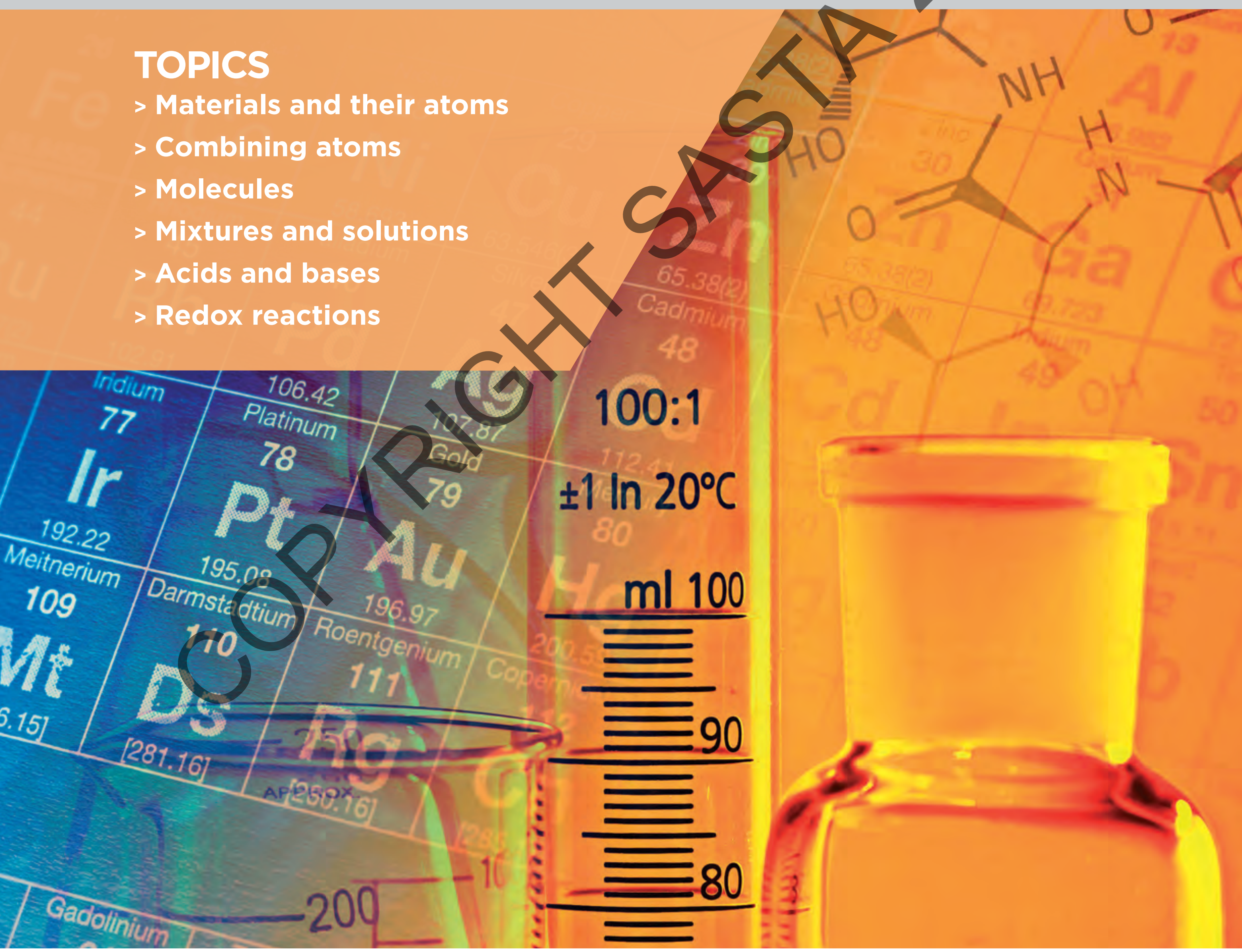
SACE STAGE 1
Australian Curriculum

WORKBOOK

Chemistry

TOPICS

> Materials and their atoms
> Combining atoms
> Molecules
> Mixtures and solutions
> Acids and bases
> Redox reactions



SASTA
South Australian Science Teachers Association
Supporting Teachers of Science
Advancing Science Education

SECOND EDITION
Rhys Lewis

Contents

CHAPTER 1

TOPIC 1: MATERIALS AND THEIR ATOMS

1.1: Properties and uses of materials

The uses of materials are related to their properties, including solubility, thermal and electrical conductivities, melting point, and boiling point.

Nanomaterials are substances that contain particles in the size range 1 – 100 nm.

- Suggest uses of materials, including nanomaterials, given their properties and vice versa.

Chemistry is the study of the composition, structure, properties and transformations of **matter** which is a term used to describe anything that has mass and takes up space. Matter is composed of **particles** including atoms, molecules, and ions that combine to form different types of **materials** including glass, plastic, metals, textiles and ceramics. The uses of a material are related to its properties including solubility, thermal and electrical conductivity, melting point, and boiling point.

Solubility

A **solution** is a homogeneous mixture formed by dissolving one material, called the **solute**, in a second material, called the **solvent**. For example, a solution of copper sulfate is formed by dissolving copper sulfate (solute) in water (solvent). **Solubility** is a measure of the extent to which a solute dissolves in a solvent at a given temperature and is expressed in units of grams per litre (g L^{-1}). The use of a material is dependent on its solubility in certain solvents.

Example 1.01

Liquid washing tablets (Figure 1.01) are commercial cleaning products used when washing dirty clothes. The tablets contain cleaning agents stored inside a water-soluble plastic capsule. The capsule dissolves in water and releases the stored cleaning agents when laundering clothes. The cleaning agents stored inside the plastic capsule have high solubility in water which allows them to dissolve and form a solution that removes dirt and stains when mixed with water inside a washing machine.

Figure 1.01: Liquid washing tablets.

Nanomaterials

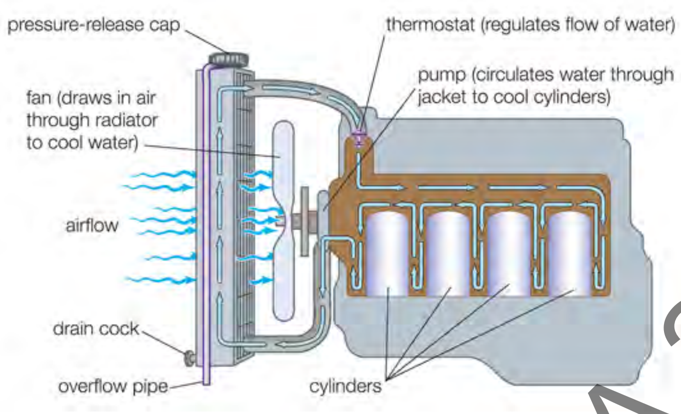
The physical and chemical properties of a material such as melting point, electrical and thermal conductivity, solubility, and chemical reactivity are dependent on the size of its particles. **Nanomaterials** are substances composed of particles that range in size from 1 – 100 nanometres. Materials constructed at this scale often exhibit distinctive physical and chemical properties that make them suitable for new and existing applications. The properties and uses of some common nanomaterials are identified below.

Nanomaterial	Properties	Use(s)
Silver nanoparticles	Antimicrobial (kills bacteria and fungi)	Medical bandages
Gold nanoparticles	Transform visible light into thermal radiation	Destroy cancer cells and shrink tumours.
Zinc oxide nanoparticles	Absorb and reflect ultraviolet radiation (UV).	Sunscreen
Titanium dioxide nanoparticles	Catalyses the breakdown of organic materials in sunlight.	Self-cleaning windows
Nanowhiskers	Repel water and oily materials	Waterproof and stain-resistant clothing and fabrics
Metal-organic frameworks	Very high surface area for the absorption of gases	Increasing the storage capacity of fuel tanks and gas cylinders
Zeolites	Porous structure with a very high surface area	Cracking of hydrocarbons, and water purification
Buckyballs	Forms a low-friction surface when bonded to other atoms	Lubricants
Carbon nanotubes	Light-weight and very high tensile strength	Construction, automotive and aircraft materials.
Nanocapsules	Non-toxic, water-soluble, hollow structures that transport materials in the body	Targeted drug delivery
Nano-robot	A cell-sized device capable of identifying and destroying target cells in the body.	Treatment of cancer and autoimmune diseases.

A defining feature of nanomaterials is their very large surface-to-volume ratio. Materials with high surface area to volume ratio have improved chemical reactivity which makes them suitable for use in batteries, fuel cells, and catalysts.

Question 1

The diagram below shows the cooling system for an internal combustion engine in a car which operates at a temperature between 100 and 200^{0}C.



(a) Car engines are made from metal alloys including cast iron and aluminium alloy.

Some physical properties of these two materials are identified below.

Material	Melting point (^{0}C)	Thermal conductivity
Cast iron	1200	High
Aluminium alloy	670	High

(1) Define the melting point of a material.

____

____ (1 mark) KA1

(2) State why materials with high melting points are used in the manufacture of engines.

____

____ (1 mark) KA2

(b) The cylinders in the engine are cooled by water as shown in the diagram.

Water has a boiling point of 100^{0}C.

(1) State why the cylinders are composed of materials with high thermal conductivity.

____

____ (1 mark) KA2

(2) State why it is important that water is circulated through the engine during operation.

____

____ (1 mark) KA2

Question 5

Silver nanoparticles are nanomaterials with applications including catalysis and medicine.

(a) Silver nanoparticles are used as a catalyst in the industrial conversion of benzene to phenol.

(1) State the property of silver nanoparticles that makes them suitable for use as a catalyst.

______________________________ (1 mark) KA1

(2) Phenol is used in the manufacture of polycarbonate plastics.

State and explain the benefit of using silver nanoparticles to manufacturers of phenol.

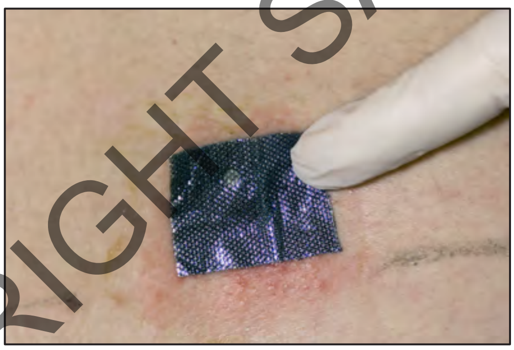
____

____

____

______________________________ (2 marks) KA2

(b) Silver nanoparticles have antimicrobial properties and are used in medical bandages as shown in the diagram below.



(1) State the benefit of using silver nanoparticles in medical bandages.

____

______________________________ (1 mark) KA2

(2) A home appliance company released a washing machine that uses silver nanoparticles to sterilise clothing and bedding.

State one advantage and one disadvantage to the environment of using silver nanoparticles in a washing machine.

____

____

____

______________________________ (2 marks) KA2

The components of a mixture are uncombined chemically, and each retains its own distinct physical and chemical properties which allows them to be separated using techniques including filtration, evaporation and distillation.

Filtration

Filtration is a technique that is used to separate a heterogeneous mixture containing solid particles in a liquid medium. The mixture is passed through a porous filter that permits the liquid to pass through while retaining the solid particles. The apparatus used in simple laboratory filtration is illustrated in Figure 1.11.

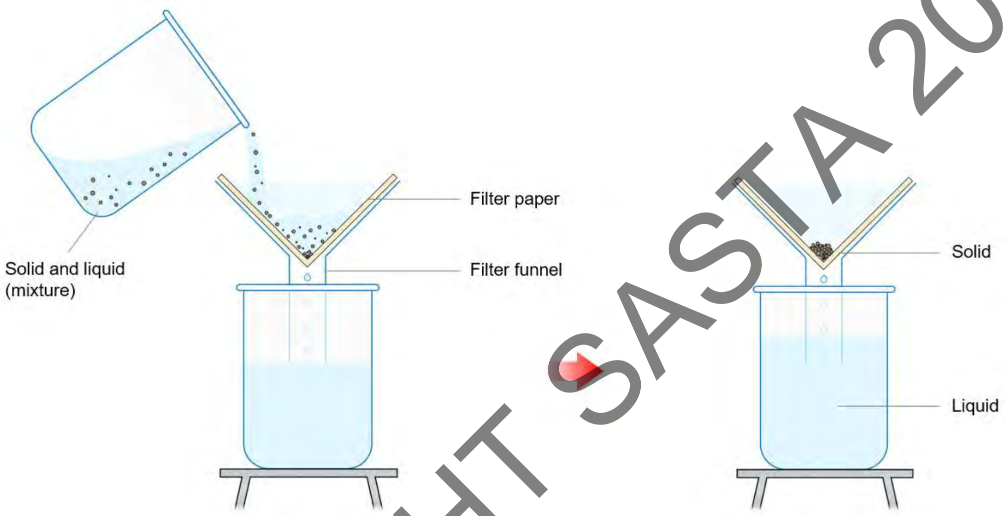


Figure 1.11: Apparatus used in simple laboratory filtration.

Evaporation

Evaporation is a technique used to separate a homogeneous mixture containing one or more solutes that are dissolved in a solvent. Heat is transferred to the mixture which causes the solvent to evaporate and leave behind the solid solute. The apparatus used in simple laboratory filtration is illustrated in Figure 1.12.

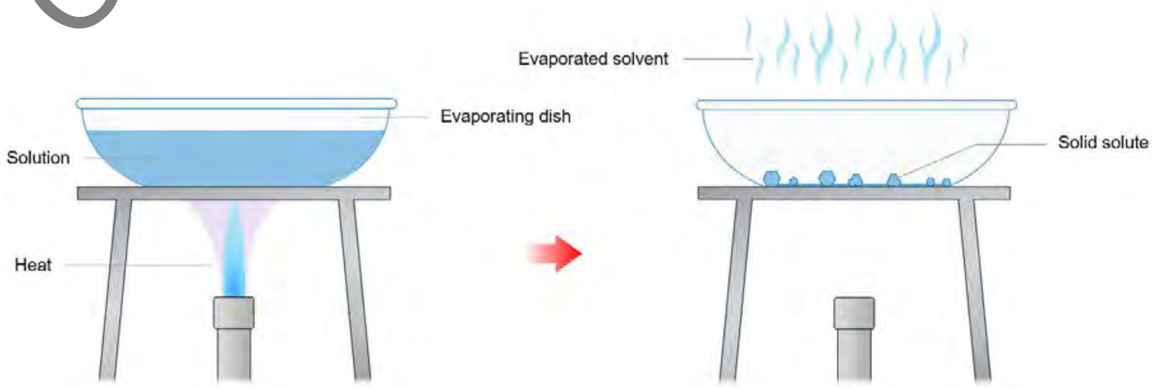


Figure 1.12: Apparatus used in simple laboratory evaporation.

Electron arrangement

The electrons of atoms orbit the nucleus within defined boundaries called **shells**. The electron shells in an atom each hold a maximum number of electrons which is calculated using the formula $2n^2$ where n is the shell number. The table below identifies the maximum number of electrons in the first four electron shells of an atom.

Shell name	Shell number	Maximum # electrons
K	1	2
L	2	8
M	3	18
N	4	32

Electrons fill lower-numbered shells first before filling progressively higher-numbered shells. Once an electron shell has been filled, electrons are placed in the next-highest shell. The arrangement of electrons is described using the **main-shell electron configuration** which denotes the number of electrons in each shell of an atom.

Example 1.09

Figure 1.16 shows the arrangement of electrons in helium, oxygen and magnesium atoms.

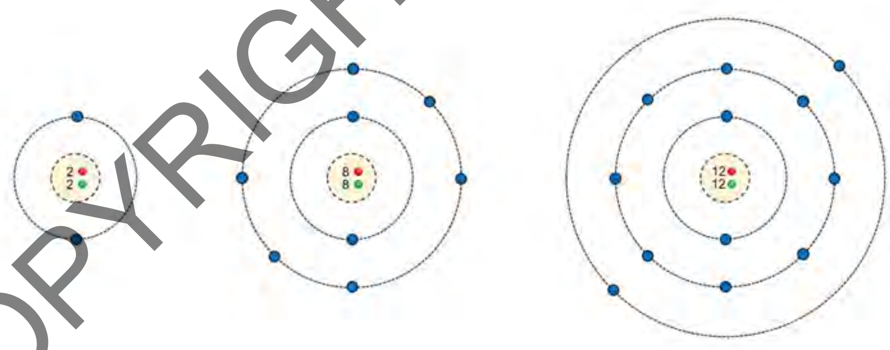


Figure 1.16: Electron arrangement in helium, oxygen and magnesium.

The main shell electron configurations of helium, oxygen and magnesium are identified in the table below.

Atom	Number of electrons	Main-shell electron configuration
Helium	2	2
Oxygen	8	2,6
Magnesium	12	2,8,2

Example 1.11

Robert Bunsen and Gustav Kirchhoff discovered two new elements between 1860 and 1861 using their newly developed spectroscope. The elements were identified by vaporising samples of spring water and comparing the emission lines to those present in the line emission spectra of the known chemical elements (Figure 1.29).

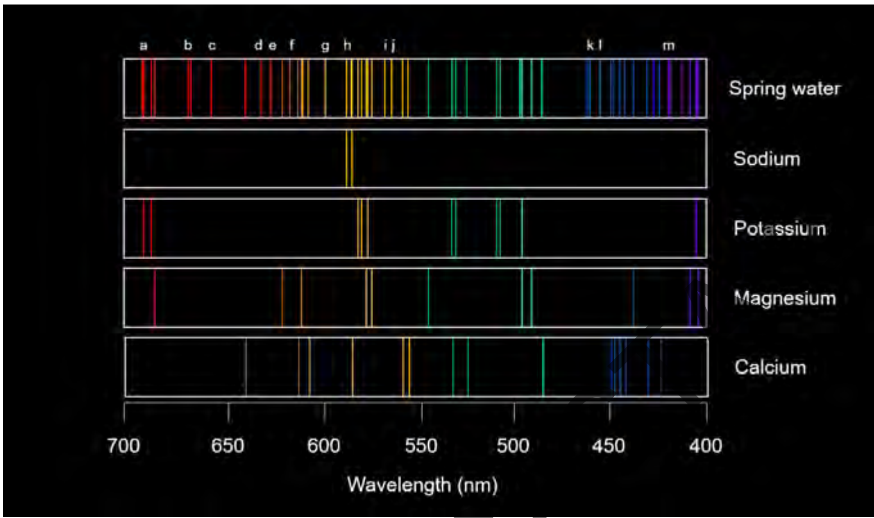


Figure 1.29: Line emission spectra for spring water and several chemical elements.

The lines marked (a)–(m) on the diagram are not featured in the emission spectra of sodium, potassium, magnesium or calcium and were identified as belonging to two previously undiscovered elements: rubidium (Rb) and caesium (Cs).

Example 1.12

Line absorption spectra are used to determine the elemental composition of the atmospheres of stars and planets. This is achieved by comparing the absorption spectrum of a star or planet with the line emission spectra of the chemical elements. Figure 1.30 shows how hydrogen is identified in the atmosphere of the star Sirius A.

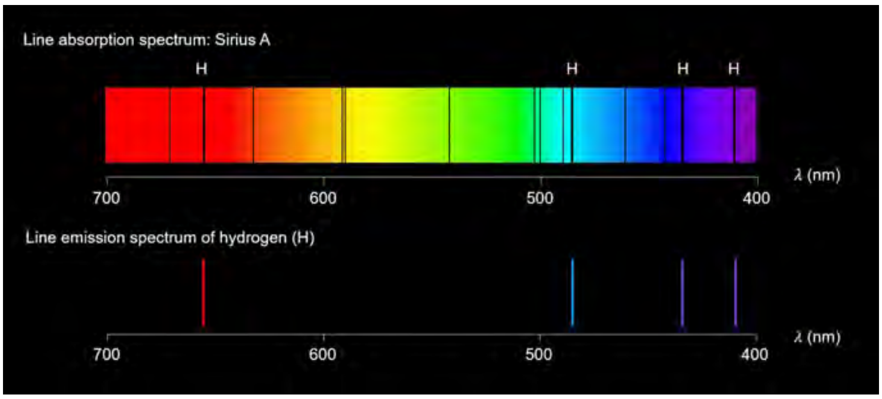


Figure 1.30: Identification of hydrogen in the atmosphere of Sirius A.

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

Many elements consist of a number of different isotopes, which have different physical properties but the same chemical properties.

- Represent isotopes of an element using appropriate notation.

A chemical **element** is a material that is composed of one type of atom. There are 92 naturally-occurring elements and 26 synthetic elements, and each is identified using a chemical symbol (X). The composition of the nucleus of an atom is described using the **atomic number** (Z) and **mass number** (A). The atomic number is the number of protons, and the mass number is the sum of protons and neutrons. The elements are represented using a combination of the chemical symbol, atomic number and mass number as shown in Figure 1.31.

$$^{A}_{Z}X \qquad ^{16}_{8}O \qquad ^{52}_{24}Cr$$

Figure 1.31 Representing atomic nuclei.

The number of neutrons (N) in an atom is calculated by subtracting the atomic number (Z) from the mass number (A).

Example 1.13

Plutonium-239 is a synthetic chemical element with an atomic number of 94 and a mass number of 239. The number of neutrons in the nucleus of an atom of plutonium-239 is calculated below.

N	=	$A - Z$
N	=	$239 - 94$
N	=	145

Example 1.14

Lead is a naturally-occurring chemical element with an atomic number of 82.

Determine the mass number of a lead atom with 128 neutrons.

N	=	$A - Z$
A	=	$N + Z$
A	=	$128 + 82$
A	=	210

Isotopes

The majority of the chemical elements have **isotopes**. Isotopes are atoms of the same chemical element that have different mass numbers. The difference in mass number is due to the variation in the number of neutrons in the nuclei of the atoms.

Example 1.15

Hydrogen has three naturally occurring isotopes. The atomic structure of the three hydrogen isotopes is depicted in Figure 1.32.

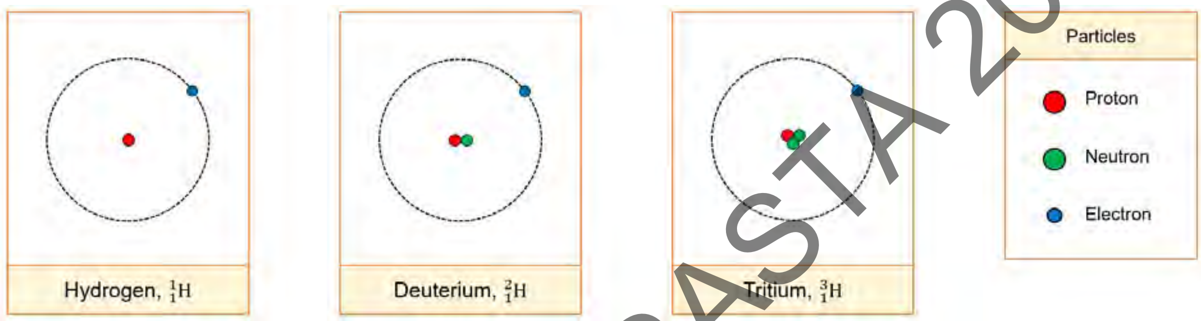


Figure 1.32: Isotopes of hydrogen.

The composition of each isotope of hydrogen is identified in the table below.

Isotope	Symbol	Protons	Neutrons	Electrons
Hydrogen	^1_1H	1	0	1
Deuterium	^2_1H	1	1	1
Tritium	^3_1H	1	2	1

The data shows that the isotopes have the same atomic number as each has the same number of protons, but the mass numbers are different as each has a different number of neutrons. The difference in mass causes each isotope to have slightly different **physical properties** including density, melting point and boiling point. The isotopes have the same **chemical properties** as each has the same number of electrons in the outermost electron shell and it is these electrons that are exchanged in a chemical reaction.

Isotopes are identified using the name of the element followed by the mass number. Some examples are identified in the table below.

Symbol	Name
$^{14}_{6}C$	carbon-14
$^{37}_{17}Cl$	chlorine-37
$^{235}_{92}U$	uranium-235

The arrangement of electrons in atoms and monatomic ions can be described in terms of shells and subshells.

- Write the electron configuration using subshell notation of an atom of any of the first 38 elements in the periodic table.

Electrons in atoms are located in shells and subshells around the nucleus with each shell and subshell having the capacity for storing a maximum number of electrons. The table below shows the maximum number of electrons stored in each of the first four shells and ten subshells in an atom.

Shell number (*n*)	Subshells	Maximum number of electrons
1	1s	2
2	2s 2p	2 6
3	3s 3p 3d	2 6 10
4	4s 4p 4d 4f	2 6 10 14

Electrons are positioned in shells and subshells in order of increasing energy. Lower energy electrons are located in lower-numbered shells and higher energy electrons are located in higher-numbered shells. Electrons begin to fill a higher energy subshell once a lower energy subshell has been filled. The energies of the first ten subshells in an atom are shown in Figure 1.33 below.

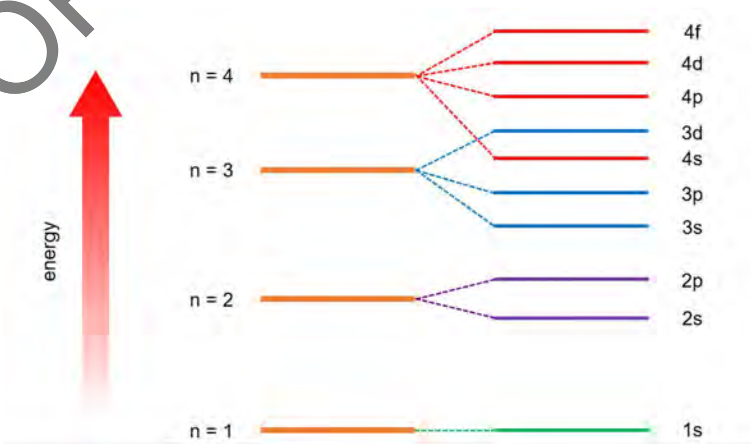


Figure 1.33: Energy of the first ten subshells in an atom.

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 from 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** such as atoms, molecules, ions, and empirical formula units*. The elementary particles that comprise different materials are identified in the table below.

Material	Elementary particle
Elements	Atoms
Molecular compound	Molecules
Continuous compound	Empirical formula unit*
Electrolyte solution	Ions

The number of elementary particles in the different materials can be conveniently compared using the **mole unit**.

Moles

The **mole** (symbol, n) is a unit of measurement used to express the quantity of elementary particles in a certain mass of material.

One mole of any material contains approximately 6.02 x 10^{23} elementary particles, which is equivalent to the number of atoms in 12 grams of carbon-12.

Some examples of a mole are identified in the table below.

Material	Example	Elementary particles in one mole
Elements	Helium	6.02 x 10^{23} helium atoms
Molecular compound	Carbon dioxide	6.02 x 10^{23} carbon dioxide molecules
Continuous compound	Silicon dioxide	6.02 x 10^{23} empirical formula units of silicon dioxide*
Electrolyte solution	Silver chloride, AgCl	6.02 x 10^{23} silver ions and 6.02 x 10^{23} chloride ions

*An empirical formula unit of a continuous compound is equivalent to the number of atoms present in one unit of its chemical formula. For example, one empirical formula unit of silicon dioxide (SiO_2) is composed of one silicon atom (Si) and two oxygen atoms (O).

Calculating moles

The number of moles in a certain quantity of material is calculated using the formula below.

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

Symbol	Variable	Unit
n	Number of moles	mol
m	Quantity mass	g
M	Molar mass	$g.mol^{-1}$

This formula and its rearrangements are used to calculate the number of moles, quantity mass and molar mass of a material.

Example 1.22

The molar mass of copper (Cu) is 63.55 $g.mol^{-1}$.

Calculate the number of moles in 15.45 g of copper.

n	=	$\frac{m}{M}$
n	=	$\frac{15.45}{63.55}$
n	=	0.24 mol

Example 1.23

The molar mass of glucose is 180.156 $g.mol^{-1}$.

Calculate the mass of 0.250 mol of glucose.

n	=	$\frac{m}{M}$
m	=	nM
m	=	0.250 x 180.156
m	=	45.039 g

The formula for calculating moles is also used when determining the number of moles of one element in a compound.

Question 15

(a) Use the periodic table of elements to calculate the molar masses of the following compounds.

(1) nitrous oxide (N_2O)	(2) sulfuric acid, H_2SO_4	(3) thioacetone, $(CH_3)_2CS$
(4) potassium nitrate, KNO_3	(5) manganese sulfate dihydrate $MnSO_4.2H_2O$	(6) phenylmercuric acetate, $CH_3COOHgC_6H_5$

(6 marks) **KA4**

(b) Calculate the number of moles (*n*) of material in the following:

(1) 0.6 g of glucose, $C_6H_{12}O_6$	(2) 1.0×10^6 g of carbon dioxide, CO_2	(3) 5.2 g of hydrated copper sulfate $CuSO_4.5H_2O$

(6 marks) **KA4**

(c) Calculate the quantity mass (*m*) of material in the following:

(1) 0.02 mol of iodine, I_2	(2) 0.92 mol of paraffin wax, $C_{31}H_{64}$	(3) 0.41 mol of iron(III) oxide, Fe_2O_3

(6 marks) **KA4**

1.4: The periodic table

In the modern periodic table, elements are arranged in order of increasing atomic number, and display periodic trends in their properties.

The position of an element in the periodic table is related to its metallic or non-metallic character.

- Identify trends in atomic radii, valencies, and electronegativities, across periods and down groups of the periodic table.
- Identify the position of an atom in the periodic table given its electron configuration.
- Identify the s, p, d, and f blocks of the periodic table.

The **periodic table of elements** is a chart displaying the chemical elements in order of increasing atomic number (Figure 1.38).

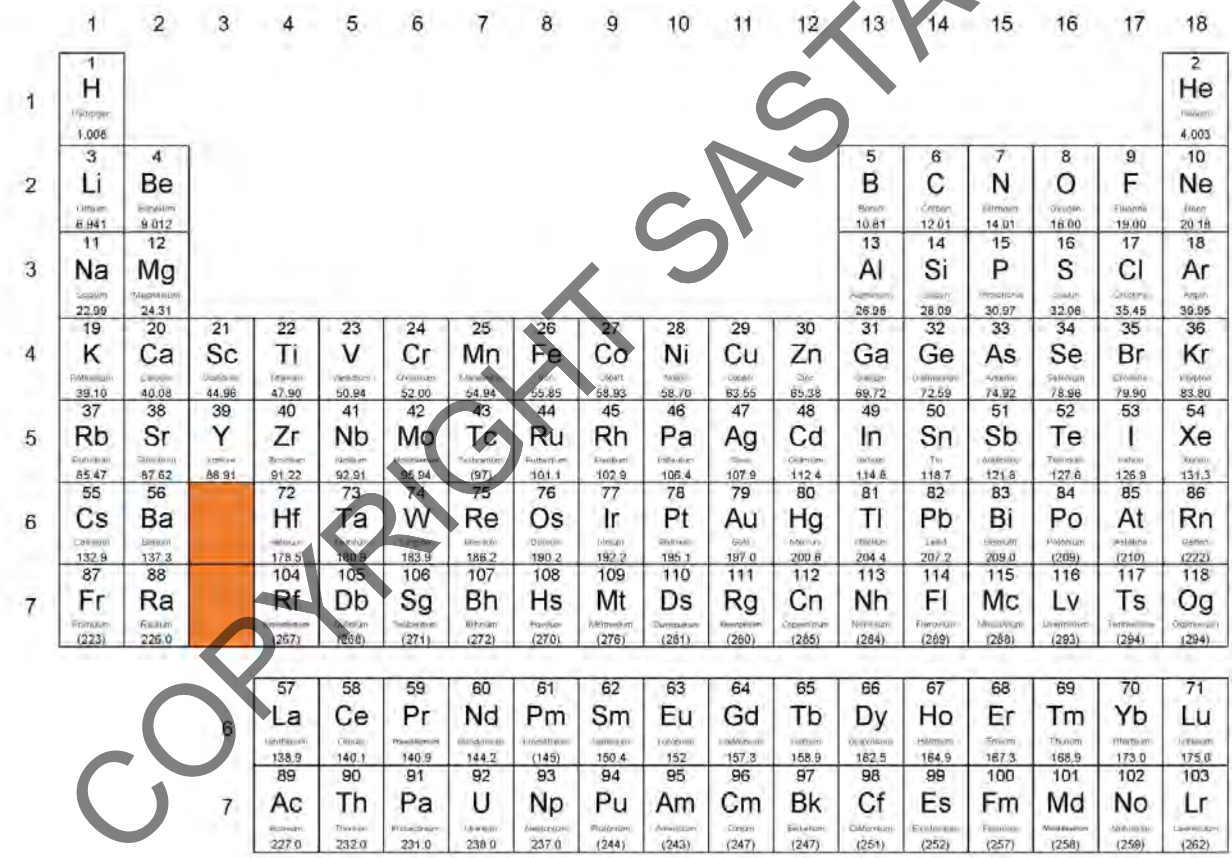


Figure 1.38: The periodic table of elements.

The 118 chemical elements are arranged into horizontal rows called **periods** in order of increasing atomic number, and in vertical columns called **groups** according to similarities in chemical properties. There are currently seven periods and 32 groups on the periodic table. However, a 15-column segment has been removed from periods 6 and 7 and placed beneath period 7 such that the main body of the periodic table contains 7 periods and 18 groups as shown in Figure 1.38.

Atomic radius

The **atomic radius** of an element is a measure of 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. The atomic radii of helium and magnesium atoms are depicted in Figure 1.40.

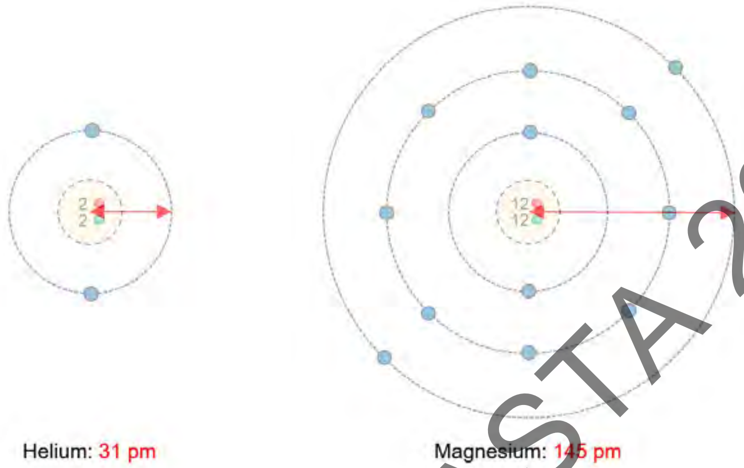


Figure 1.40: Atomic radius of helium and magnesium atoms measured in picometres (x 10^{-12} m)

Trends in the atomic radii of the elements can be identified from the position of an element on the periodic table as shown in Figure 1.41.

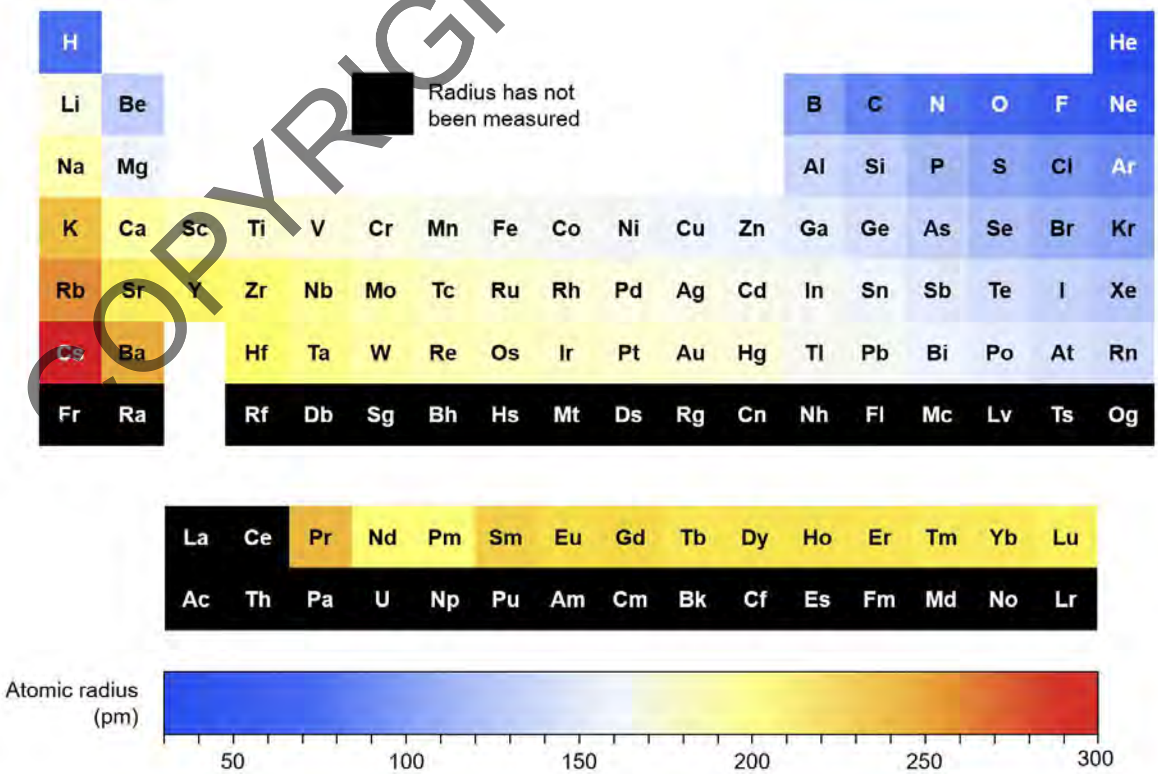


Figure 1.41: Trends in the atomic radii of the elements.

Figure 1.41 shows that the atomic radii of the elements increase down a group and decrease from left to right across a period.

Example 1.30

Figure 1.42 shows the change in atomic radii down group 2 on the periodic table.

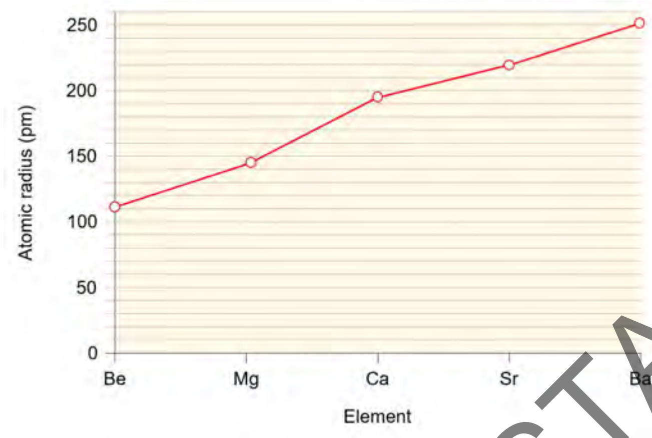


Figure 1.42: Change in atomic radii down group 2.

The increase in atomic radius down a group is due to the increase in the number of electron shells in the atoms of the elements. For example, barium (Ba) has six electron shells and has a larger atomic radius than strontium (Sr) which has five electron shells.

Example 1.31

Figure 1.43 shows the change in atomic radii from left to right across period 3.

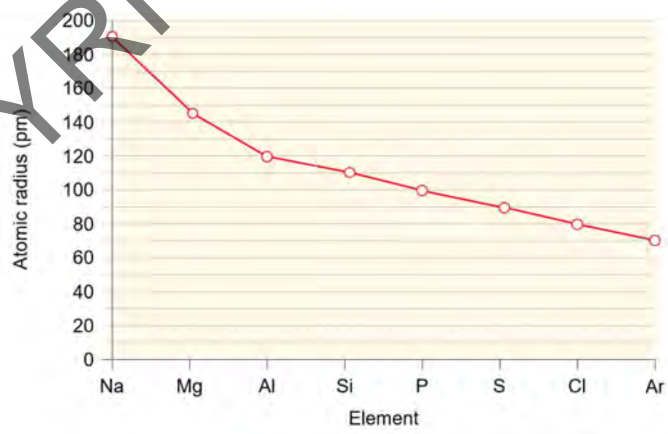


Figure 1.43: Change in atomic radii across period 3

The decrease in atomic radius is due to the increase in **nuclear charge*** on the atoms of the elements from left to right across a period. For example, the atomic radius of argon (Ar) is shorter than the atomic radius of sodium (Na) as the nucleus of argon has more protons that attract the electron shells toward the nucleus and effectively lower the atomic radius.

*Nuclear charge is the charge on the nucleus and is equal to the number of protons in an atom.

Metallic and non-metallic character

The elements on the periodic table are classified as metals, non-metals and metalloids depending on how readily electrons are added or removed from the valence shell of an atom in a chemical reaction.

Element	Definition
Metal	Have valence electrons removed in a chemical reaction.
Non-metal	Have valence electrons added in a chemical reaction.
Metalloid	Have valence electrons added or removed in a chemical reaction.

Figure 1.48 shows the positions of the metals, non-metals and metalloids on the periodic table of elements. It should be noted that the classification of some elements is still disputed.

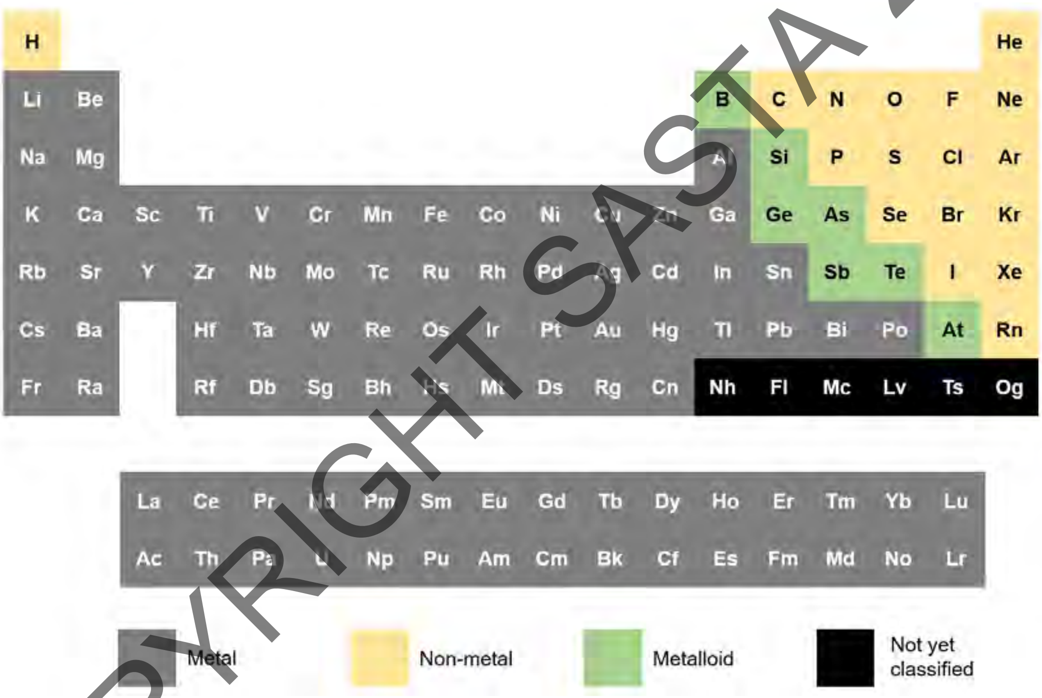


Figure 1.48: Positions of metals, non-metals and metalloids on the periodic table.

The terms **metallic character** and **non-metallic character** are used to define how readily valence electrons are removed or added to an atom in a chemical reaction. Elements with high metallic character lose valence electrons readily and form positively-charged ions (cations), and elements with high non-metallic character gain valence electrons and form negatively-charged ions (anions).

Trends in the metallic and non-metallic character of the elements can be identified on the periodic table. The trends are summarised in the table below.

Character	Periodic Trend	Main group trend
Metallic	Decreases from left to right	Increases down a group
Non-metallic	Increases from left to right	Decreases down a group

Review Test 1

Question 1

Seawater is a homogeneous mixture composed of several elements that are dissolved in water.

The elemental composition of a one-litre volume of seawater is given below.

Element	Mass (x 10^{-3} g)
sodium	11
calcium	400
potassium	380
chlorine	?
bromine	65

(a) Describe a homogeneous mixture using seawater as an example.

(2 marks) **KA2**

(b) Describe and explain how the elements may be separated from water using simple distillation.

(2 marks) **KA2**

(c) Sodium and chlorine are abundant in seawater in the form of sodium cations (Na^+) and chloride anions (Cl^-).

(1) Calculate the number of moles of sodium ions in one-litre of seawater.

(2 marks) **KA4**

(2) There are 3.22 x 10^{20} chloride ions in one-litre of seawater.

Determine the mass of chloride ions in one-litre of seawater.

(4 marks) **KA4**

(d) Seawater contains two isotopes of bromine (Br).

The relative atomic mass and abundance of each isotope are provided below.

Isotope	Relative atomic mass	Abundance (%)
bromine-79	78.918	50.69
bromine-81	80.916	49.31

(1) Define the term isotope using the isotopes of bromine as an example.

(2 marks) KA1

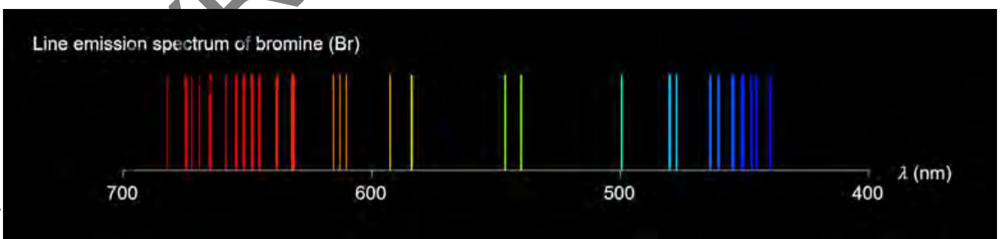
(2) Write the chemical symbol for the most abundant isotope of bromine, identifying the atomic and mass numbers.

(2 marks) KA1

(3) Calculate the relative atomic mass of bromine using the information above.

(2 marks) KA4

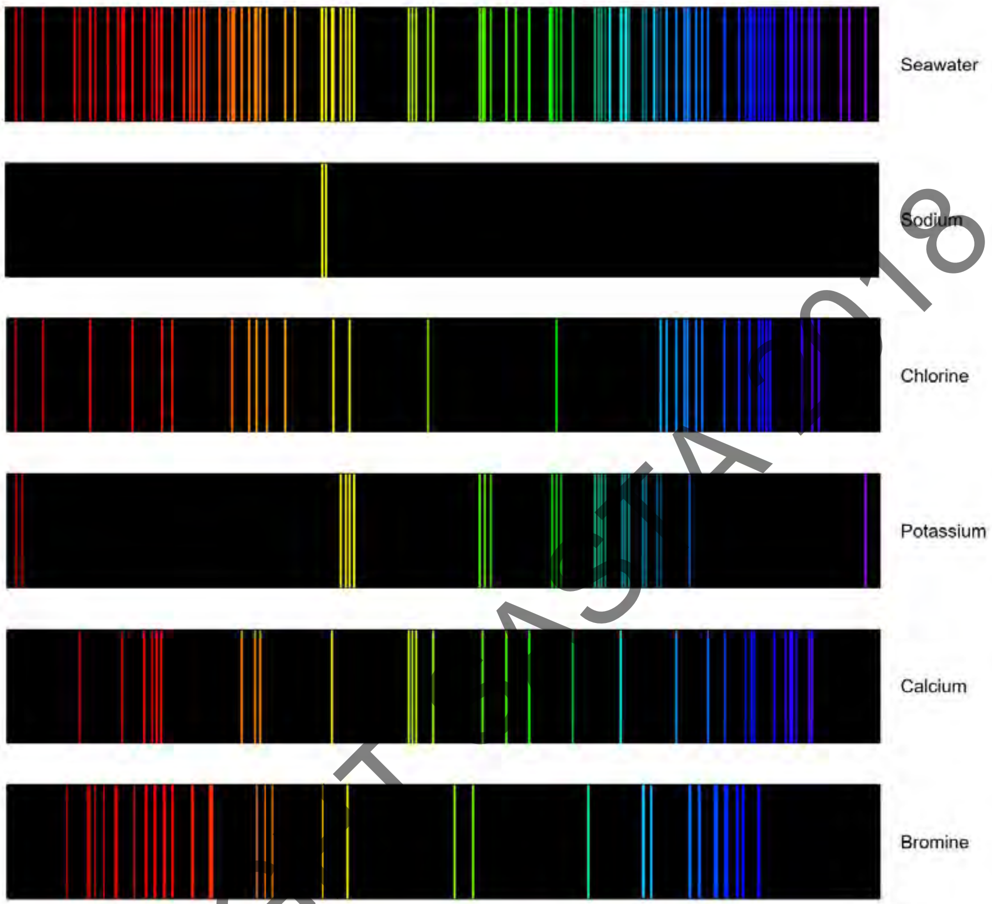
(4) The diagram below shows the line emission spectrum of bromine.



Explain how the coloured lines in the spectrum above provide evidence that electrons in atoms are arranged in distinct energy levels.

(2 marks) KA1

(e) The diagram below shows a line emission spectrum for a sample of seawater and several elements present in seawater.



(1) State the evidence from the diagram that seawater contains sodium.

____

____ (1 mark) KA2

(2) In 1811, French chemist Bernard Courtois discovered a new element called iodine in seawater.

The presence of iodine was confirmed over 50 years later using spectroscopy.

Describe and explain how line emission spectra make possible the identification of previously undiscovered chemical elements such as iodine.

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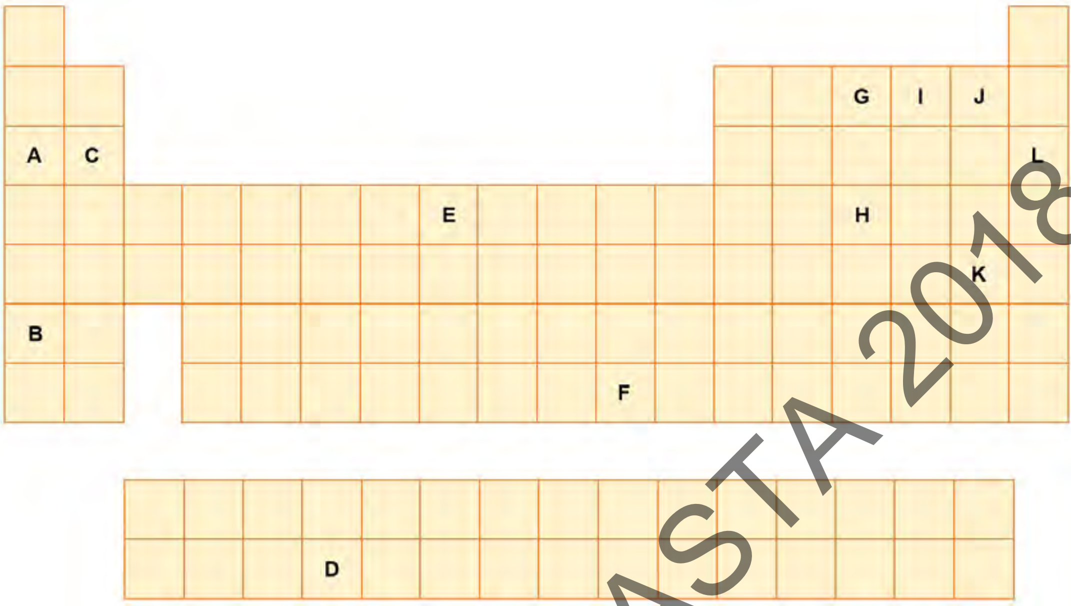
____ (2 marks) KA2

Question 2

The diagram below shows a periodic table with several elements represented using different letters of the alphabet.



(a) State the letter which represents an element in f-block.

____ (1 mark) **KA1**

(b) State the number of letters in the diagram that represent elements in p-block.

____ (1 mark) **KA1**

(c) State the letter which represents the element with the lowest molar mass.

____ (1 mark) **KA1**

(d) State the letter representing an element with a valency of 2.

____ (1 mark) **KA1**

(e) State the letter representing the element with the greatest metallic character.

____ (1 mark) **KA1**

(f) Element E has 26 electrons.

(1) Write the electron configuration of element E.

____ (2 marks) **KA1**

(2) State the block where element E is located on the periodic table.

____ (1 mark) **KA1**

(g) Elements G, I and J are non-metals.

State and explain which element has the greatest non-metallic character given their positions on the periodic table.

(2 marks) KA2

(h) Elements G and H are in main group 5 of the periodic table.

(1) State the property of elements located in main group 5 of the periodic table.

(1 mark) KA1

(2) State and explain which element has the highest electronegativity.

(3 marks) KA1

(i) Elements A, C and L are in period 3 on the periodic table.

(1) State the property of elements located in period 3 of the periodic table.

(1 mark) KA1

(2) State and explain which element has the shortest atomic radius given their positions on the periodic table.

(3 marks) KA1

CHAPTER 6

6.1: Concepts of oxidation and reduction

A range of reactions, including reactions of metals, combustion, and electrochemical processes, can be considered as 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.
- Determine the oxidation states of atoms in elements and monatomic ions, and in compounds and polyatomic ions.

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

Combination with oxygen

The concept of oxidation and reduction was introduced by French scientist Antoine Lavoisier (Figure 6.01) in the late 18th century.

Figure 6.01: Antoine Lavoisier (1743–1794)

Lavoisier introduced the term oxidation to describe the gain of oxygen by a substance in a chemical reaction. For example, hydrogen is oxidised in the reaction below as it gains oxygen.

	$2H_2$	+	O_2	→	$2H_2O$

Lavoisier used the term reduction to describe the reduction in mass of a substance when oxygen is removed. For example, potassium chlorate is reduced in the reaction below as it loses oxygen.

	$2KClO_3$	→	$2KCl$	+	$3O_2$

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

Transfer of electrons

Oxidation and reduction are also defined in terms of the transfer of electrons in a chemical reaction.

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 describe the processes of oxidation and reduction separately. Each half-reaction is described using an equation that identifies the number of electrons that are lost or gained by a substance in a reaction. Equations that describe either oxidation or reduction are called **half-equations** as they describe half a redox reaction.

Example 6.01

Figure 6.02 shows the reaction of iron and chlorine.

Figure 6.02: Reaction of iron and chlorine.

The reaction is described in the equation below.

$$2Fe + 3Cl_2 \rightarrow 2FeCl_3$$

Iron is oxidised in the reaction and chlorine is reduced. The oxidation of iron is described in the half-equation below.

$$Fe \rightarrow Fe^{3+} + 3e^-$$

The half-equation shows that each iron atom loses three electrons in the reaction. The three electrons appear on the right-hand side of the equation to show that they are lost in the reaction.

The reduction of chlorine is described in the half-equation below.

$$Cl_2 + 2e^- \rightarrow 2Cl^-$$

The half-equation shows that each chlorine molecule gains two electrons in the reaction. The electrons appear on the left-hand side of the equation to show that they are gained in the reaction.

Question 97

Chlorine and bromine are elements with many different oxidation states.

(a) Determine the oxidation number of chlorine (Cl) in the following substances.

ClO_3^-: ______	$HClO_4$: ______	KCl: ______
HClO: ______	ClO_2^-: ______	ClO_2: ______

(6 marks) **KA1**

(b) Determine the oxidation number of bromine (Br) in the following substances.

BrO^-: ______	BrO_2: ______	$NaBrO_4$: ______
$HBrO_2$: ______	$KBrO_3$: ______	Br^-: ______

(6 marks) **KA1**

(c) The diagram below shows the reaction of chlorine with sodium bromide solution.

The reaction is described in the equation below.

$$Cl_2 + 2NaBr \rightarrow 2NaCl + Br_2$$

(1) State the change in oxidation number of chlorine in the reaction.

______________________________ (2 marks) **KA2**

(2) State whether chlorine is oxidised or reduced in the reaction.

______________________________ (1 mark) **KA2**

(3) State the change in oxidation number of bromine in the reaction.

______________________________ (2 marks) **KA2**

(4) State whether bromine is oxidised or reduced in the reaction.

______________________________ (1 mark) **KA2**

- Write oxidation and reduction half-equations, in neutral and acidic conditions, given reactant and product species.
- Combine half-equations to write a chemical equation.

Half-equations are written to describe the processes of oxidation and reduction in a redox reaction. A method of writing half-equations in neutral conditions* is described in the example below.

Example 6.02

The equation below shows the redox reaction of aluminium and iodine.

$$2Al + 3I_2 \rightarrow 2AlI_3$$

Aluminium is oxidised in the reaction as its oxidation number increases from 0 to +3. A half-equation for the oxidation of aluminium is written by stating the formula for aluminium (Al) on the left-hand side of the arrow and the oxidation-product of aluminium (Al^{3+}) on the right-hand side of the arrow.

$$Al \rightarrow Al^{3+}$$

The number of aluminium atoms is the same on both sides of the arrow, so the final step is to balance electric charge by adding electrons (symbol, e-) to the side of the arrow which has the greater positive charge.

Al	→	Al^{3+}
0		3+

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

$$Al \rightarrow Al^{3+} + 3e^-$$

Iodine is reduced in the reaction as its oxidation number decreases from 0 to −1 in the reaction. A half-equation for the reduction of iodine is written by stating the formula for iodine (I_2) on the left-hand side of the arrow and the reduction-product of iodine (I^-) on the right-hand side of the arrow.

$$I_2 \rightarrow 2I^-$$

Electrons are added to the side of the arrow with the greater positive charge.

I_2	→	$2I^-$
0		2−

In this example, two electrons are added to the left-hand side of the arrow.

$$I_2 + 2e^- \rightarrow 2I^-$$

*neutral conditions refers to a redox reaction that occurs at pH 7 and does not require acid or base.

Combining half-equations

Half-equations are combined when writing the overall reaction equation. A method of combining half-equations is described in the below.

The half-equations below describe the oxidation of aluminium and the reduction of iodine in the reaction described in Example 6.02.

$$Al \rightarrow Al^{3+} + 3e^-$$

$$I_2 + 2e^- \rightarrow 2I^-$$

The first step is to ensure that each half-equation has the same number of electrons. This is done by finding the lowest common multiple of the electron-numbers in the half-equations and by multiplying all species by a factor that results in each half-equation having the same number of electrons. In this example, the lowest common multiple of the electron-numbers in the half-equations (3 and 2) is 6, so all species in the oxidation half-equation are multiplied by 2 and all species in the reduction half-equation are multiplied by 3 such that each equation has 6 electrons.

$$2Al \rightarrow 2Al^{3+} + 6e^-$$

$$3I_2 + 6e^- \rightarrow 6I^-$$

The second, and final step is to combine the half-equations by stating the formulae of all reactants on the left of the arrow and all products to the right of the arrow. The electrons cancel out and are not included in the overall equation as the same number are present on both sides of the arrow.

$$2Al + 3I_2 + \cancel{6e^-} \rightarrow 2Al^{3+} + 6I^- + \cancel{6e^-}$$

The overall equation is written below.

$$2Al + 3I_2 \rightarrow 2Al^{3+} + 6I^-$$

Oxidisers and reducers

The terms **oxidiser** and **reducer** are used in the context of a redox reaction to identify the species that gained electrons and the species that lost electrons.

An oxidiser gains electrons, and a reducer loses electrons in a redox reaction.

Oxidisers are also known as **oxidants** and **oxidising agents**, and reducers are also known as **reductants** and **reducing agents**.

Question 100

The equation below shows the redox reaction of calcium and fluorine.

$$Ca + F_2 \rightarrow CaF_2$$

(a) State the evidence that fluorine is reduced in the reaction.

(1 mark) KA2

(b) Write a half-equation to show the reduction of fluorine in the reaction.

(2 marks) KA4

(c) State the number of electrons gained by each fluorine molecule in the reaction.

(1 mark) KA2

(d) Write a half-equation to show the oxidation of calcium in the reaction.

(2 marks) KA4

(e) Write the formula of the oxidiser in the reaction.

(1 mark) KA2

Question 101

The equation below shows the redox reaction of antimony (Sb) and chlorine.

$$2Sb + 5Cl_2 \rightarrow 2SbCl_5$$

(a) State the evidence that antimony is oxidised in the reaction.

(1 mark) KA2

(b) Write a half-equation to show the oxidation of antimony in the reaction.

(2 marks) KA4

(c) State the number of electrons gained by each antimony atom in the reaction.

(1 mark) KA2

(d) Write a half-equation to show the reduction of chlorine in the reaction.

(2 marks) KA4

(e) Write the formula of the reducer in the reaction.

(1 mark) KA2

Question 102

The equation below shows the redox reaction of zinc and iodine.

$$Zn + I_2 \rightarrow ZnI_2$$

(a) Write half-equations to show the processes of oxidation and reduction.

(4 marks) KA4

(b) Combine the half-equations and write the overall equation for the reaction.

(2 marks) KA4

(c) Name the formula of the reactant that is reduced in the reaction.

(1 mark) KA2

(d) Write the formula of the reducer in the reaction.

(1 mark) KA2

Question 103

The equation below shows the redox reaction of aluminium and bromine.

$$2Al + 3Br_2 \rightarrow 2AlBr_3$$

(a) Write half-equations to show the processes of oxidation and reduction.

(4 marks) KA4

(b) Combine the half-equations and write the overall equation.

Show all working.

(3 marks) KA4

(c) Name the formula of the reactant that is oxidised in the reaction.

(1 mark) KA2

(d) Write the formula of the oxidiser in the reaction.

(1 mark) KA2

Writing half-equations in acidic conditions

Certain redox reactions will only occur in acidic conditions. The function of the acid (H^+) is to remove oxygen atoms from the reactant that is reduced in the reaction. The half-equation below shows the reduction of permanganate ions (MnO_4^-) to manganese ions (Mn^{2+}) in acidic conditions.

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

A method of writing half-equations in acidic conditions is described in the example below.

Example 6.03

Ethanol is oxidised by an acidic solution containing dichromate ions. Dichromate ions are reduced to chromium(III) ions according to the unbalanced half-equation below.

$$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$$

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 which has fewer oxygen atoms. In this example, seven water molecules are added to the right-hand side of the arrow to balance oxygen.

$$Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$$

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

$$Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$$

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

$Cr_2O_7^{2-}$	+	$14H^+$	→	$2Cr^{3+}$	+	$7H_2O$
2−		14+		6+		0
Total charge = 12+				Total charge = 6+		

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

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$$

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

$$C_2H_5OH \rightarrow CH_3COOH$$

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 which has fewer oxygen atoms. In this example, one water molecule is added to the left-hand side of the arrow to balance oxygen.

$$C_2H_5OH + H_2O \rightarrow CH_3COOH$$

The second step is to balance hydrogen by adding protons (H^+) to the side of the arrow which has fewer hydrogen atoms. In this example, four protons are added to the right-hand side of the arrow.

$$C_2H_5OH + H_2O \rightarrow CH_3COOH + 4H^+$$

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

C_2H_5OH	+	H_2O	$\rightarrow$	CH_3COOH	+	$4H^+$
0		0		0		4+
Total charge = 0				Total charge = 4+		

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

$$C_2H_5OH + H_2O \rightarrow CH_3COOH + 4H^+ + 4e^-$$

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.

$$2Cr_2O_7^{2-} + 28H^+ + 12e^- \rightarrow 4Cr^{3+} + 14H_2O$$

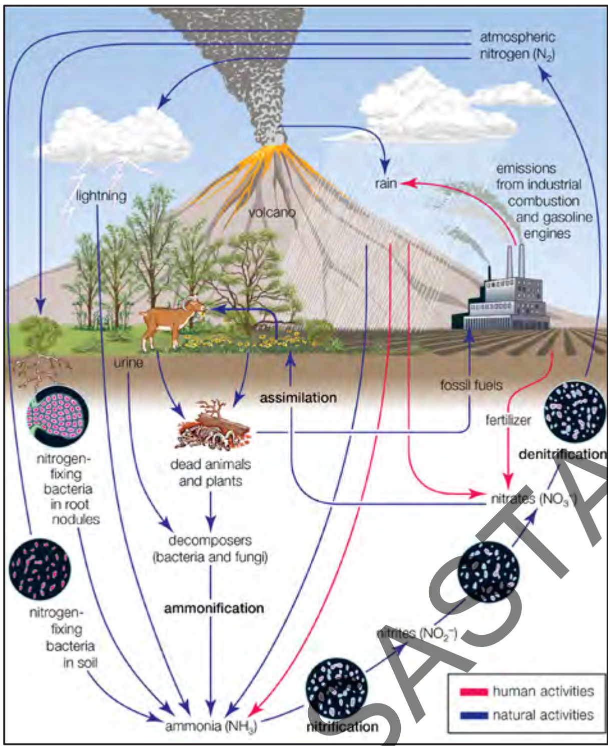
$$3C_2H_5OH + 3H_2O \rightarrow 3CH_3COOH + 12H^+ + 12e^-$$

Protons, water molecules and electrons appear on both sides of the arrow and these are cancelled out when combining the half-equations. This is done by removing all protons or water molecules from the side of the arrow which has fewer and by removing an equal number of that species from the opposing side of the arrow. For example, 3 water molecules are removed from both sides to cancel water and 12 protons are removed from both sides to balance oxygen in this reaction.

$$3C_2H_5OH + 2Cr_2O_7^{2-} + 16H^+ \rightarrow 3CH_3COOH + 4Cr^{3+} + 11H_2O$$

Question 104

The diagram below shows the nitrogen cycle.



Several process in the nitrogen cycle involve the redox reactions in acidic conditions.

Balance the following half-equations in acidic conditions.

(a) Nitrogen-fixation

$$N_2 \rightarrow 2NH_3 + H_2$$

(2 marks) **KA2**

(b) Nitrification part one.

$$NH_3 + O_2 \rightarrow NH_2OH$$

(2 marks) **KA2**

(c) Nitrification part two.

$$NH_2OH \rightarrow NO_2^-$$

(2 marks) **KA2**

(d) Nitrification part three.

$$NO_2^- \rightarrow NO_3^-$$

(2 marks) **KA2**

(e) Denitrification

$$NO_3^- \rightarrow N_2$$

(2 marks) **KA2**

6.2: Metal reactivity

Metals differ in their tendency to lose electrons; more reactive metals lose electrons more easily.

A more reactive metal is able to donate electrons to the ion of a less active metal in a displacement reaction.

Differences in metal reactivity can be represented as a metal activity series.

- Write equations and half-equations for reactions between a metal and the ion of a less active metal.
- Determine whether a reaction will occur between a metal and a solution containing the ions of another metal, given a metal activity series containing both metals.

Metals are elements composed of atoms that lose one or more electrons in a chemical reaction. The reactivity of metals can be defined in terms of the tendency to lose electrons. The most reactive metals lose electrons rapidly in a chemical reaction and the least reactive metals rarely lose an electron under standard conditions. The reactivity of metals is represented as a **metal activity series** which ranks the metals from most reactive to least reactive. The metal activity series below shows the activity of 13 metals in aqueous solution.

	Metal	Oxidation half-equation
Most reactive	Potassium	$K_{(s)} \rightarrow K^{+}_{(aq)} + e^{-}$
	Calcium	$Ca_{(s)} \rightarrow Ca^{2+}_{(aq)} + 2e^{-}$
	Sodium	$Na_{(s)} \rightarrow Na^{+}_{(aq)} + e^{-}$
	Magnesium	$Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + 2e^{-}$
	Aluminium	$Al_{(s)} \rightarrow Al^{3+}_{(aq)} + 3e^{-}$
	Zinc	$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^{-}$
	Iron	$Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^{-}$
	Nickel	$Ni_{(s)} \rightarrow Ni^{2+}_{(aq)} + 2e^{-}$
	Tin	$Sn_{(s)} \rightarrow Sn^{2+}_{(aq)} + 2e^{-}$
	Lead	$Pb_{(s)} \rightarrow Pb^{2+}_{(aq)} + 2e^{-}$
	Copper	$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^{-}$
	Silver	$Ag_{(s)} \rightarrow Ag^{+}_{(aq)} + e^{-}$
Least reactive	Gold	$Au_{(s)} \rightarrow Au^{3+}_{(aq)} + 3e^{-}$

The metal activity series is used when predicting the products of the reactions of metals with acids, water, and with each other.

Metal displacement reaction

A **metal displacement reaction** is the reaction of a metal with a compound or solution containing cations of a less active metal. In all cases, electrons are transferred from the more active to the less active metal which displaces the less active metal from the compound or solution.

Example 6.04

Figure 6.03 shows the reaction of zinc (Zn) and solution containing copper ions (Cu^{2+}).

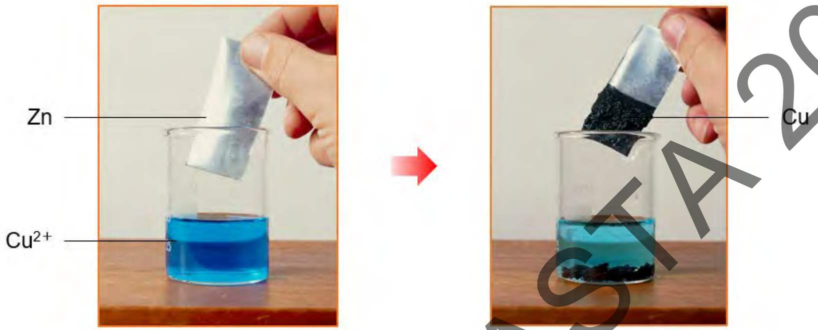


Figure 6.03: Reaction of zinc and a solution of copper ions.

Zinc is more reactive than copper and electrons are transferred from zinc atoms to copper ions which displaces copper from solution according to the equation below.

$$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$$

Metal displacement reactions are redox reactions and as electrons are transferred from the more active metal to the less active metal. The oxidation of zinc and the reduction of copper ions are described in the half-equations below.

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^{-}$$

$$Cu^{2+}_{(aq)} + 2e^{-} \rightarrow Cu_{(s)}$$

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.

$$Mg + 2Ag^{+} \rightarrow Mg^{2+} + 2Ag$$

$$Ag + Mg^{2+} \rightarrow \text{No reaction}$$

Question 108

The products of a metal displacement reaction can be predicted from the positions of two metals on the metal activity series.

Most reactive

Metal
Potassium
Calcium
Sodium
Magnesium
Aluminium
Zinc
Iron
Nickel
Tin
Lead
Copper
Silver
Gold

Least reactive

(a) Use the metal activity series to determine whether the following reactions will occur.

(1) Lead nitrate solution and iron.

____ (1 mark) **KA1**

(2) Nickel sulfate solution and copper.

____ (1 mark) **KA1**

(3) Copper sulfate solution and magnesium.

____ (1 mark) **KA1**

(b) Write a fully balanced ionic equation for the metal displacement reactions below.

(1) Zinc and tin(II) chloride solution.

____ (2 marks) **KA1**

(2) Iron and copper sulfate solution.

____ (2 marks) **KA1**

(3) Aluminium and nickel nitrate solution.

____ (2 marks) **KA1**

Reaction of metals and dilute acids

Active metals react with dilute acids forming a salt and hydrogen according to the unbalanced general equation below.

M	$+$	HX	$\rightarrow$	MX	$+$	H_2
metal		acid		salt		hydrogen

Calcium is an active metal that reacts with hydrochloric acid according to the fully balanced and ionic equations below.

$$Ca + 2HCl \rightarrow CaCl_2 + H_2$$

$$Ca + 2H^+ \rightarrow Ca^{2+} + H_2$$

The reaction of a metal and acid is a redox reaction. The oxidation of the metal and the reduction of protons (H^+) are described in the half-equations below.

$$M \rightarrow M^+ + e^-$$

$$2H^+ + 2e^- \rightarrow H_2$$

The activity of a metal can be determined from its reaction with a dilute acid. Figure 6.04 shows the reactions of four different metals with dilute hydrochloric acid solution.

Figure 6.04: Reactions of metals with hydrochloric acid (left to right: magnesium, zinc, iron, lead).

The following conclusions are drawn from the diagram:

1. Magnesium is the most reactive metal in the diagram as it reacts most vigorously with hydrochloric acid solution.
2. Iron is less reactive than both magnesium and zinc as it reacts less vigorously with hydrochloric acid solution.
3. Lead is the least reactive metal in the diagram as it does not react with dilute hydrochloric acid solution.

Question 112

The diagram below shows the reactions of four metals with dilute hydrochloric acid solution.

The metals are calcium (left), magnesium, zinc and copper (right).

(a) Identify the most reactive metal in the diagram and give a reason for your choice.

____

____________________________________ (2 marks) KA2

(b) State the evidence from the diagram that zinc is more reactive than copper.

____

____________________________________ (1 mark) KA2

(c) Write a fully balanced chemical equation for the reaction of calcium and hydrochloric acid solution.

____________________________________ (2 marks) KA2

(d) Write a fully balanced ionic equation for the reaction of magnesium and hydrochloric acid solution.

____________________________________ (2 marks) KA2

(e) Write half-equations showing the processes of oxidation and reduction in the reaction of zinc and hydrochloric acid solution.

____

____________________________________ (4 marks) KA1

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 that either produce or are initiated by a flow of electrons. Electrochemical reactions occur in **electrolytic cells*** and **galvanic cells**.

Galvanic cell

A galvanic cell is an electrochemical cell that generates an electric current from spontaneous redox reactions. The cell is named after 18th century Italian scientist Luigi Galvani. The components of a simple galvanic cell are identified in Figure 6.05.

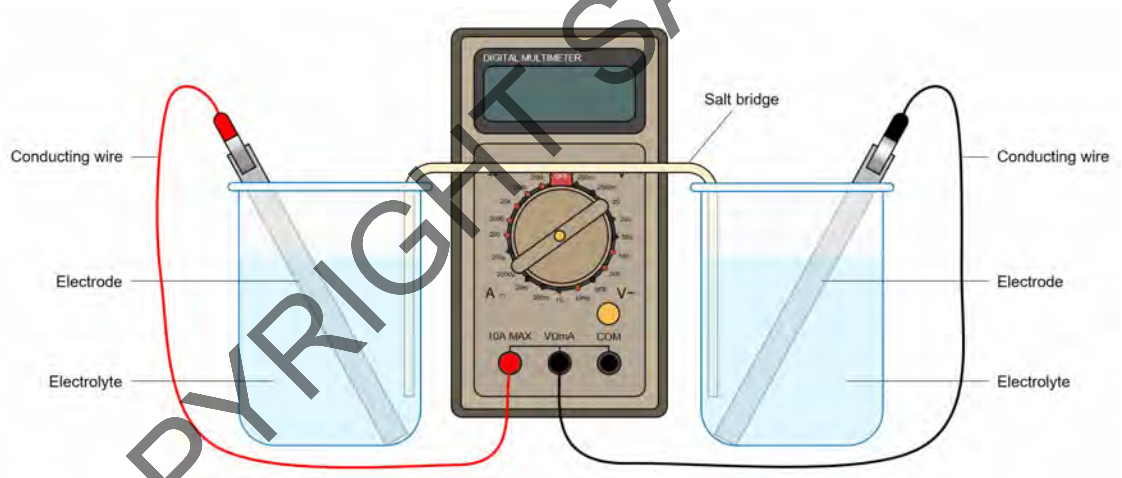


Figure 6.05: Components of a simple galvanic cell.

The function of each component is summarised in the table below.

Component	Function
Electrode	An electrical conductor that allows electrons to flow either towards or away from the electrolyte.
Electrolyte	An electrically-conductive solution that participates in a redox reaction.
Conducting wires	Permit the flow of electrons between the two electrodes.
Salt bridge	Permits the flow of ions between the two electrolytes.

*An electrolytic cell uses a flow of electrons to drive a non-spontaneous redox reaction. The function and properties of electrolytic cells are discussed in Stage 2 chemistry.

Galvanic cell operation

A simple galvanic cell is composed of two **half-cells** that are connected by conducting wires and a salt bridge. The simplest galvanic cells are composed of half-cells containing an electrode submerged in an electrolyte. Each electrode is composed of a reactive metal element and the electrolyte contains cations of the metal element. Each half-cell contains a different metal electrode submerged in a solution of its ions. The more reactive of the two metals is called the **anode** (negative electrode) and the less reactive metal is the **cathode** (positive electrode). Figure 6.06 shows a simple Galvanic cell composed of magnesium and copper half-cells.

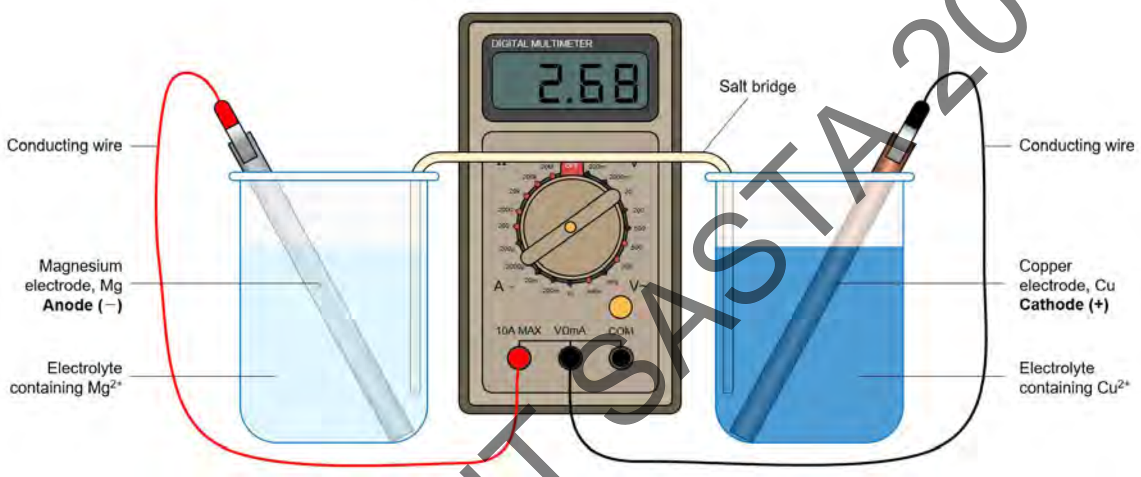


Figure 6.06: Components of a magnesium-copper Galvanic cell.

Magnesium is the anode in this cell as it is more reactive than copper. The magnesium atoms in the anode electrode undergo oxidation when placed in the electrolyte solution.

$$Mg \rightarrow Mg^{2+} + 2e^-$$

The magnesium ions enter the electrolyte solution and the electrons flow through the conducting wires to the cathode half-cell. The cations in the cathode electrolyte undergo reduction as they gain electrons at the surface of the cathode electrode.

$$Cu^{2+} + 2e^- \rightarrow Cu$$

The overall cell reaction is described in the equation below.

$$Cu^{2+} + Mg \rightarrow Cu + Mg^{2+}$$

Positive charge accumulates in the anode electrolyte due to the addition of magnesium ions, and a negative charge accumulates in the cathode electrolyte due to the removal of copper cations. The accumulation of charges in the electrolytes can inhibit the proper functioning of the galvanic cell, and for this reason, a **salt bridge** is placed between the two electrolytes.

The salt bridge contains cations and anions that flow freely into the anode and cathode electrolytes during the normal operation of the galvanic cell. Anions from the salt bridge flow to the anode electrolyte to neutralise positive charge and cations from the salt bridge flow to the cathode electrolyte to neutralise negative charge (Figure 6.07). The flow of ions from the salt bridge completes the electric circuit between the anode and cathode half-cells in a galvanic cell.

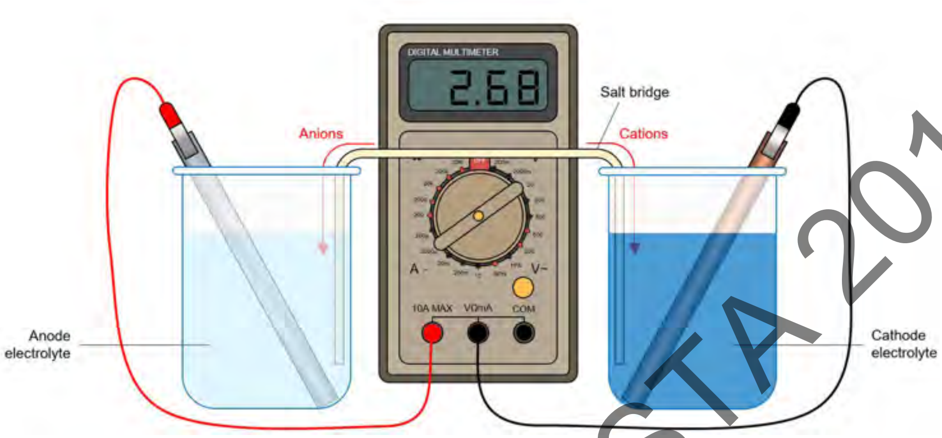


Figure 6.07: Flow of ions from the salt bridge.

An **ammeter** or **voltmeter** is used to measure the **electric current** or **potential difference** in a galvanic cell. An ammeter measures electric current which is the rate at which electrons flow from the anode to the cathode through the conducting wires, and a voltmeter measures the potential difference* between the anode and cathode.

The properties of a galvanic cell composed of half-cells containing two different metals are summarised in the table below.

Component	Summary
Anode	More reactive metal
Cathode	Less reactive metal
Negative electrode	Anode
Positive electrode	Cathode
Anode reaction	Oxidation
Cathode reaction	Reduction
Electron flow	Anode to cathode
Ion flow from salt bridge	Anions to anode and cations to cathode

*The oxidation and reduction reactions occurring at the anode and cathode are associated with an **electric potential** which is a measure of the electrical energy per unit charge at the interface between the electrode and electrolyte. There is an electric potential difference between the two electrodes provided that different reactions are occurring at the anode and cathode. This is why two different metals are required to produce an electric potential difference in a simple galvanic cell. The difference in electric potential between the half-cells is what drives electrons in the electrodes and conducting wires to flow from the anode to the cathode. A voltmeter measures the difference in electric potential between the two electrodes.

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

- Compare the operation of different types of batteries.

A **battery** is an assembly of two or more galvanic cells that is capable of transforming chemical energy into electrical energy. The primary function of batteries is to store and release electrical energy that can be used to power electronic devices. The oxidation and reduction reactions occurring in a battery produce an electric potential difference that is used to drive 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 6.07–6.09.

Example 6.07

Figure 6.12 shows the structure of a mercury battery.

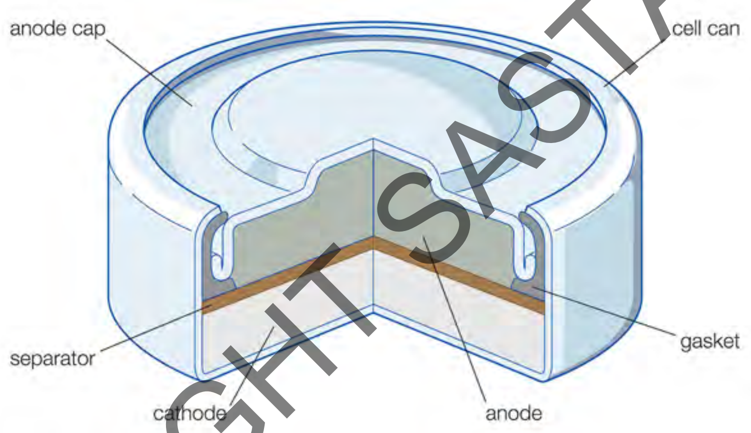


Figure 6.12: Components of a mercury battery.

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

$$Zn + 2OH^- \rightarrow ZnO + H_2O + 2e^-$$

The cathode is 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.

$$HgO + H_2O + 2e^- \rightarrow Hg + 2OH^-$$

The overall cell reaction is:

$$HgO + Zn \rightarrow Hg + ZnO$$

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

Example 6.08

Figure 6.13 shows the structure of an alkaline battery.

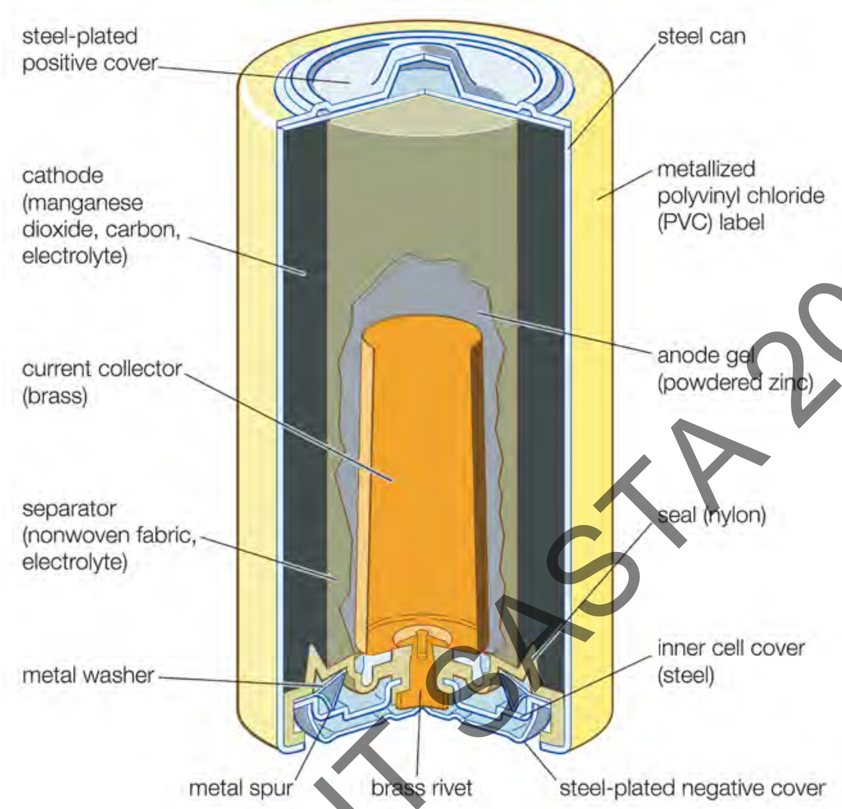


Figure 6.13: Components of an alkaline battery.

The anode is composed of zinc powder in a gel containing potassium hydroxide. Zinc is oxidised at the anode according to the half-equation below.

$$Zn + 2OH^- \rightarrow ZnO + H_2O + 2e^-$$

The cathode is a mixture of manganese dioxide with powdered graphite to increase the electrical conductivity. Manganese dioxide is reduced at the cathode according to the half-equation below.

$$2MnO_2 + H_2O + 2e^- \rightarrow Mn_2O_3 + 2OH^-$$

The overall cell reaction is:

$$2MnO_2 + Zn \rightarrow Mn_2O_3 + ZnO$$

The overall cell reaction produces a potential difference of approximately 1.5 volts between the anode and cathode which is sufficient to power many household items such as remote controls, portable radios, toys, and clocks.

Example 6.10

Figure 6.15 shows the components of a nickel-cadmium battery.

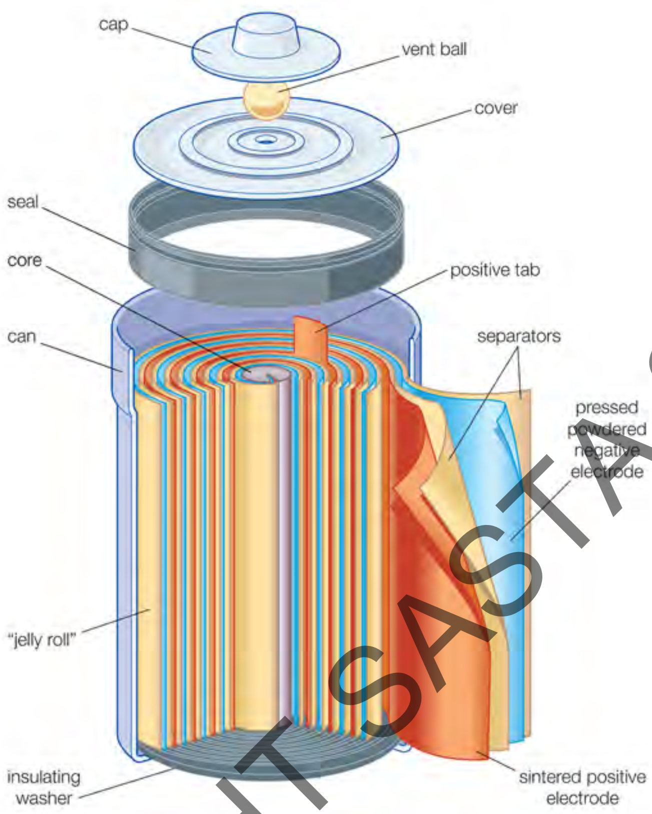


Figure 6.15: Components of a nickel-cadmium battery.

The anode is composed of powdered cadmium that is contact with an alkaline electrolyte (OH^-). Cadmium is oxidised when the cell is discharging according to the half-equation below.

$$Cd + 2OH^- \rightarrow Cd(OH)_2 + 2e^-$$

The cathode is composed of nickel oxide hydroxide which is reduced when the cell is discharging according to the half-equation below.

$$2NiO(OH) + 2H_2O + 2e^- \rightarrow 2Ni(OH)_2 + 2OH^-$$

The overall reaction when the cell is discharging is:

$$2NiO(OH) + 2H_2O + Cd \rightarrow 2Ni(OH)_2 + Cd(OH)_2$$

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

$$2Ni(OH)_2 + Cd(OH)_2 \rightarrow 2NiO(OH) + 2H_2O + Cd$$

Example 6.11

Figure 6.16 shows the components of a lead-acid battery.

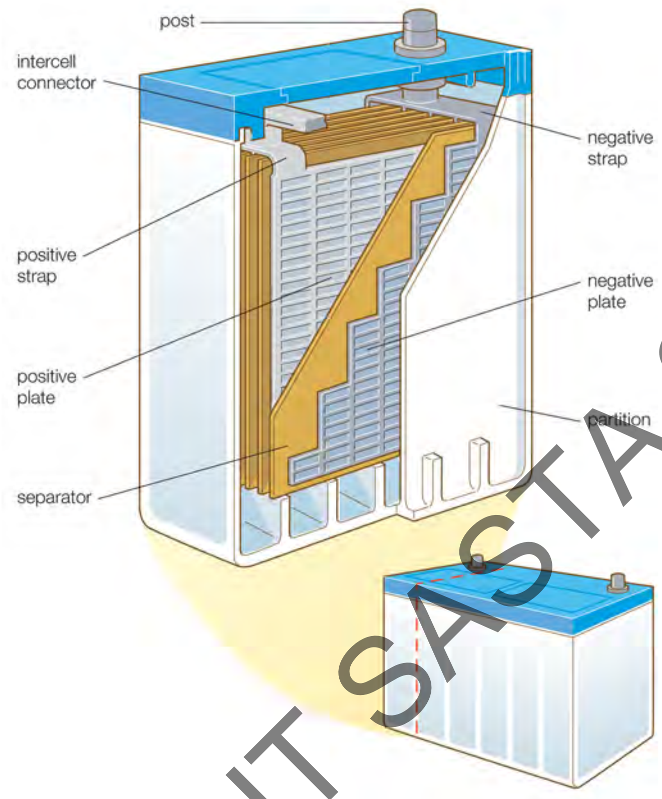


Figure 6.16: Components of a lead-acid battery.

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

$$Pb + SO_4^{2-} \rightarrow PbSO_4 + 2e^-$$

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

$$PbO_2 + SO_4^{2-} + 4H^+ + 2e^- \rightarrow PbSO_4 + 2H_2O$$

The overall reaction when the cell is discharging is:

$$PbO_2 + 2SO_4^{2-} + 4H^+ + Pb \rightarrow 2PbSO_4 + 2H_2O$$

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

$$2PbSO_4 + 2H_2O \rightarrow PbO_2 + 2SO_4^{2-} + 4H^+ + Pb$$

Question 119

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 is discharging.

	$LiC_6 \rightarrow C_6 + Li^+ + e^-$

	$CoO_2 + Li^+ + e^- \rightarrow LiCoO_2$

(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.

	$Li^+ + e^- + LiCoO_2 \rightarrow Li_2O + CoO$

(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

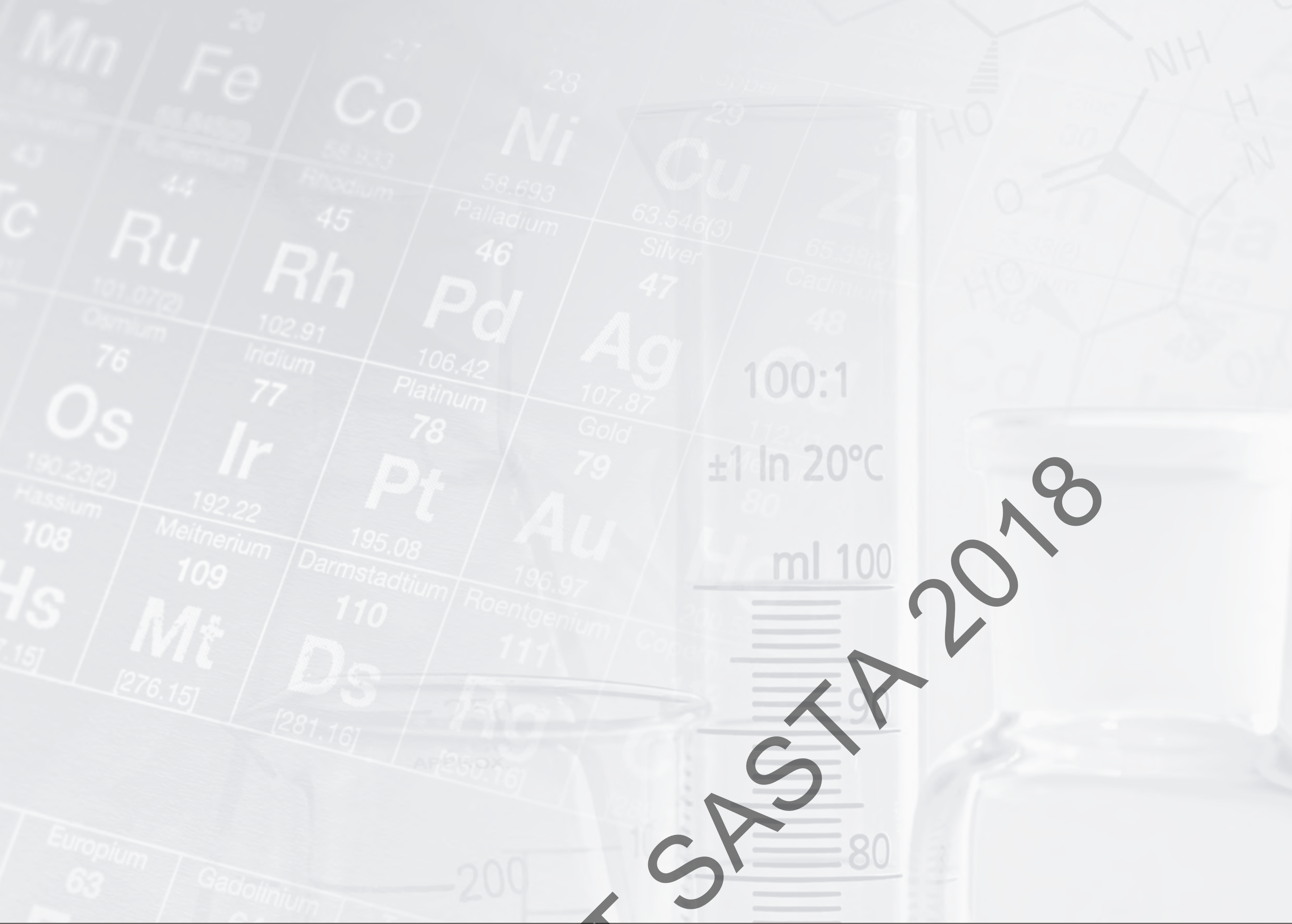
(d) Lithium-ion batteries are rechargeable.

(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

Solutions: Chapter One				
Question	Part		Author's response	Marks
1	(a)	(1)	The minimum temperature at which the state of a material changes from solid to liquid.	1
		(2)	To ensure that the material does not melt at normal operating temperature.	1
	(b)	(1)	Materials with high thermal conductivity are used to maximise the rate at which heat is transferred to the coolant and prevent the cylinders from overheating.	1
		(2)	The water in contact with the cylinders would boil and evaporate at normal operating temperature (>100^{o}C) which would cause the engine to overheat due to coolant loss.	1
2	(a)		Thermal insulators reduce the rate of heat transfer by conduction between the interior and exterior of a building.	1
			This minimises the costs associated with 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 material.	1
	(c)		Isopentane has a low boiling point meaning that it expands and forms a gas readily when heated by energy from the reaction.	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
		(2)	**Any one:** To prevent electric shock when handling the wire; To prevent charge from leaking away from the device during operation; To prevent the wire from corrosion over time.	1
	(b)		High electrical conductivity allows electrons to flow through the heating element where electrical energy is transformed 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)	Solubility is a measure of the degree to which a solute such as PVOH dissolves in a solvent such as water.	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 which reduces the possibility of transmitting infection.	1
		(3)	Detergents must dissolve and form a solution in water that is capable of removing dirt, stains, bacteria, and bodily fluids from contaminated clothing.	1
	(b)		The sutures are water-soluble and dissolve naturally over time as the wound heals.	1
			This saves the surgeon and patient time and money and minimises the risk of infection associated with the surgery that would otherwise be used to remove the sutures after the wound has healed.	1

5	(a)	(1)	Very high surface area to volume ratio.	1
		(2)	The silver nanoparticle catalyst increases the rate of formation of phenol;	1
			More phenol can be produced in a shorter period which reduces operating costs and increases the amount of product that may be sold to manufacturers of polycarbonate plastics for profit.	1
	(b)	(1)	Silver nanoparticles kill infectious bacteria and fungi which reduces the risk of infection associated with open wounds.	1
		(2)	**Advantage:** Can wash clothes in cold water which eliminates the need to heat water using electrical energy; No need to use detergents to sterilise clothing which reduces emissions of harmful detergents into waterways.	1
			Disadvantage: Silver nanoparticles could potentially endanger aquatic organisms if introduced to waterways in wastewater; Silver nanoparticles could potentially kill helpful bacteria used in waste-water treatment facilities.	1
6	(a)		The individual pores in the material are less than 100 nm in diameter.	1
	(b)		Very high surface area to volume ratio provides many surfaces for the adsorption of gas molecules.	1
	(c)		The highly porous structure of MOF-5 increases the capacity for storing hydrogen gas in the fuel tank of a car;	1
			More fuel can be stored in the car which increases the distance the owner of a car can travel before having to refuel.	1
	(d)		MIL-101 could be mixed with pollutant gases produced in power plants and factories before they are released into the environment.	1
			MIL-101 can potentially absorb a high concentration of the pollutant gases which reduces emissions into the environment.	1
7	(a)	(1)	The presence of CNTs reduces the chemical reactivity of the paint which minimises the rate of corrosion.	1
		(2)	CNTs reduce the mass of the car which reduces the energy required to move the car between two locations.	1
			The rate at which fuel is consumed decreases which minimises fuel costs to the owner and reduces the rate at which pollutants are emitted into the environment.	1
	(b)	(1)	Low solubility in water ensures that buckyballs do not dissolve in bodily fluids which are mostly water.	1
			If the buckyballs dissolved, the anticancer drugs would be released into the body before they reach a tumour which would minimise their effectiveness.	1
		(2)	Must ensure that buckyballs are non-toxic and do not cause negative health effects.	1

Solutions: Review Test 6				
Question	Part		Author's response	Marks
1	(a)		$FeSO_4$: +2	1
			$FeCl_3$: +3	1
			K_2FeO_4: +6	1
	(b)	(1)	Oxidation states: +3 → 0 (Fe); +2 → +4 (C) $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$	1+1+1+1
		(2)	Carbon monoxide gains oxygen in the reaction.	1
	(c)	(1)	$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	2
			$Fe^{2+} \rightarrow Fe^{3+} + e^-$	2
		(2)	$6Fe^{2+} \rightarrow 6Fe^{3+} + 6e^-$	1
			$Cr_2O_7^{2-} + 14H^+ + 6Fe^{2+} \rightarrow 2Cr^{3+} + 7H_2O + 6Fe^{3+}$	2
	(d)		Fe^{2+}	1
	(e)		$Cr_2O_7^{2-}$	1
2	(a)	(1)	Sodium displaces aluminium in the reaction.	1
		(2)	Na	1
		(3)	The colour becomes less bright as the concentration of sodium atoms decreases over time due to the displacement reaction with aluminium oxide.	1
	(b)	(1)	$2Na + 2H_2O \rightarrow 2NaOH + H_2$	2
		(2)	$Na \rightarrow Na^+ + e^-$	2
		(3)	$2H_2O + 2e^- \rightarrow 2OH^- + H_2$	2
		(4)	Hydrogen	1