) ions according to the unbalanced half-equation below.



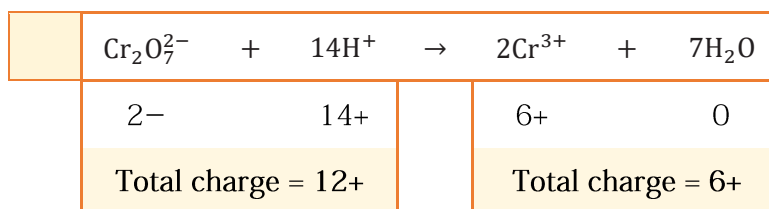
The first step in balancing the redox half-equation in acidic conditions is to balance oxygen by adding water molecules to the side of the arrow with fewer oxygen atoms. In this example, seven water molecules are added to the right-hand side of the arrow to balance oxygen.



The second step is to balance hydrogen by adding protons, H^+ , to the side of the arrow with fewer hydrogen atoms. In this example, 14 protons are added to the left of the arrow to balance hydrogen.



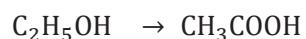
The third and final step is to balance electric charge by adding electrons to the side of the arrow with the greater positive charge. The charges on both sides of the arrow have been identified below.



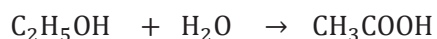
In this example, six electrons are added to the left-hand side of the arrow to balance the electric charge on both sides of the arrow in the equation.



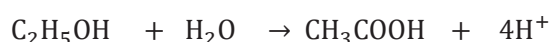
The oxidation of ethanol in the reaction is described in the unbalanced half-equation below.



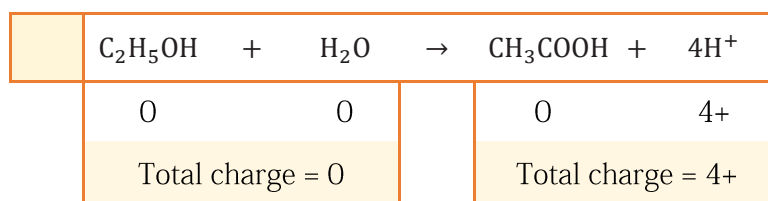
The first step in balancing the half-equation in acidic conditions is to balance oxygen by adding water molecules to the side of the arrow with fewer oxygen atoms. In this example, one water molecule is added to the left-hand side of the arrow to balance oxygen.



The second step is to balance hydrogen by adding protons to the side of the arrow with fewer hydrogen atoms. In this example, four protons are added to the right of the arrow to balance hydrogen.



The third and final step is to balance the electric charge by adding electrons to the side of the arrow with the greater positive charge. The charges on both sides of the arrow have been identified below.



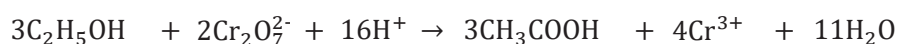
In this example, four electrons are added to the right-hand side of the arrow.



The half-equations are combined when writing the overall reaction equation. The first step is to ensure that the half-equations have the same number of electrons as shown below.



Protons, water molecules and electrons appear on both sides of the arrow, and these are cancelled out when combining the half-equations. Cancelling is done by removing all protons and water molecules from the side of the arrow with the least number of each. For example, three water molecules are removed from both sides to cancel water from the left-hand side of the arrow, and 12 protons are removed from both sides to cancel protons from the right-hand side of the arrow in this equation. The overall equation for this redox reaction is given below.



Metal Displacement Reaction

When a reactive metal is placed in contact with a compound containing ions of a less reactive metal, the reactive metal atoms transfer electrons to the less reactive metal ions in the compound. As a result, the reactive metal is oxidised, forming cations and the less reactive metal ions are reduced, forming atoms. This redox reaction is called a **metal displacement reaction**, as one element replaces another in a compound.

Example 6.05

Figure 6.04 shows the reaction of aluminium, Al, and solution containing copper ions, Cu^{2+} .

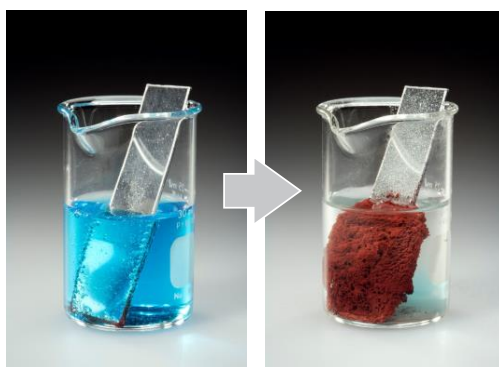
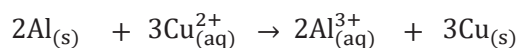
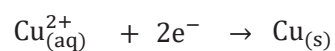
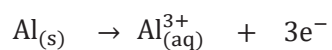


Figure 6.04: Reaction of aluminium and a solution of copper ions.

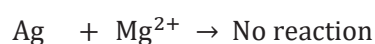
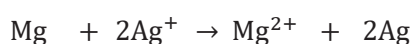
Aluminium is more reactive than copper, and electrons are transferred from aluminium atoms to copper ions, displacing copper from the solution according to the equation below.



The aluminium atoms in the metal strip transfer electrons to the copper ions in the solution, forming copper atoms and aluminium ions. The aluminium ions dissolve in the solution, and the copper atoms assemble into a solid lattice seen as a red mass on the surface of the metal strip. In addition, the solution changes from blue to colourless as the blue-coloured copper ions are reduced to copper.

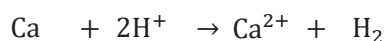
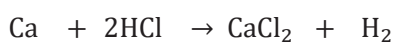


A metal displacement reaction will only occur when atoms of the more reactive metal are reacted with cations of the less reactive metal. For example, magnesium will displace silver as magnesium is more reactive than silver, but silver cannot displace magnesium.

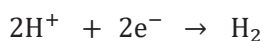
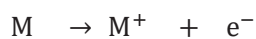


The Reaction of Metals and Dilute Acids

When a reactive metal is reacted with dilute acid, the metal atoms transfer electrons to the protons, H^+ in the acidic solution, forming hydrogen molecules, H_2 . This is another type of displacement in which a reactive metal displaces hydrogen from an acidic solution. For example, calcium is a reactive metal that reacts with hydrochloric acid according to the fully balanced and ionic equations below.



The reaction of a reactive metal with dilute acid is a redox reaction in which the metal atoms are oxidised, and the protons are reduced according to the half equations below.



The reactivity of metals is determined from their reaction with dilute acids. **Figure 6.07** shows the reactions of four different metals with a dilute hydrochloric acid solution.

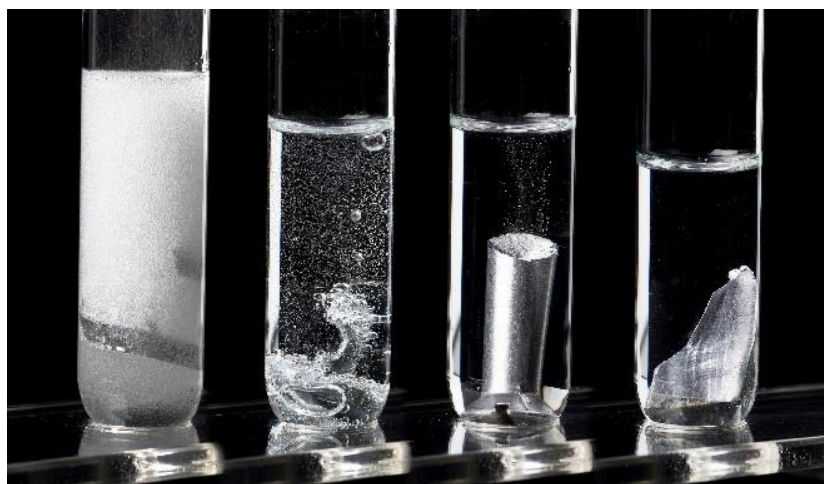


Figure 6.07: Reactions of metals with hydrochloric acid (left to right: magnesium, zinc, iron, lead).

The following conclusions are drawn from the diagram:

1. Hydrogen gas forms rapidly when magnesium is reacted with the hydrochloric acid solution, indicating that magnesium is the most reactive of the four metals in **Figure 6.07**.
2. Iron reacts less vigorously with the hydrochloric acid solution than magnesium and zinc, indicating that iron is less reactive than these two metals.
3. No hydrogen gas is produced when lead is placed in hydrochloric acid solution, indicating that lead is the least reactive of the four metals in **Figure 6.07**.

6.3: Electrochemistry

Electrochemical reactions involve a flow of electrons during a chemical reaction.

Galvanic cells produce electrical energy from spontaneous redox reactions.

- Identify the anode and cathode and their charges, and the direction of ion and electron flow, in a galvanic cell, given sufficient information.
- Draw a diagram of a galvanic cell, given sufficient information.
- Write electrode half-equations for a galvanic cell, given sufficient information.

Electrochemical reactions are redox reactions either caused or accompanied by the flow of electrons between the reacting substances. When the electrochemical reactions proceed in contact with two electric conductors, separated by conducting wires, in a suitably organised electrochemical cell, electrons flow through the wires, producing an electric current. This electric current can be used to power electronic devices and is the basis for the operation of a battery.

Galvanic Cell

A **galvanic cell** (Figure 6.08) is an electrochemical cell that generates a flow of electrons from spontaneous redox reactions. The cell contains two electrolyte solutions, each containing ions that participate in either oxidation or reduction reactions. Each electrolyte contains an electrode, an electrical conductor that serves two functions in the cell. The first is to facilitate redox reactions occurring at their surface. The second is to carry charge between the electrolyte and the external circuit. For this reason, electrodes are composed of electrically conductive materials, such as metals and graphite. Together, the electrolyte and electrode are called a **half-cell**. Oxidation occurs at the electrode surface in one half-cell, releasing electrons that flow through conducting wires to the other half-cell, where they are gained at the electrode surface by cations or molecules in the electrolyte. The two half cells are connected by a **salt bridge**, a medium containing a soluble salt solution. The electrode reactions alter the ionic composition of the electrolytes, producing a charge imbalance that is neutralised by the flow of ions from the salt bridge.

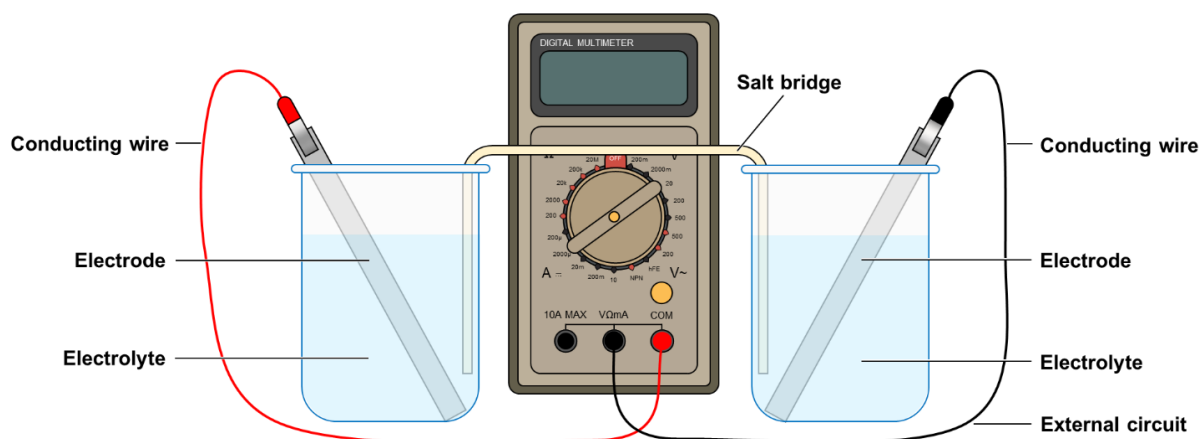


Figure 6.08: Components of a simple galvanic cell.

Galvanic Cell Operation

A galvanic cell contains two half-cells, each comprised of an electrode and electrolyte, connected to an external circuit via conducting wires. The electrodes have different functions in different galvanic cell types. In some galvanic cells, the electrodes are part of the redox reaction, and in others, they act as a conductive medium to facilitate electron transfer. One type of galvanic cell, described in **Example 6.06**, has two dissimilar metal electrodes, each suspended in a solution of their common cations. The two metals have different chemical reactivity, with the more reactive of the two being oxidised more readily. The more reactive of the two metals is the negative electrode, called the **anode**, and the less reactive metal is the positive electrode, called the **cathode**. When placed in contact with their electrolytes, oxidation occurs at the anode surface, releasing electrons that flow to the cathode, reducing substances at its surface.

Example 6.06

Figure 6.09 shows a simple galvanic cell composed of magnesium and copper half-cells.

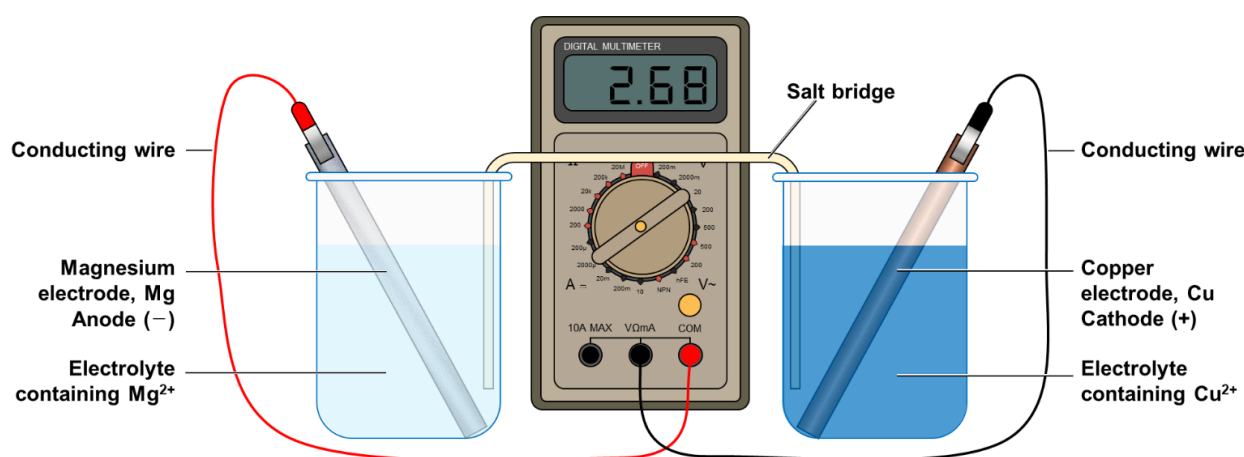
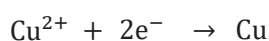


Figure 6.09: Components of a magnesium-copper galvanic cell.

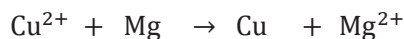
The magnesium electrode is the anode in this galvanic cell as magnesium is more reactive than copper. Therefore, the magnesium atoms in the electrode are oxidised according to the half-equation below. As a result, the magnesium ions, Mg^{2+} , dissolve in the anode electrolyte, reducing the mass of the magnesium electrode, and increasing the positive charge of the anode electrolyte.



The electrons flow through conducting wires to the cathode half-cell, where they reduce copper ions, Cu^{2+} at the cathode surface according to the half-equation below. As a result, the copper ions are transformed into copper atoms, forming solid copper that increases the mass of the copper electrode.



The overall cell reaction is described in the equation below.



The formation of magnesium ions at the anode increases the magnitude of positive charge in the anode electrolyte, and the removal of copper ions at the cathode reduces the magnitude of positive charge in the cathode electrolyte. The positive charge in the anode electrolyte and negative charge in the cathode electrolyte are neutralised by the flow of ions from the salt bridge. The salt bridge contains high concentrations of dissolved cations and anions that flow to the different half-cells to neutralise this charge imbalance. Anions from the salt bridge flow to the anode to neutralise the positive charge, and cations from the salt bridge flow to the cathode to neutralise the negative charge. The flow of ions from the salt bridge to the electrolytes completes the electric circuit, ensuring electron flow can continue within the conducting wires (Figure 6.10).

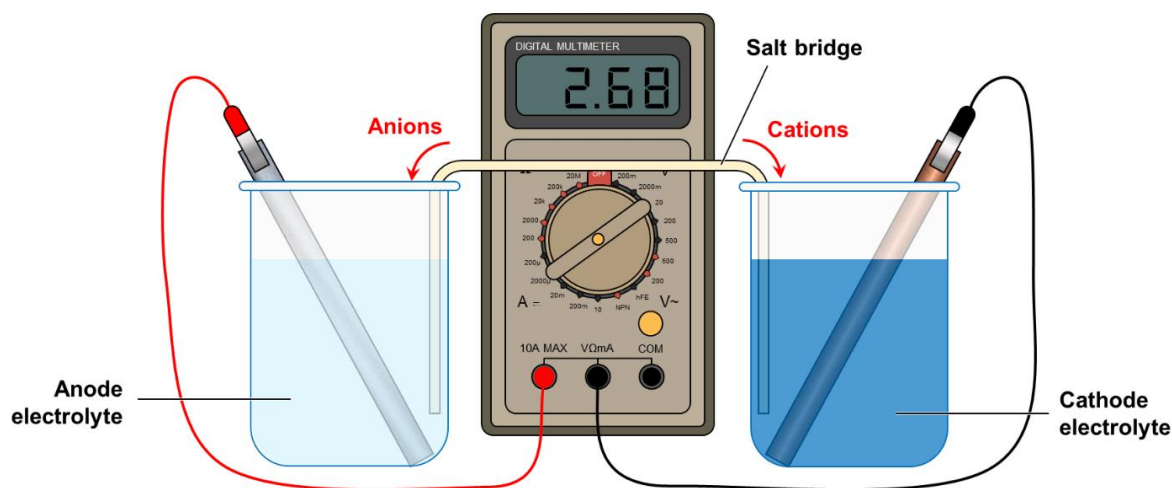


Figure 6.10: Flow of ions from the salt bridge.

The electrode reactions produce a potential difference that is measured with a voltmeter. The potential difference drives electron flow in the external circuit, producing a current that is measured with an ammeter. The properties of a galvanic cell composed of half-cells containing two different metals are summarised in the table below.

Component	Summary
Anode	Negative electrode composed of the more reactive metal.
Cathode	Positive electrode composed of the less reactive metal.
Anode reaction	Oxidation.
Cathode reaction	Reduction.
Electron flow	Anode to cathode.
Ion flow from the salt bridge	Anions to anode and cations to cathode.

Galvanic cells are commonly used as portable sources of electric current.

- Compare the operation of different types of batteries.

A **battery** assembles two or more galvanic cells into a power supply capable of transforming chemical energy into electrical energy. The primary function of batteries is to store and release electrical energy to power electronic devices. Smaller batteries can power low-voltage devices, while larger batteries can power cities. The oxidation and reduction reactions occurring at the electrodes in a battery produce an electric potential difference that drives the flow of electrons from anode to cathode through an electronic device. The operation of some non-rechargeable batteries, called **primary cells**, are described in **Examples 6.09, 6.10, and 6.11**.

Example 6.09

Figure 6.14 shows the structure and components of a mercury battery.

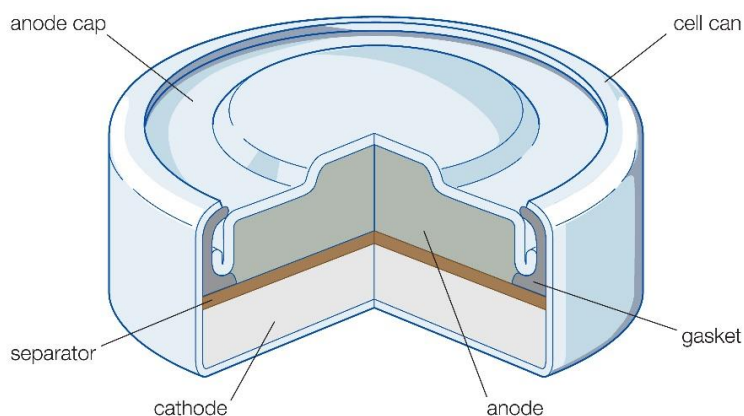
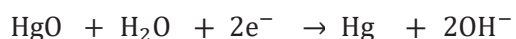


Figure 6.14: Components of a mercury battery.

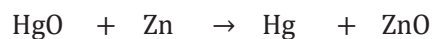
The anode is composed of zinc separated from the cathode by an electrolyte containing hydroxide ions. Zinc is oxidised at the anode according to the half-equation below.



The cathode is a mixture of mercury(II) oxide with powdered graphite to increase the electrical conductivity. Mercury(II) oxide is reduced at the cathode according to the half-equation below.



The overall cell reaction is:



The overall cell reaction produces a potential difference of approximately 1.35 volts between the anode and cathode, which can power watches, calculators, and hearing aids.

Example 6.13

Figure 6.19 shows the components of a lead-acid battery.

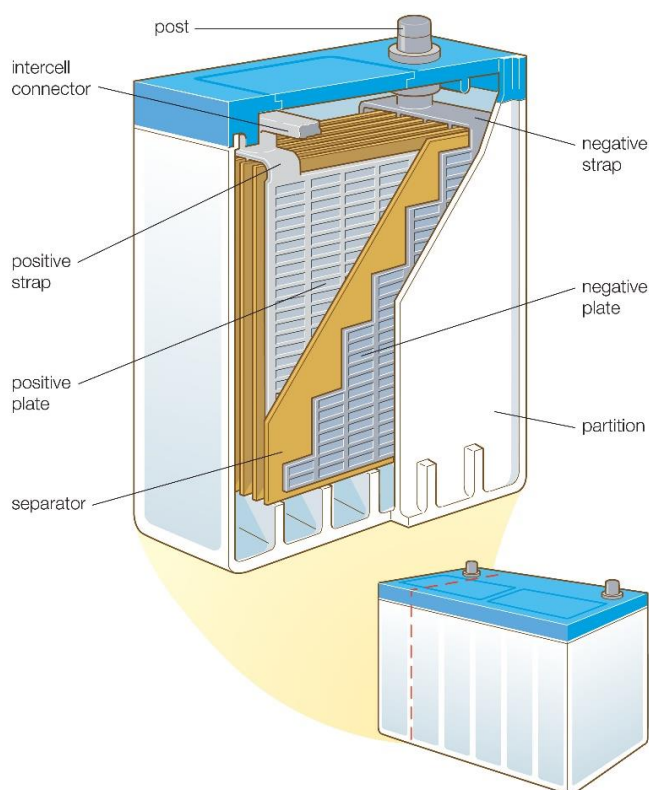
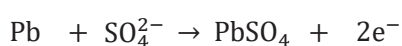


Figure 6.19: Components of a lead-acid battery.

The anode is composed of lead that is in contact with the sulfuric acid electrolyte. Lead is oxidised when the cell is discharging according to the half-equation below.



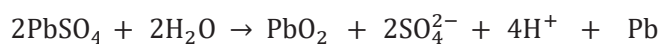
The cathode is composed of lead(IV) oxide in contact with sulfuric acid. Lead(IV) oxide is reduced when the cell discharges according to the half-equation below.



The overall reaction when the cell is discharging is:



The anode and cathode reactions are reversed when the cell is recharging. The overall reaction when the cell is recharging is:

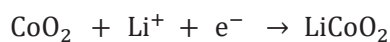
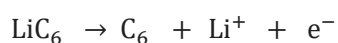


Question 174

The diagram below shows a lithium-ion battery that is used in smartphones.



The half-equations below describe the oxidation and reduction reactions when the cell discharges.



(a) State the charge on the anode when the cell is discharging.

_____ (1 mark) KA1

(b) Write an equation for the overall reaction when the cell is discharging.

_____ (2 marks) KA1

(c) Over-discharging leads to the formation of insoluble lithium and cobalt oxides.



(1) State how this reaction reduces the effectiveness of the battery over time.

 _____ (1 mark) KA2

(2) State a simple method of preventing the reaction above from occurring.

_____ (1 mark) KA2

(d) Lithium-ion batteries are rechargeable.

(1) Write an equation for the overall reaction when the cell is recharging.

_____ (2 marks) KA1

(2) State the charge on the cathode when the cell is recharging.

_____ (1 mark) KA1



**SOLUTIONS TO
CHAPTER QUESTIONS
AND REVIEW TESTS**

Question	Part	Author's response	Marks
1	(a)	(1) The temperature at which a material changes from solid to liquid.	1
		(2) To ensure that the material does not melt at normal operating temperature.	1
	(b)	(1) The high thermal conductivity maximises the rate of conductive heat transfer from the cylinders to the coolant.	1
		(2) The high boiling point ensures water remains liquid inside the engine, allowing it to circulate, absorb heat and prevent the engine from overheating.	1
2	(a)	Thermal insulators reduce the rate of conductive heat transfer between the interior and exterior of a building.	1
		This minimises the costs of heating and cooling the building to maintain a habitable temperature for the occupants.	1
	(b)	Polyurethane.	1
		Greatest amount of insulation per unit area of the material.	1
	(c)	The low boiling point ensures isopentane forms a gas readily when heated.	1
		The heat from the reaction readily boils and expands isopentane, allowing it to function as an effective blowing agent.	1
3	(a)	(1) Metals such as copper are good electrical conductors as these materials contain free-moving electrons.	1
		The electrons in copper flow through the wire and toaster when the wire is connected to a source of electric potential difference such as the wall-outlet.	1
		Any one:	1
	(b)	(2) To prevent electric shock when handling the wire.	
		To prevent charge from leaking away from the device during operation.	
		To prevent the wire from corroding over time.	
(b)	High electrical conductivity allows electrons to flow through the heating element, transforming electricity into heat.	1	
	High thermal conductivity ensures that heat flows rapidly from the heating element to the bread when using the device.	1	
4	(a)	(1) A water-soluble plastic dissolves in water when the water temperature is sufficiently high.	1
		(2) The PVOH bag dissolves and releases the contaminated clothing and bedding directly into the detergent-filled washing machine.	1
		This practice minimises contact between hospital workers and the contaminated clothing and bedding, reducing the possibility of transmitting infection.	1
	(b)	(3) Detergents must dissolve in water and form a solution to remove dirt, stains, bacteria, and bodily fluids from contaminated clothing.	1
		PGA and PLA sutures dissolve naturally over time as the wound heals.	1
		Using PGA and PLA save the surgeon and patient time and money, minimising the risk of infection associated with the follow-up surgery to remove the sutures.	1

172	(a)	(1)	$\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$	2	
		(2)	Anode	1	
	(b)	(1)	+4 to +3	1+1	
		(2)	$\text{Zn} + 2\text{MnO}_2 + 2\text{NH}_4^+ \rightarrow \text{Zn}^{2+} + 2\text{MnO}(\text{OH}) + 2\text{NH}_3$	2	
	(c)	Graphite is an electrical conductor which increases the conductivity of the manganese dioxide (positive) electrode.		1	
	(d)	Any one:		1	
The potential difference produced by the electrode reactions is limited, limiting the output power. Primary cells can only be used once and not recharged. Therefore, they must be replaced more frequently at the consumer's expense. The raw materials often include highly corrosive alkaline substances that are hazardous to consumers if stored incorrectly or misused.					
173	(a)	Zn		1	
	(b)	Ag and OH^-		1	
	(c)	$\text{Zn} + \text{Ag}_2\text{O} \rightarrow \text{ZnO} + 2\text{Ag}$		2	
	(d)	Powdered graphite is an electrical conductor which increases the conductivity of the silver oxide cathode.		1	
174	(a)	Negative		1	
	(b)	$\text{LiC}_6 + \text{CoO}_2 \rightarrow \text{C}_6 + \text{LiCoO}_2$		2	
	(c)	(1)	The insoluble oxides are poor electrical conductors, which reduces the anode and cathode's conductivity and the cell reactions' rate.		1
		(2)	By charging the battery before it goes flat.		1
	(d)	(1)	$\text{C}_6 + \text{LiCoO}_2 \rightarrow \text{LiC}_6 + \text{CoO}_2$		2
		(2)	Negative		1
175	(a)	PbO_2		1	
	(b)	Positive		1	
	(c)	$\text{PbO}_2 + 2\text{SO}_4^{2-} + 4\text{H}^+ + \text{Pb} \rightarrow 2\text{PbSO}_4 + 2\text{H}_2\text{O}$		2	
	(d)	Water is formed in the discharging reaction, which freezes when the temperature is below zero in winter.		1	
	(e)	(1)	$2\text{PbSO}_4 + 2\text{H}_2\text{O} \rightarrow \text{PbO}_2 + 2\text{SO}_4^{2-} + 4\text{H}^+ + \text{Pb}$		2
		(2)	Positive		1

Solutions: Review Test 1				
Question	Part	Author's response	Marks	
1	(a)	Homogeneous mixtures are those in which the components are mixed uniformly such that all parts have the same chemical composition.	1	
		Seawater is a uniform mixture with all parts containing the same composition of the different elemental salts and water.	1	
	(b)	A sample of seawater is transferred to a reaction flask connected to a condenser.	1	
		Heat is transferred to the mixture until all the water has evaporated, leaving behind solid elemental salts that are collected and purified.	1	
	(c)	(1)	$n = \frac{m}{M}$	
			$n = \frac{11 \times 10^{-3}}{22.99}$	1
			$n = 4.78 \times 10^{-4} \text{ mol}$	1
		(2)	$N_{\text{Cl}^-} = n_{\text{Cl}^-} \times N_A$	
			$n_{\text{Cl}^-} = \frac{N_{\text{Cl}^-}}{N_A}$	
			$n_{\text{Cl}^-} = \frac{3.22 \times 10^{20}}{6.02 \times 10^{23}}$	1
			$n_{\text{Cl}^-} = 5.35 \times 10^{-4} \text{ mol}$	1
			$m_{\text{Cl}^-} = n_{\text{Cl}^-} \times M_{\text{Cl}^-}$	
			$m_{\text{Cl}^-} = 5.35 \times 10^{-4} \times 35.45$	1
			$m_{\text{Cl}^-} = 1.90 \times 10^{-2} \text{ g (0.0190)}$	1
	(d)	(1)	Isotopes are atoms of the same chemical element with different mass numbers due to differences in neutron numbers.	1
			Bromine-79 and bromine-81 are bromine atoms with neutron numbers of 44 and 46, respectively.	1
		(2)	${}_{35}^{79}\text{Br}$	2
		(3)	$A_r = (A_r({}_{35}^{79}\text{Br}) \times \%({}_{35}^{79}\text{Br})) + (A_r({}_{35}^{81}\text{Br}) \times \%({}_{35}^{81}\text{Br}))$	
$A_r = (78.918 \times 0.5069) + (80.916 \times 0.4931)$ $A_r = 79.90$	1 1			
(e)	(1)	2,8,18,7	1	
	(2)	The emission lines are evidence that the electrons of bromine atoms emit specific wavelengths of visible light.	1	
		The energies of the emitted wavelengths are equal to the difference in energy between two electron energy levels in a bromine atom.	1	
(f)	(1)	The positions of the coloured lines in the line emission spectrum of sodium perfectly match those present in the line emission spectrum of seawater.	1	
	(2)	The line emission spectrum of seawater is compared with the line emission spectra of the known chemical elements.	1	
		Any coloured lines in the emission spectrum of seawater that do not match those belonging to an element that has been discovered are likely to belong to a previously undiscovered element such as iodine.	1	