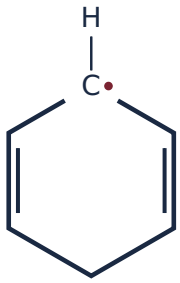
CAMBRIDGE O-LEVEL · CHEMISTRY 6092

O-Level Chemistry

Complete Personalised Revision Notes — Chapters 1 to 22

Compiled and structured personal revision notes, covering the full O-Level Chemistry syllabus from Experimental Chemistry through to Maintaining Air Quality, for end-to-end self-study.



cyclohexadienyl radical

Table of Contents

The Periodic Table of Elements

Group																	
1	2											13	14	15	16	17	18
							1 H hydrogen 1										2 He helium 4
3 Li lithium 7	4 Be beryllium 9											5 B boron 11	6 C carbon 12	7 N nitrogen 14	8 O oxygen 16	9 F fluorine 19	10 Ne neon 20
11 Na sodium 23	12 Mg magnesium 24	3	4	5	6	7	8	9	10	11	12	13 Al aluminium 27	14 Si silicon 28	15 P phosphorus 31	16 S sulfur 32	17 Cl chlorine 35.5	18 Ar argon 40
19 K potassium 39	20 Ca calcium 40	21 Sc scandium 45	22 Ti titanium 48	23 V vanadium 51	24 Cr chromium 52	25 Mn manganese 55	26 Fe iron 56	27 Co cobalt 59	28 Ni nickel 59	29 Cu copper 64	30 Zn zinc 65	31 Ga gallium 70	32 Ge germanium 73	33 As arsenic 75	34 Se selenium 79	35 Br bromine 80	36 Kr krypton 84
37 Rb rubidium 85	38 Sr strontium 88	39 Y yttrium 89	40 Zr zirconium 91	41 Nb niobium 93	42 Mo molybdenum 96	43 Tc technetium –	44 Ru ruthenium 101	45 Rh rhodium 103	46 Pd palladium 106	47 Ag silver 108	48 Cd cadmium 112	49 In indium 115	50 Sn tin 119	51 Sb antimony 122	52 Te tellurium 128	53 I iodine 127	54 Xe xenon 131
55 Cs caesium 133	56 Ba barium 137	57–71 lanthanoids	72 Hf hafnium 178	73 Ta tantalum 181	74 W tungsten 184	75 Re rhenium 186	76 Os osmium 190	77 Ir iridium 192	78 Pt platinum 195	79 Au gold 197	80 Hg mercury 201	81 Tl thallium 204	82 Pb lead 207	83 Bi bismuth 209	84 Po polonium –	85 At astatine –	86 Rn radon –
87 Fr francium –	88 Ra radium –	89–103 actinoids	104 Rf rutherfordium –	105 Db dubnium –	106 Sg seaborgium –	107 Bh bohrium –	108 Hs hassium –	109 Mt meitnerium –	110 Ds darmstadtium –	111 Rg roentgenium –	112 Cn copernicium –	113 Nh nihonium –	114 Fl flerovium –	115 Mc moscovium –	116 Lv livermorium –	117 Ts tennessine –	118 Og oganesson –

Key

proton (atomic) number
atomic symbol
name
relative atomic mass

lanthanoids	57 La lanthanum 139	58 Ce cerium 140	59 Pr praseodymium 141	60 Nd neodymium 144	61 Pm promethium –	62 Sm samarium 150	63 Eu europium 152	64 Gd gadolinium 157	65 Tb terbium 159	66 Dy dysprosium 163	67 Ho holmium 165	68 Er erbium 167	69 Tm thulium 169	70 Yb ytterbium 173	71 Lu lutetium 175
actinoids	89 Ac actinium –	90 Th thorium 232	91 Pa protactinium 231	92 U uranium 238	93 Np neptunium –	94 Pu plutonium –	95 Am americium –	96 Cm curium –	97 Bk berkelium –	98 Cf californium –	99 Es einsteinium –	100 Fm fermium –	101 Md mendelevium –	102 No nobelium –	103 Lr lawrencium –

The volume of one mole of any gas is 24 dm^3 at room temperature and pressure (r.t.p.).
The Avogadro constant, $L = 6.02 \times 10^{23}\ mol^{-1}$.

Chapter 1: Experimental Chemistry

Measurement, gas collection/drying, separation techniques, and purity

1.1 | Measurement of Physical Quantities

Measuring Time

A stopwatch is used to measure time. The SI unit for time is the **second (s)**; minutes (min) and hours (h) may also be used. Depending on the type of digital stopwatch, readings can be taken to 2 decimal places (e.g. 20.01 s).

Table 1.1.1 — Precision and uncertainty of digital stopwatches

Precision / s	Uncertainty / s	D.P.	Example reading / s
0.1	±0.1	1	28.1
0.01	±0.01	2	28.00, 28.11

> ⓘ **NOTE — PRECISION VS UNCERTAINTY**
>
> Precision is set by the smallest division on the instrument. Uncertainty of measurement is the range within which the true value could lie. If a stopwatch reading of 28.00 s has an uncertainty of ±0.01 s, the true value lies between 27.99 s and 28.01 s.

Measuring Temperature

A thermometer (alcohol or mercury) is used to measure temperature, read to the nearest 0.5°C. The SI unit is the **kelvin (K)**; degrees Celsius (°C) is also commonly used (K = °C + 273).

A **datalogger connected to a temperature sensor** can also be used. Advantages over a thermometer:

- More accurate than an alcohol/mercury thermometer
- Can record data continuously over a period of time
- Saves data digitally, which can be used to plot graphs/charts

Table 1.1.2 — Precision and uncertainty of laboratory thermometers

Precision / °C	Uncertainty / °C	D.P.	Example / °C
1	±0.5	1	23.0, 23.5

Measuring Length

A metre rule or measuring tape is used. SI unit: **metre (m)**. Other units: mm, cm, dm.

Table 1.1.3 — Precision and uncertainty of a metre rule

Precision / mm	Uncertainty / mm	D.P.	Example / mm
1	±0.5	1	20.5

Measuring Mass

An electronic (mass) balance is used, recorded to the nearest 0.001 g depending on model. SI unit: **kilogram (kg)**; grams (g) and milligrams (mg) also used.

Table 1.1.4 — Precision and uncertainty of an electronic balance

Precision / g	Uncertainty / g	D.P.	Example / g
0.001	±0.001	3	2.253

Measuring Volume of Liquids/Solutions

Table 1.1.5 — Instruments used to measure volume of liquids

Apparatus	Precision / cm^3	Uncertainty / cm^3	D.P.	Primary use
Measuring cylinder	1	±0.5	1	Approximate volumes of liquids/solutions
Pipette	—	±0.1	1	Accurate, fixed volumes (e.g. 25.0, 10.0 cm^3)
Burette	0.1	±0.05	2	Very precise, variable volumes (e.g. 25.45 cm^3)
Volumetric flask	—	—	—	Accurate, large fixed volumes (e.g. 100, 250 cm^3)

> ⓘ **NOTE — PARALLAX ERROR**
>
> When reading a burette or measuring cylinder, position your eye level with the meniscus to avoid parallax error. For coloured liquids, read from the top of the meniscus if the bottom cannot be seen.

To measure the volume of a **gas**, a **gas syringe** is used (typically up to 100 cm^3).

1.2 | Collection and Drying of Gases

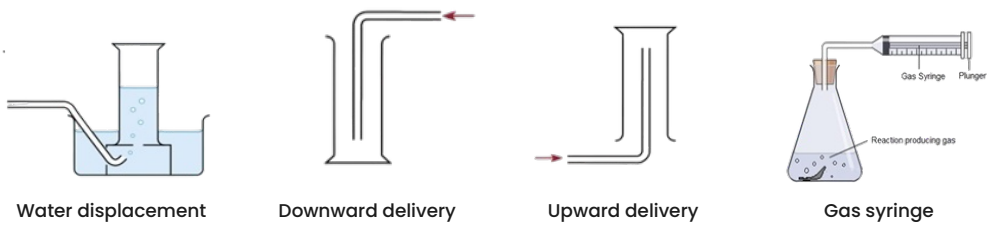


Fig 1.2.1–1.2.4. *The four main methods of collecting gases, chosen by density and water-solubility of the gas.*

Table 1.2.1 — Methods of collecting gases

Method	Used for	Description	Example gases
Water displacement (downward displacement of water)	Insoluble/slightly soluble gases (density irrelevant)	Gas is bubbled into an inverted, water-filled test tube/gas jar; the gas displaces water as it collects	O_2, H_2, CO_2
Downward delivery (upward displacement of air)	Gases denser than air	Gas sinks into the (unsealed) gas jar and pushes air out from the top	Cl_2, HCl, SO_2, NO_2
Upward delivery (downward displacement of air)	Gases less dense than air	Gas rises into the (unsealed) gas jar and pushes air out from the bottom	NH_3, H_2
Gas syringe	Any gas — also measures volume	Gas produced pushes the plunger outward; volume read directly off the syringe	Any

> ⓘ **NOTE**
>
> The approximate relative molecular mass (M_r) of air is 29. Gases with $M_r > 29$ are collected by downward delivery; gases with $M_r < 29$ by upward delivery.

Table 1.2.2 — Density, solubility and collection methods of common gases

Gas	Density vs air	Solubility in water	Method(s) of collection
Hydrogen, H_2	Less dense	Insoluble	Upward delivery & water displacement
Ammonia, NH_3	Less dense	Highly soluble	Upward delivery
Oxygen, O_2	Slightly denser	Slightly soluble	Downward delivery & water displacement
Hydrogen chloride, HCl	Denser	Highly soluble	Downward delivery
Carbon dioxide, CO_2	Denser	Slightly soluble	Downward delivery & water displacement
Nitrogen dioxide, NO_2	Denser	Highly soluble	Downward delivery
Sulfur dioxide, SO_2	Denser	Highly soluble	Downward delivery
Chlorine, Cl_2	Denser	Highly soluble	Downward delivery

Drying Gases

Gases are classified as **neutral** (O_2, N_2, H_2), **acidic** (HCl, CO_2, NO_2, SO_2) or **alkaline** (NH_3) — a drying agent must not react with the gas being dried.

Table 1.2.3 — Drying agents

Drying agent	Suitable for	Not suitable for	Reason unsuitable
Concentrated sulfuric acid	Neutral, acidic gases	Alkaline gases (NH_3)	Reacts with NH_3 (neutralisation/acid-base reaction)
Quicklime, CaO (freshly heated)	Neutral, alkaline gases	Acidic gases	Reacts with acidic gases (neutralisation)
Fused (anhydrous) calcium chloride (freshly heated)	Neutral, acidic gases	NH_3	Anhydrous $CaCl_2$ reacts with NH_3 to form a complex, $CaCl_2.8NH_3$

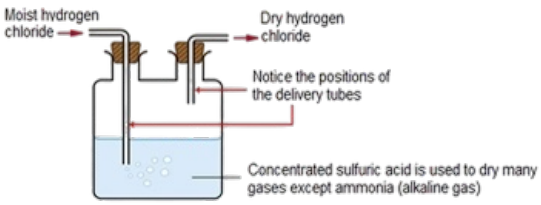


Fig 1.2.5. *Drying a gas using concentrated sulfuric acid.*

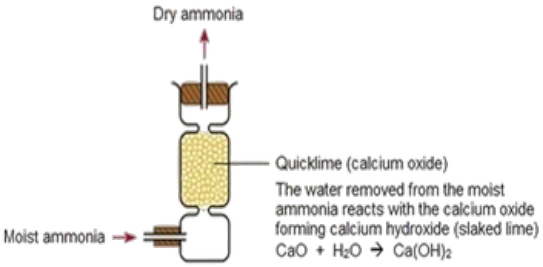


Fig 1.2.6. *Drying a gas using quicklime (calcium oxide).*

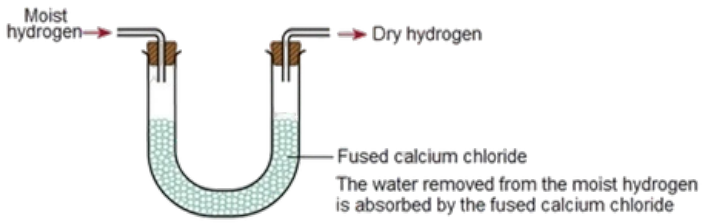


Fig 1.2.7. *Drying a gas using fused (anhydrous) calcium chloride.*

KEY QNS 1.2.1 — DRYING HYDROGEN

Q. Concentrated sulfuric acid should not be used to dry hydrogen gas even though hydrogen is neutral. Suggest why.
A. When concentrated sulfuric acid absorbs water vapour from the gas, a large amount of heat is given out. This heat could ignite the hydrogen gas, which is highly flammable.

ⓘ NOTE

Quicklime (CaO) absorbs both moisture and CO_2 from air and so must be freshly heated (to drive off absorbed water/CO_2) immediately before use. The same applies to fused calcium chloride, which readily absorbs moisture from the air.

1.3 | How Are Substances in Mixtures Separated?

DEFINITION 1.3.1 — MIXTURE

A mixture is a substance made up of two or more substances that are *not* chemically combined (i.e. physically combined, in no fixed ratio).

DEFINITION 1.3.2 — PURE SUBSTANCE

A pure substance is made up of only one element or compound; it is not mixed with any other substance.

Table 1.3.1 — Methods of separating different types of mixtures

Solid–Solid	Solid–Liquid	Liquid–Liquid	Gaseous–Gaseous
Magnetic attraction	Filtration (insoluble solid)	Separating funnel (immiscible)	Fractional distillation (after liquefying)
Sieving	Evaporation to dryness	Chromatography (miscible, different Rf)	—
Using a suitable solvent	Crystallisation	Fractional distillation (different b.p.)	—
Sublimation	Simple distillation	—	—

Solid–Solid Mixtures

- **Magnetic attraction** — separates a magnetic solid (e.g. iron filings) from a non-magnetic solid using a magnet.
- **Sieving** — separates solids of different particle sizes using a sieve/mesh.
- **Using a suitable solvent** — one solid dissolves in the solvent (e.g. water for salt, ethanol for iodine) while the other does not; the mixture is then filtered and the dissolved solid recovered by crystallisation/evaporation.
- **Sublimation** — separates a solid that sublimes (e.g. iodine, ammonium chloride) from one that does not, by heating the mixture so only the subliming solid turns to vapour and re-solidifies elsewhere.

Solid–Liquid Mixtures

- **Filtration** — separates an insoluble solid (residue) from a liquid (filtrate) using filter paper.
- **Evaporation to dryness** — recovers a dissolved solid by heating the solution until all the solvent evaporates. Suitable only if the solid does not decompose on strong heating.
- **Crystallisation** — the solution is heated until saturated (a hot, concentrated solution), then allowed to cool slowly so pure crystals of the solid form; used when the solid decomposes on strong/continuous heating, or when a well-formed crystalline solid is desired.
- **Simple distillation** — recovers the solvent (e.g. pure water) from a solution, based on differences in boiling point between solvent and dissolved solid.

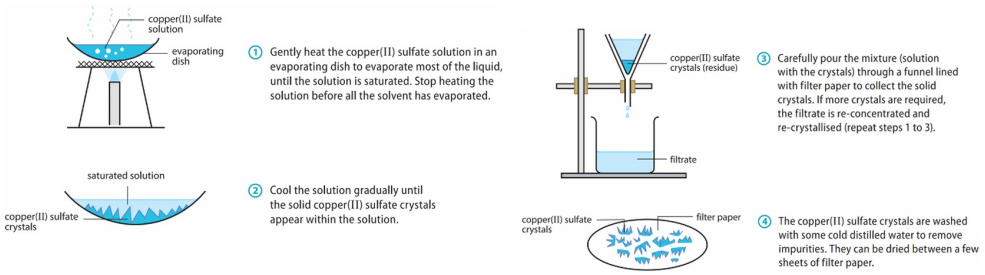


Fig 1.3.2. *Crystallisation of copper(II) sulfate from solution.*

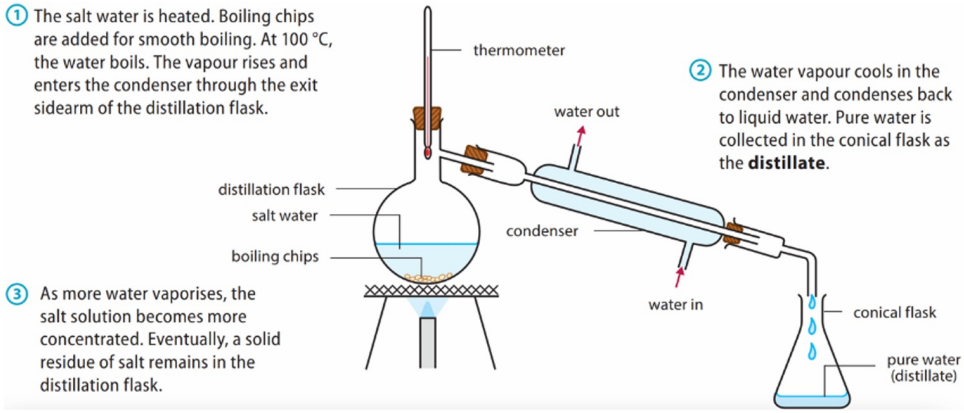
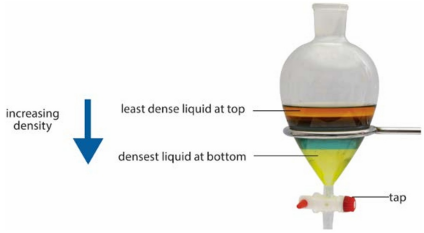


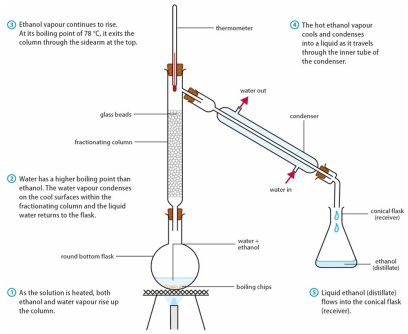
Fig 1.3.3. *Simple distillation, used to recover pure water from saltwater.*

Liquid–Liquid Mixtures

- **Separating funnel** — separates immiscible liquids of different densities (e.g. oil and water); the denser liquid is run off from the bottom tap.
- **Fractional distillation** — separates miscible liquids with different boiling points (e.g. ethanol and water) using a fractionating column, which provides a temperature gradient so that a series of vaporisation-condensation cycles occur, allowing separation of components with close boiling points.
- **Chromatography** — separates miscible substances (in solution) with different solubilities/degrees of adsorption to a stationary phase, and hence different R_f values.



Separating funnel



Fractional distillation

Fig 1.3.4–1.3.5. *Separating funnel used to split immiscible liquids, and a fractional distillation setup used to separate miscible liquids (e.g. ethanol and water) by boiling point.*

DEFINITION 1.3.3 — R_F VALUE

R_f = (distance travelled by solute) ÷ (distance travelled by solvent). R_f values are always between 0 and 1, and are unique to a substance in a given solvent, allowing identification by comparison with known/reference substances.

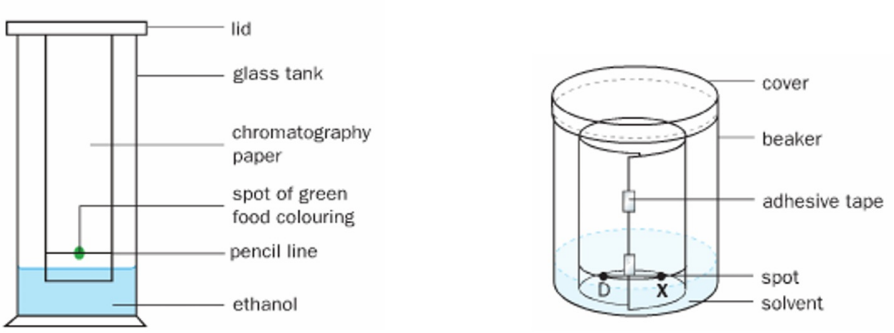


Fig 1.3.6. *Two setups for paper chromatography, used to separate mixtures such as food colourings or amino acids.*

WORKED EXAMPLE — CHROMATOGRAPHY

A dye mixture is spotted onto chromatography paper below the solvent line (pencil, not ink, is used as pencil marks do not dissolve/run). The paper is dipped in solvent (below the spots) in a covered container (to keep the atmosphere saturated with solvent vapour and prevent evaporation). As the solvent rises up the paper, components separate based on differing solubility in the solvent and affinity for the paper. Number of spots = number of components; matching R_f values/spot positions to known substances allows identification.

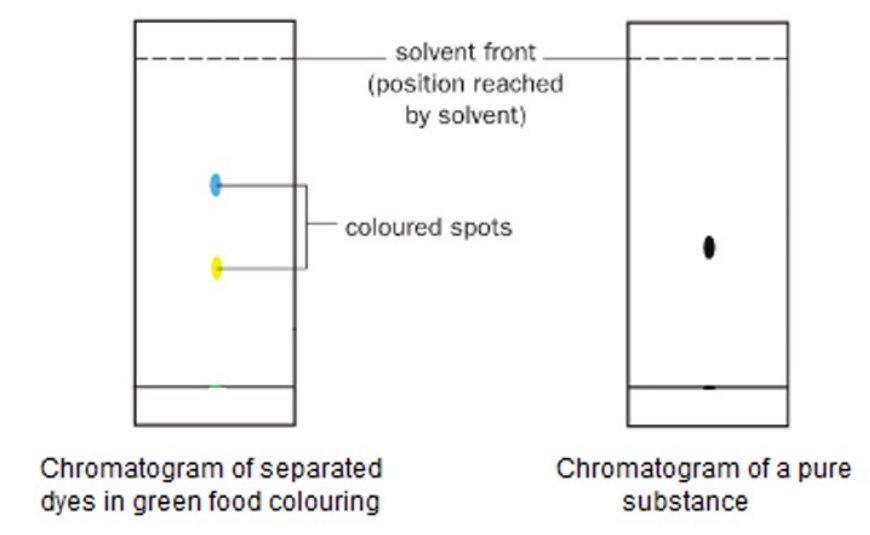


Fig 1.3.7. *Comparing the chromatogram of a pure substance (single spot) against a mixture of dyes (multiple spots, each with its own R_f).*

Gaseous–Gaseous Mixtures

Fractional distillation of liquid air — air is first cooled and compressed until it liquefies, then fractionally distilled; the components (mainly N_2 and O_2) are separated according to their different boiling points (N_2 b.p. −196°C, O_2 b.p. −183°C), with nitrogen (lower b.p.) boiling off first.

1.4 | Determining Purity of Substances

A pure substance has a **fixed (sharp) melting and boiling point**. An impure substance melts/boils over a **range of temperatures**, and impurities generally **lower the melting point** and **raise the boiling point** compared to the pure substance.

Table 1.4.1 — Melting/boiling point behaviour: pure vs impure substances

	Pure substance	Impure substance (with soluble impurity)
Melting point	Fixed, sharp; melts at a single temperature	Lower than pure substance; melts over a range of temperatures
Boiling point	Fixed, sharp; boils at a single temperature	Higher than pure substance; boils over a range of temperatures

ⓘ NOTE

Melting point and boiling point determination is used to check the purity of a substance obtained after separation/purification — the closer the observed melting/boiling point (and the sharper the range) is to the known value for the pure substance, the purer the sample.

COMMON EXAM PITFALL

"Pure" in chemistry has a stricter meaning than in everyday language (e.g. "pure orange juice" is a mixture in the chemistry sense). Always define purity in terms of a single element/compound not mixed with any other substance, and justify using melting/boiling point behaviour rather than taste, appearance, etc.

CHAPTER 1 QUICK RECAP: Choose apparatus based on required precision (mass > volume > length ≈ temp > time in general precision); gas collection method depends on density (vs air, M_r=29) and solubility; drying agents must not react with the gas; separation technique depends on the states and properties (particle size, magnetism, solubility, boiling point) of the mixture's components; purity is confirmed via sharp, fixed melting/boiling points.

Chapter 2: Kinetic Particle Theory

States of matter, changes of state, and particle movement

2.1 | How Are Solids, Liquids and Gases Different?

DEFINITION 2.1.1 — KINETIC PARTICLE THEORY

All matter is made up of tiny particles (atoms, molecules or ions) that are in constant, random motion. These particles have mass and occupy space — this is why matter has mass and occupies space.

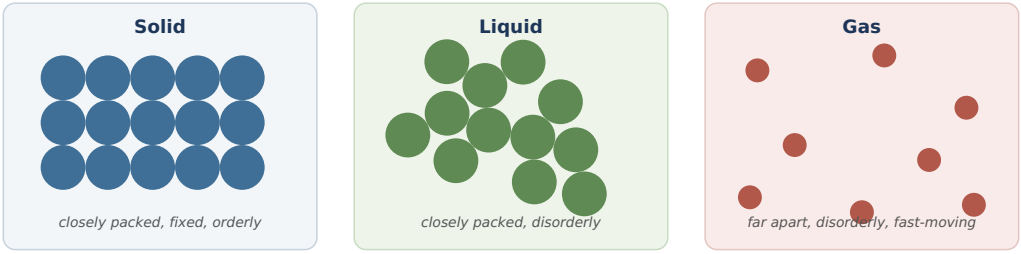


Fig 2.1.1. *Particle arrangement, spacing and movement in solids, liquids and gases.*

Table 2.1.1 — Distance, arrangement and movement of particles in the three states

State	Attractive forces	Arrangement / spacing	Movement
Solid	Very strong	Very closely packed; orderly arranged, fixed positions	Particles vibrate and rotate about fixed positions only
Liquid	Strong	Closely packed; disorderly arranged	Particles slide past one another freely throughout the liquid
Gas	Weak	Spaced far apart; disorderly arranged	Particles move randomly, freely and rapidly in any direction, occupying any available space

KEY QNS 2.1.1 — DENSITY OF BROMINE

Q. Why does liquid bromine have a much higher density than gaseous bromine?
A. In the liquid state, bromine molecules are closely packed together; in the gaseous state, molecules are far apart. In the same unit volume, there are far more bromine particles in the liquid than in the gas. Thus the total mass per unit volume (density) of liquid bromine is much higher than that of gaseous bromine.

DEFINITION 2.1.2 — TEMPERATURE

Temperature is a measure of the average kinetic energy of the particles in a substance.

Heating: thermal energy is transferred to the substance; some is converted to kinetic energy of particles, raising temperature.
Cooling: thermal energy is transferred from the substance to the surroundings; kinetic energy of particles decreases, lowering temperature.

KEY QNS 2.1.2 — CHANGES OF STATE

Q. Why does temperature remain constant during a change of state?
A. The average kinetic energy of the particles does not change during a change of state. Instead, the thermal energy supplied (or removed) is used entirely to overcome (or strengthen/re-form) the attractive forces between particles, not to change particle speed.

The Six Changes of State

Table 2.1.2 — Definitions and particle explanations of the six changes of state

Change	Definition	Particle explanation	Heat
Melting (solid → liquid, at m.p.)	A solid changes into a liquid at the melting point	Particles vibrate faster on heating and gain enough energy to overcome the forces holding them in fixed positions, breaking away to move around one another	Absorbed (endothermic)
Freezing (liquid → solid, at f.p.)	A liquid changes into a solid at the freezing point	Particles lose kinetic energy on cooling and eventually settle into fixed, orderly positions	Released (exothermic)
Evaporation (liquid → gas, below b.p.)	A liquid changes into a gas at temperatures below the boiling point, at the liquid surface only	Surface particles with sufficient energy overcome attractive forces and escape into the air as vapour	Absorbed (endothermic)
Boiling (liquid → gas, at b.p.)	A liquid changes into a gas throughout the liquid, at the boiling point	Particles gain enough energy to completely overcome attractive forces; particles become far apart and move freely in all directions	Absorbed (endothermic)
Condensation (gas → liquid)	A gas changes into a liquid	Kinetic energy is transferred to surroundings; spacing between particles reduces until close enough to form a liquid	Released (exothermic)
Sublimation (solid → gas, no liquid stage)	A solid changes directly into a gas without passing through the liquid state	Surface particles gain enough kinetic energy to completely overcome attractive forces and escape directly as gas	Absorbed (endothermic)
Vapour deposition (gas → solid, no liquid stage)	A gas changes directly into a solid without passing through the liquid state	Gas particles cool, lose kinetic energy and arrange directly into a solid	Released (exothermic)

Table 2.1.3 — Evaporation vs boiling

Evaporation	Boiling
Occurs only at the surface of the liquid	Occurs throughout the liquid
Occurs at any temperature	Occurs only at the boiling point
A slow process	A fast process

Expansion and Contraction of Solids

Below the melting point, heating a solid does not let particles spread out freely (still strongly attracted); instead, particles vibrate faster with a slightly greater spacing — the solid **expands**. On cooling, particles vibrate slower and move closer together — the solid **contracts**. (Real-life applications: expansion gaps in bridges/railway tracks, bimetallic strips in thermostats.)

2.2 | How Do Particles Move?

> **DEFINITION 2.2.1 — DIFFUSION**
> Diffusion is the net movement of particles from a region of higher concentration to a region of lower concentration.

Simple diffusion occurs mainly in liquids and gases (particles move freely); vacancy/interstitial diffusion occurs (much more slowly) in solids such as metals and alloys.

Everyday examples of diffusion: the spread of perfume or cooking aromas through a room (diffusion in a gas); tea or coffee grains/sugar spreading through water without stirring (diffusion in a liquid).

Table 2.2.1 — Factors affecting rate of diffusion

Condition	Effect on rate	Reason
Increase temperature	Increases	More thermal energy is converted to kinetic energy of the particles, so they move faster
Increase M_r	Decreases	Particles with greater mass require more kinetic energy to move at the same speed, so heavier particles diffuse more slowly at a given temperature

WORKED EXAMPLE — POROUS POT EXPERIMENT (H_2 VS AIR)

A porous pot connected to water via a tube is surrounded by hydrogen gas. Since M_r of H_2 (2) is much lower than the approximate M_r of air (29), hydrogen diffuses into the pot faster than air diffuses out. This raises the pressure inside the porous pot above atmospheric pressure, pushing the water level down at the open end of the tube.

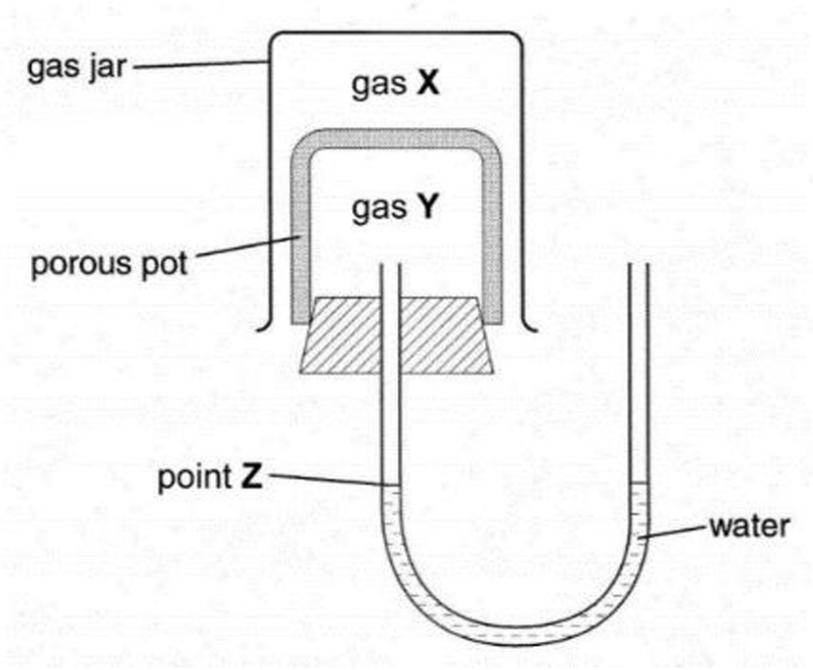


Fig 2.2.1. *Porous pot experiment comparing the rate of diffusion of hydrogen and air.*

WORKED EXAMPLE — DIFFUSION OF IONS IN A GEL (KI + $PB(NO_3)_2$)

Crystals of potassium iodide (KI) and lead(II) nitrate, $Pb(NO_3)_2$, are placed at opposite ends of a gel/agar strip soaked in water. Both dissolve and their ions diffuse from a region of higher to lower concentration. A yellow precipitate of lead(II) iodide, PbI_2, forms where the two ions meet — closer to the lead(II) nitrate crystal, because Pb^{2+} (M_r = 207) is heavier than I^- (M_r = 127) and so diffuses more slowly.

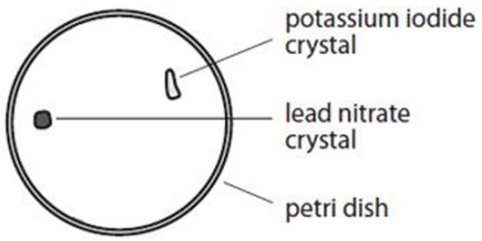


Fig 2.2.2. *Diffusion of K^+/I^- and Pb^{2+}/NO_3^- ions through a gel, forming a yellow band of PbI_2 where they meet.*

CHAPTER 2 QUICK RECAP: Solids/liquids/gases differ in the strength of attractive forces and hence particle spacing and movement. Temperature = average KE of particles; it stays constant during a change of state because energy goes into changing bonds, not particle speed. Diffusion rate increases with temperature and decreases with M_r (heavier particles diffuse slower).

Chapter 3: Atomic Structure

Subatomic particles, electronic configuration, ions and isotopes

3.1 | What Is an Atom Made Of?

DEFINITION 3.1.1 — ELEMENT

An element is a pure substance that cannot be broken down into two or more simpler substances by chemical methods. There are 118 known elements, each made of a unique type of atom.

DEFINITION 3.1.2 — ATOM

An atom is the smallest particle that can still have the chemical characteristics of an element.

DEFINITION 3.1.3 — COMPOUND

A compound is a pure substance containing two or more types of elements chemically combined in a fixed ratio.

Atoms are made up of three types of subatomic particles: protons and neutrons (found together in the central **nucleus**, and hence collectively called **nucleons**), and electrons (found in electron shells orbiting the nucleus).

Table 3.1.1 — Relative mass, charge and location of subatomic particles

Subatomic particle	Relative mass / unit	Relative charge	Location
Proton (p)	1	1+	Nucleus
Neutron (n)	1	0	Nucleus
Electron (e^-)	1/1840	1−	Electron shells

KEY QNS 3.1.1 — WHY ARE ATOMS ELECTRICALLY NEUTRAL?

Protons have charge 1+, electrons have charge 1−, neutrons have charge 0. In an atom, the number of protons equals the number of electrons, so the positive and negative charges balance exactly — hence atoms are electrically neutral.

KEY QNS 3.1.2 — WHY DOES ATOMIC MASS COME MAINLY FROM PROTONS AND NEUTRONS?

The mass of an electron (1/1840 unit) is negligible compared to that of a proton or neutron (1 unit each). Hence nearly all the mass of an atom is concentrated in the nucleus — this is why nucleon number is also called "mass number".

3.2 | How Many Subatomic Particles Does an Atom Have?

DEFINITIONS 3.2.1

Proton (atomic) number, Z — the number of protons in an atom's nucleus (= number of electrons in a neutral atom). Unique to each element.
Nucleon (mass) number, A — the total number of protons and neutrons in the nucleus.
Number of neutrons = nucleon number (A) − proton number (Z)

Nuclide Notation

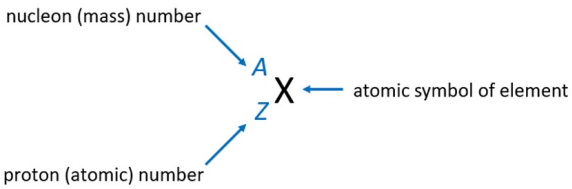


Fig 3.1.1. *Nuclide notation showing nucleon number (top-left) and proton number (bottom-left) of an atom* X.

$$^{A}_{Z}\mathrm{X} \text{ where A = nucleon number (top left), Z = proton number (bottom left)}$$

For example, $^{12}_{6}\mathrm{C}$ represents a carbon atom with 12 nucleons and 6 protons (and therefore 6 neutrons). Other carbon atoms may be represented as $^{13}_{6}\mathrm{C}$ and $^{14}_{6}\mathrm{C}$.

> **WORKED EXAMPLE — DEDUCING PARTICLE NUMBERS**
>
> For $^{40}_{20}\mathrm{Ca}$ (calcium): proton number = 20, so number of electrons = 20 (neutral atom); nucleon number = 40, so number of neutrons = 40 − 20 = 20.

3.3 | How Are Electrons Arranged in an Atom?

Electron shells have different energy levels, the shell closest to the nucleus being lowest in energy. The further a shell is from the nucleus, the higher its energy and the less stable its electrons. **Electrons fill the lowest-energy (innermost) shell first.**

- 1st shell: max. 2 electrons
- 2nd shell: max. 8 electrons
- 3rd shell (as outermost/valence shell, for the first 20 elements): max. 8 electrons

The outermost occupied shell is the **valence shell**; its electrons are **valence electrons**.

> **ⓘ NOTE — PERIOD AND GROUP RELATIONSHIP**
>
> The **Period number** = number of electron shells in the atom.
> The **Group number** = number of valence electrons (for Groups 1–2: valence electrons = group number; for Groups 13–18: valence electrons = group number − 10, except helium which has only 2 valence electrons despite being in Group 18).
> Elements in the same group have the same number of valence electrons and hence similar chemical properties.

Table 3.3.1 — Electronic configuration of Group 18 elements (noble gases)

	Helium	Neon	Argon
Proton number	2	10	18
Period	1	2	3
Electronic configuration	2	2,8	2,8,8
Valence electrons	2	8	8

Noble gas atoms have a completely filled valence shell — a stable electronic configuration — and so do not react spontaneously with other atoms under standard conditions; they exist as single atoms, not molecules or ions.

> **WORKED EXAMPLE — DEDUCING ELECTRONIC CONFIGURATION (SODIUM)**
>
> Sodium: proton number 11 ⇒ 11 electrons. Sodium is in Period 3 ⇒ 3 electron shells. Fill 1st shell (2), 2nd shell (8), remaining 1 electron goes to the 3rd (valence) shell. Electronic configuration: **2, 8, 1** — hence 1 valence electron, confirming Group 1.

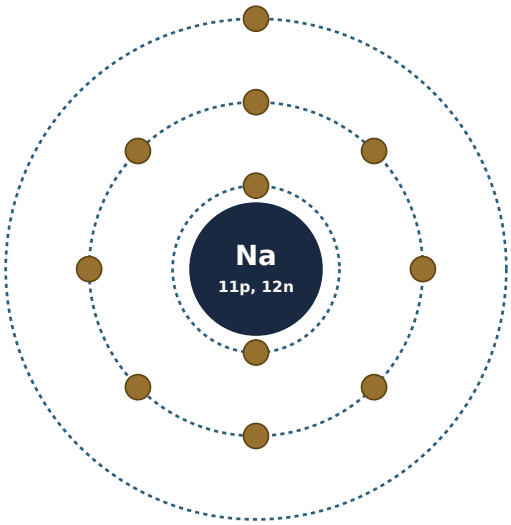


Fig 3.3.1. *Electron shell diagram of sodium: 2, 8, 1 — 1 valence electron in the outermost (3rd) shell.*

WORKED EXAMPLE — DECODING "2, 8, 6"

3 numbers ⇒ 3 shells ⇒ Period 3. Last number (6) = valence electrons ⇒ Group 16. Total electrons = 2+8+6 = 16 ⇒ proton number 16 ⇒ the element is sulfur.

3.4 | Ions

DEFINITION 3.4.1 — ION

An ion is a charged particle formed when an atom (or group of atoms) gains or loses electron(s); the number of protons and neutrons remains unchanged.

An atom without a completely filled valence shell does not have a stable electronic configuration, and will lose, gain, or share valence electrons with other atoms to attain the stable noble gas configuration.

- Losing electron(s) → forms a positive ion (**cation**)
- Gaining electron(s) → forms a negative ion (**anion**)

WORKED EXAMPLE — FORMATION OF A CATION (CA → CA^{2+})

Calcium (proton number 20): electronic configuration 2,8,8,2 — 2 valence electrons, not a full shell. Calcium atom loses 2 electrons to attain the stable configuration 2,8,8 (same as argon), forming Ca^{2+}.

	Ca atom (^{20}Ca)	Ca^{2+} ion
Protons	20	20
Neutrons	20	20
Electrons	20	18
Electronic config.	2,8,8,2	2,8,8
Overall charge	0	2+

WORKED EXAMPLE — FORMATION OF AN ANION (O → O^{2-})

Oxygen (proton number 8): electronic configuration 2,6 — 6 valence electrons. Oxygen atom gains 2 electrons to attain the stable configuration 2,8 (same as neon), forming O^{2-}.

	O atom ($^{16}_{8}$O)	O^{2-} ion
Protons	8	8
Neutrons	8	8
Electrons	8	10
Electronic config.	2,6	2,8
Overall charge	0	2−

ⓘ NOTE

The charge of an ion tells you exactly how many electrons were gained/lost: a 2+ ion has lost 2 electrons; a 2− ion has gained 2 electrons. Protons and neutrons never change when an ion forms.

Table 3.4.1 — Common ions and their chemical formulae

Cations	Formula	Anions	Formula
Silver	Ag^{+}	Hydroxide	OH^{-}
Ammonium	NH_4^{+}	Nitrate	NO_3^{-}
Hydrogen	H^{+}	Ethanoate	CH_3COO^{-}
Zinc	Zn^{2+}	Sulfate	SO_4^{2-}
—	—	Sulfite	SO_3^{2-}
—	—	Carbonate	CO_3^{2-}
—	—	Thiosulfate	$S_2O_3^{2-}$
—	—	Phosphate	PO_4^{3-}
—	—	Manganate(VII)	MnO_4^{-}

ⓘ NOTE

Copper, iron, lead and chromium (transition metals) can form multiple ions of different charge — the Roman numerals (II), (III) refer to 2+ and 3+ charges respectively (e.g. iron(II) = Fe^{2+}, iron(III) = Fe^{3+}).

3.5 | Isotopes

DEFINITION 3.5.1 — ISOTOPES

Isotopes are atoms of the same element with the same number of protons (and electrons) but different numbers of neutrons.

Table 3.5.1 — Isotopes of hydrogen

Isotope	Nuclide	Protons	Neutrons	Electrons	% abundance
Hydrogen-1 (protium)	$_{11}H$	1	0	1	99.99%
Hydrogen-2 (deuterium)	$_{21}H$	1	1	1	0.01%
Hydrogen-3 (tritium)	$_{31}H$	1	2	1	trace, unstable

Properties of Isotopes

- **Chemical properties:** similar, since isotopes have the same number of valence electrons and undergo the same types of reactions.
- **Physical properties:** different, since isotopes have different masses (different numbers of neutrons) — density, melting point and boiling point differ slightly.

Table 3.5.2 — Applications of isotopes (radioisotopes)

Isotope	Application
Technetium-99	Detection of tumours
Iodine-131	Treatment of thyroid disorders
Californium-252	Detection of explosives
Americium-241	Smoke detectors
Carbon-14	Carbon dating (estimating age of once-living things)
Uranium-238	Estimating the age of rocks

Relative Atomic Mass from Isotopic Abundance

$$A_r = \Sigma\ (\%\ \text{abundance of isotope} \times \text{isotopic mass}) \div 100$$

WORKED EXAMPLE — CALCULATING A_R OF PHOSPHORUS

^{30}P, ^{31}P, ^{32}P have abundances 23%, 56%, 21% respectively.
$A_r = (23/100 \times 30) + (56/100 \times 31) + (21/100 \times 32) =$ **31.0** (3 s.f.) — no unit, since A_r is a ratio.

WORKED EXAMPLE — FINDING ABUNDANCE FROM A_R

Sulfur has isotopes ^{32}S and ^{34}S with $A_r = 32.12$. Let x = % abundance of ^{32}S:
$32.12 = (x/100 \times 32) + [(100-x)/100 \times 34]$
$3212 = 32x + 3400 - 34x \Rightarrow 2x = 188 \Rightarrow$ **x = 94%** (^{32}S), so ^{34}S = 6%.

COMMON EXAM PITFALL

When a question asks about the "structure of the nucleus," only compare protons and neutrons — electrons are *not* located in the nucleus and should not be included in that comparison.

CHAPTER 3 QUICK RECAP: Atoms = protons + neutrons (nucleus) + electrons (shells). Proton number = electrons in a neutral atom; nucleon number − proton number = neutrons. Electronic configuration follows 2,8,8 filling order; period = number of shells, group = valence electrons. Ions form when atoms lose/gain electrons to reach a noble gas configuration. Isotopes share proton/electron number but differ in neutrons, giving similar chemical but different physical properties; A_r is a weighted average over isotopic abundances.

Chapter 4: Chemical Bonding — Ionic Bonding

Elements, compounds, mixtures and the ionic lattice

4.1 | Elements, Compounds and Mixtures Revisited

Table 4.1.1 — Forms in which elements exist

Atoms of the same element	Molecules of the same element	Metal lattice + sea of delocalised electrons
e.g. individual noble gas atoms (He, Ne, Ar)	e.g. H_2, N_2, O_2, Cl_2	e.g. Na, Fe, Cu

Table 4.1.2 — Types of compounds

Ionic compounds	Covalent compounds
e.g. NaCl exists as cations and anions, formed when a metal atom loses electrons to a non-metal atom to achieve a stable noble gas configuration	e.g. HCl exists as diatomic molecules formed when two atoms share electrons to achieve a stable noble gas configuration

DEFINITION 4.1.1 — MIXTURE

A mixture is an impure substance containing two or more substances (elements and/or compounds) physically mixed in any ratio.

Table 4.1.3 — Compounds vs mixtures

	Compounds	Mixtures
Made up of	Two or more elements chemically combined	Two or more elements and/or compounds physically combined
Ratio of constituents	Fixed ratio	No fixed ratio
Properties	Different from constituent elements	Similar to constituent elements
Melting/boiling point	Fixed	Range of temperatures

Separation of compounds requires a chemical method: thermal decomposition (strong heating) or electrolysis (passing an electric current through the compound) — unlike mixtures, which can be separated by physical methods (Chapter 1).

4.2 | Formation of Ionic Bonds

DEFINITIONS 4.2.1

Cation — a positively charged particle formed when an atom loses electrons.
Anion — a negatively charged particle formed when an atom gains electrons.
Ionic compound — a pure substance made up of cations and anions held together by strong ionic bonds.
Ionic bonding — the strong electrostatic force of attraction between cations and anions.

When a metal reacts with a non-metal, electrons transfer from metal atoms to non-metal atoms. The metal atom loses electrons to form a cation; the non-metal atom gains electrons to form an anion. The resulting compound is held together by strong electrostatic attraction between oppositely charged ions.

Table 4.2.1 — Giant ionic lattice structure

Type of matter	Ionic compounds
Type of structure	Giant ionic lattice structure
Particles	Cations and anions
Forces between particles	Strong electrostatic forces of attraction (ionic bonds)
Arrangement/movement	Ions orderly arranged in fixed positions, closely packed in a fixed ratio; ions vibrate about fixed positions only

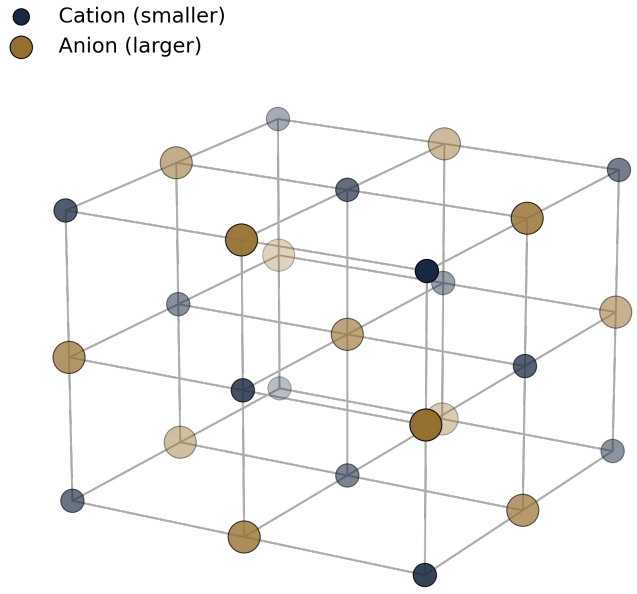


Fig 4.2.1. *Giant ionic lattice (rock-salt structure): cations and anions alternate in a fixed, orderly 3-D arrangement, held together by strong electrostatic forces in every direction.*

Properties of Ionic Compounds

PROPERTY 1 — HIGH MELTING AND BOILING POINTS

Explain, in terms of structure, why calcium oxide is a solid at r.t.p.

Calcium oxide is an ionic compound with a giant ionic lattice structure. In this structure, the calcium ions and oxide ions are held together by strong electrostatic forces of attraction (ionic bonds). A lot of energy is required to overcome these strong forces to melt CaO. Thus, CaO has a high melting point and is a solid at r.t.p.

EXPLAIN, IN TERMS OF STRUCTURE, WHY CALCIUM OXIDE HAS A HIGHER MELTING POINT THAN SODIUM CHLORIDE.

The charge of the calcium ion is 2+ and the charge of the oxide ion is 2−, whereas the charge of the sodium ion is 1+ and the charge of the chloride ion is 1−. Thus, the electrostatic force of attraction between Ca^{2+} and O^{2-} ions is stronger than the force between Na^{+} and Cl^{-} ions. More energy is therefore required to overcome the stronger forces to melt CaO than NaCl. Hence, CaO has a higher melting point than NaCl.

PROPERTY 2 — HARD, WITH REGULAR SHAPES

Explain, in terms of structure, why magnesium oxide crystals are hard and have regular shapes.

Magnesium oxide is an ionic compound with a giant ionic lattice structure. To separate the Mg^{2+} ions and O^{2-} ions from each other, a lot of force must be applied to overcome the strong electrostatic forces of attraction which hold these ions in fixed positions. Hence, MgO is hard. The strong electrostatic forces of attraction hold the ions closely packed together and orderly arranged in fixed positions, so the crystal of MgO also has a regular shape.

PROPERTY 3 — CONDUCT ELECTRICITY ONLY WHEN MOLTEN OR AQUEOUS

Explain, in terms of structure, why sodium bromide crystals do not conduct electricity, but aqueous sodium bromide is a good conductor of electricity.

Sodium bromide crystal is an ionic compound with a giant ionic lattice structure. In solid sodium bromide, the sodium and bromide ions are **not mobile** as they are held in fixed positions by strong electrostatic forces of attraction, so they cannot act as charge carriers. In aqueous sodium bromide, the sodium and bromide ions become **mobile** and are free to move throughout the solution. Hence, these ions can act as charge carriers and conduct electricity.

PROPERTY 4 — USUALLY SOLUBLE IN WATER (WITH EXCEPTIONS)

Explain, in terms of structure, why sodium nitrate is soluble in water.

Sodium nitrate is an ionic compound with a giant ionic lattice structure. The force of attraction between the Na^+/NO_3^- ions and water molecules is stronger than the electrostatic forces of attraction (ionic bonds) holding the ions in the lattice. Thus, the ions are pulled away from the lattice by water molecules, and sodium nitrate dissolves.

Explain, in terms of structure, why magnesium oxide is insoluble in water.

Magnesium oxide is an ionic compound with a giant ionic lattice structure. The electrostatic forces of attraction (ionic bonds) between the Mg^{2+} and O^{2-} ions in the lattice are stronger than the force of attraction between these ions and water molecules. Thus, the ions remain held in the lattice, and magnesium oxide does not dissolve in water.

4.3 | Covalent Compounds and Molecules of Elements

DEFINITION 4.3.1 — MOLECULE

A molecule is an electrically neutral particle made of two or more atoms chemically combined by sharing electrons.

DEFINITIONS 4.3.2

Molecule of an element — atoms of the same kind chemically combined by electron sharing.
Molecule of a compound — atoms of different elements chemically combined by electron sharing.

Table 4.3.1 — Examples of molecules

Molecules of elements	Molecules of compounds
H_2, N_2, F_2, O_2, I_2, Cl_2, Br_2, At_2, O_3 (ozone)	CH_4, NH_3, HCl, HNO_3, H_2SO_4, CH_3COOH, C_2H_5OH, H_2O_2

DEFINITION 4.3.3 — COVALENT BOND

The strong electrostatic force of attraction between the positively charged nuclei of two atoms and the electron pair(s) shared between them.

Table 4.3.2 — Types of covalent bonds

Bond type	Examples
Single bond	F–F (fluorine); H–Cl (hydrogen chloride)
Double bond	O=O (oxygen); O=C=O (carbon dioxide)
Triple bond	N ≡ N (nitrogen); H–C ≡ C–H (ethyne/acetylene)

DEFINITION 4.3.4 — VALENCY

The number of electron(s) an atom must lose, gain or share to attain the stable noble gas electronic configuration.

> ⓘ **NOTE — NAMING BINARY COVALENT COMPOUNDS**
>
> 1. Identify the element furthest to the left/bottom of the Periodic Table (named first, unchanged).
> 2. Add an "–ide" ending to the second element's name.
> 3. Use a prefix (mono-, di-, tri-, tetra- ...) to indicate the number of atoms of each element, e.g. carbon dioxide (CO_2), sulfur trioxide (SO_3).

Polyatomic Ions and Dative (Coordinate) Bonding

A polyatomic ion contains more than one atom held together by covalent bonds; its overall charge arises from the ion as a whole gaining/losing electrons (e.g. NH_4^+, OH^-, SO_4^{2-}). A **dative (coordinate) bond** is a covalent bond in which both shared electrons come from the same atom (e.g. the fourth N–H bond in NH_4^+, formed when NH_3 donates its lone pair to H^+). Once formed, a dative bond behaves identically to a normal covalent bond.

CHAPTER 4 QUICK RECAP: Metal + non-metal → electron transfer → ionic bond (strong electrostatic attraction between cations/anions) → giant ionic lattice → high m.p./b.p., hard, conducts only molten/aqueous, often water-soluble. Non-metal + non-metal → electron sharing → covalent bond, forming molecules (simple, giant, or macromolecular structures — Chapter 5).

Chapter 5: Structure and Properties of Materials

Covalent structures (simple, giant, macromolecular) and metallic bonding

5.1 | Structures of Covalent Substances

Covalent substances can exist in three structural forms: **simple covalent molecules**, **giant covalent (molecular) structures**, and **macromolecules (polymers)**.

Table 5.1.0 — Comparing the three covalent structural forms

	Simple Molecular	Giant Covalent	Macromolecular
Particles	Individual molecules	Atoms (one giant network)	Long-chain molecules
Bonding within particle	Strong covalent bonds	Strong covalent bonds throughout	Strong covalent bonds along the chain
Forces between particles	Weak intermolecular forces	N/A — one continuous network	Weak intermolecular forces between chains
Melting/boiling point	Low	Very high	No fixed point; softens over a range
Electrical conductivity	Poor (exceptions apply)	Poor, except graphite	Poor
Examples	H_2O, CO_2, CH_4	Diamond, graphite, SiO_2	Poly(ethene), starch, proteins

Simple Covalent Molecules

Type of matter	Covalent substance
Type of structure	Simple covalent molecules
Particles	Molecules
Forces between particles	Weak intermolecular forces of attraction between molecules (the strong covalent bonds are *within* each molecule, not between molecules)

> ⓘ **NOTE**
>
> Empirical formula gives the simplest whole-number ratio of atoms in a covalent substance, derived from its molecular formula (e.g. molecular formula C_2H_4 → empirical formula CH_2).

> **PROPERTY 1 — LOW MELTING AND BOILING POINTS**
>
> **Explain, in terms of structure, why methane is a gas at r.t.p.**
> Methane is a covalent substance that exists as simple covalent molecules. In this structure, there are only weak intermolecular forces of attraction between methane molecules. Only a small amount of thermal energy is required to overcome these weak forces to melt and boil CH_4. Hence, CH_4 has low melting and boiling points, and is a gas at r.t.p.
>
> **Explain, in terms of structure, why ethanol is a volatile liquid at r.t.p.**
> Ethanol is a covalent compound that exists as simple covalent molecules. In this structure, there are only weak intermolecular forces of attraction between ethanol molecules. Only a small amount of thermal energy is required to overcome these weak forces to melt and boil ethanol. As ethanol has low melting and boiling points, it is a volatile liquid, evaporating quickly at r.t.p.

PROPERTY 2 — POOR CONDUCTORS OF ELECTRICITY (WITH EXCEPTIONS)

Explain, in terms of structure, why pure water is a poor conductor of electricity.

Pure water exists as simple covalent water molecules that are electrically neutral. Pure water does not have freely moving ions or delocalised electrons that can act as charge carriers to conduct electricity.

Explain, in terms of structure, why hydrogen chloride gas cannot conduct electricity, but once the gas is dissolved in water, the aqueous solution can conduct electricity.

Hydrogen chloride gas exists as simple covalent HCl molecules that are electrically neutral, and does not have freely moving ions or delocalised electrons to act as charge carriers. When hydrogen chloride gas dissolves in water, an aqueous solution of hydrochloric acid is formed; in the aqueous state, HCl molecules ionise in water to form H^+ and Cl^- ions which are mobile. These ions can act as charge carriers and conduct electricity.

Property 3 — Generally insoluble in water (exceptions: sugar, ethanol, HCl gas), but soluble in organic solvents (oil, benzene, esters, tetrachloromethane, dichloromethane, hexane).

Giant Covalent (Molecular) Structures

Type of matter	Covalent substance
Type of structure	Giant covalent structure
Particles	Atoms
Forces between particles	Strong covalent bonds between atoms throughout the structure
Arrangement	Atoms held in fixed positions throughout the giant, rigid structure

Table 5.1.1 — Substances with giant covalent structures

Substance	Structure / bonding	Everyday use
Diamond (an allotrope of carbon)	Each carbon atom bonded to 4 other carbon atoms in a rigid tetrahedral network	Drill tips, cutting tools
Graphite (an allotrope of carbon)	Planar hexagonal layers; each carbon bonded to 3 others within a layer; layers held by weak intermolecular forces; 1 delocalised electron per carbon atom	Dry lubricant, pencil "lead", electrodes
Silicon dioxide, SiO_2	Each Si bonded to 4 O atoms; each O bonded to 2 Si atoms (ratio Si:O = 1:2)	Glassware, sand

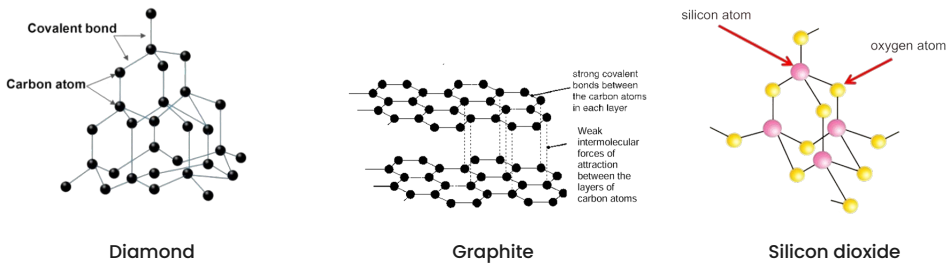


Fig 5.1.2. *Giant covalent structures: diamond's rigid tetrahedral network, graphite's hexagonal layers held by weak forces between layers, and silicon dioxide's Si–O tetrahedral network.*

DEFINITION 5.1.1 — ALLOTROPES

Allotropes are different structural forms of the same element in the same physical state (e.g. diamond and graphite are both allotropes of carbon).

PROPERTY 1 — HIGH MELTING AND BOILING POINTS

Explain, in terms of structure, why sand is non-volatile.
Sand is mainly made up of silicon dioxide, which is a covalent substance with a giant covalent structure. Each silicon atom is bonded to four oxygen atoms by strong covalent bonds throughout the giant rigid tetrahedral structure; each oxygen atom is bonded to two silicon atoms by strong covalent bonds throughout the structure. A large amount of thermal energy is required to overcome these strong covalent bonds to melt and boil silicon dioxide. Thus, sand has a high melting and boiling point range, and is non-volatile.

Explain, in terms of structure, why graphite has high melting and boiling points.
Graphite is a covalent substance with a giant covalent structure, made up of a giant network of planar hexagonal ring layers of carbon atoms. Within each layer, each carbon atom is bonded to three other carbon atoms by strong covalent bonds. A large amount of thermal energy is required to break these strong covalent bonds to melt and boil graphite. Thus, graphite has high melting and boiling points.

PROPERTY 2 — SOME ARE HARD, OTHERS ARE SOFT

Explain, in terms of structure, why silicon dioxide is hard.
Silicon dioxide is a covalent substance with a giant covalent structure. Each silicon atom is bonded to four oxygen atoms by strong covalent bonds throughout the giant rigid tetrahedral structure; each oxygen atom is bonded to two silicon atoms by strong covalent bonds throughout the structure. A large amount of force must be applied to overcome the strong covalent bonds to displace the atoms from their fixed positions in the structure. Thus, silicon dioxide is hard.

Explain, in terms of structure, why graphite is used to make dry lubricant.
Graphite is a covalent substance with a giant covalent structure, made up of a giant network of planar hexagonal ring layers of carbon atoms. The layers of carbon atoms are held together by weak intermolecular forces of attraction. Only a small amount of force needs to be applied to overcome these weak forces for the layers of atoms to slide past each other. Thus, graphite is soft and slippery, and can be used as a dry lubricant for machinery.

PROPERTY 3 — MOSTLY POOR CONDUCTORS, EXCEPT GRAPHITE

Explain, in terms of structure, why silicon dioxide is a poor conductor of electricity.
Silicon dioxide is a covalent substance with a giant covalent structure. In this structure, each silicon atom is bonded to four oxygen atoms by strong covalent bonds throughout the giant rigid tetrahedral structure, and each oxygen atom is bonded to two silicon atoms by strong covalent bonds throughout the structure. All the valence electrons of each silicon atom and oxygen atom are involved in covalent bonding. There are no delocalised electrons that can act as charge carriers and move freely throughout the structure. Thus, silicon dioxide is a poor conductor of electricity.

Explain, in terms of structure, why graphite is a good conductor of electricity.
Graphite is a covalent substance with a giant covalent structure, made up of a giant network of planar hexagonal ring layers of carbon atoms; within each layer, each carbon atom is bonded to three other carbon atoms by strong covalent bonds. Each carbon atom has one valence electron that is not used to form covalent bonds. These valence electrons are delocalised and can move freely throughout the layers of carbon atoms in the structure, acting as charge carriers. Thus, graphite is a good conductor of electricity.

Property 4 — Giant covalent structures are usually insoluble in both water and organic solvents, since the attraction between the structure's atoms and the solvent cannot overcome the strong covalent bonds within the giant structure.

Macromolecules (Polymers)

A macromolecule (polymer) is a long chain of covalently-bonded repeating units (monomers).

Type of matter	Covalent substance
Type of structure	Macromolecules
Particles	Long-chain molecules
Forces between particles	Weak intermolecular forces between long-chain molecules

Natural polymers: silk, wool, rubber, starch, proteins, lipids. **Man-made (synthetic) polymers:** polyester, poly(ethene), nylon,

polystyrene, and other plastics.

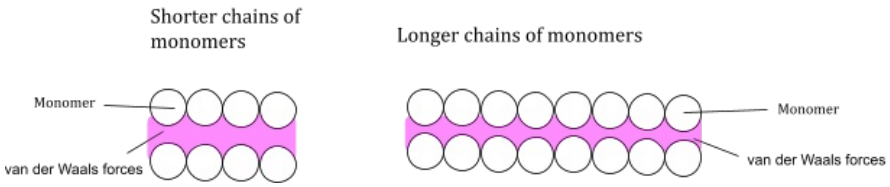


Fig 5.1.3. *Macromolecules are long chains of covalently-bonded repeating units, with weaker forces acting between neighbouring chains.*

> **PROPERTY 1 — NO FIXED MELTING/BOILING POINT (RELATIVELY LOW RANGE)**
>
> Chain length (number of monomers joined) is not fixed and may vary in size, so intermolecular forces between chains vary in strength; polymers soften gradually over a range of temperatures (the **softening temperature**) rather than melting sharply. Longer chains generally have stronger intermolecular forces (more surface area of contact) and hence higher softening temperatures.
>
> Although the intermolecular forces are individually weak, most polymers are still solid at r.t.p. because the very large chain length means the cumulative intermolecular forces between chains are strong enough to keep the m.p./b.p. range above 25°C.

Property 2 — Poor electrical conductors (in solid, liquid or gaseous state), as polymer molecules are electrically neutral with no mobile ions or delocalised electrons.

Property 3 — Most polymers are insoluble in water but soluble in organic solvents (oil, benzene, esters, tetrachloromethane, dichloromethane, hexane).

5.2 | Metallic Bonding

Table 5.2.1 — Physical properties of metals

Physical properties	
Shiny (lustrous)	Mostly solid at r.t.p. (except mercury, liquid)
Usually high m.p./b.p.	Malleable and ductile
Good thermal conductor	Good electrical conductor
Usually high density (except Li, Na, K)	

> **DEFINITION 5.2.1 — METALLIC BONDING**
>
> The strong electrostatic force of attraction between positively charged metal ions and a sea of delocalised, mobile, negatively charged valence electrons.

Type of matter	Metals
Type of structure	Giant metallic lattice structure
Particles	Positively charged metal ions + a "sea" of delocalised electrons
Forces between particles	Strong metallic bonds

In the lattice, metal atoms lose their valence electrons to become positive ions; these electrons no longer belong to any single atom and move freely between the ions, like a "sea of delocalised mobile electrons."

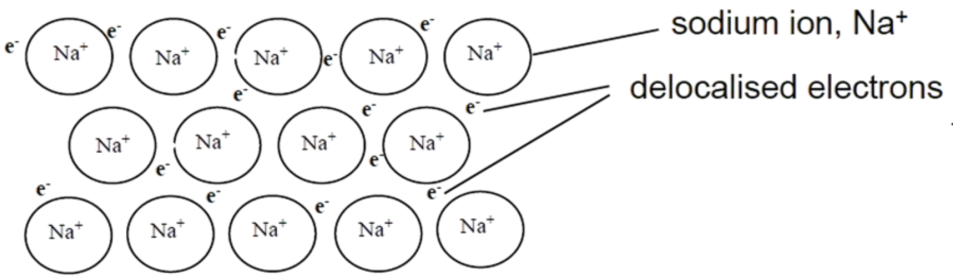


Fig 5.2.1. *Giant metallic lattice: positively charged metal ions held in fixed, orderly positions by a surrounding sea of delocalised valence electrons.*

ⓘ NOTE

Since metals are solid at r.t.p., the metal ions must be orderly arranged in the lattice, with spaces occupied by the sea of delocalised electrons.

PROPERTY 1 — HIGH MELTING AND BOILING POINTS

Explain, in terms of structure, why aluminium is a solid at r.t.p.

Aluminium is a metal with a giant metallic lattice structure. In this structure, there are strong electrostatic forces of attraction between positively charged aluminium ions and the negatively charged sea of delocalised, mobile valence electrons. A large amount of thermal energy is required to overcome these strong forces to melt Al. Hence, Al has a high melting point and is a solid at r.t.p.

Use the structure of metals to explain why magnesium has a higher melting point (650°C) than sodium (97.79°C).

Sodium atoms have an electronic configuration of 2,8,1 and magnesium atoms have an electronic configuration of 2,8,2. This means each magnesium atom loses 2 valence electrons to form a more highly charged metal ion, Mg^{2+}, compared to a sodium atom, which loses only 1 valence electron to form a singly-charged metal ion, Na^{+}. Hence, the electrostatic force of attraction between metal ions and delocalised valence electrons is stronger in magnesium than in sodium. A large amount of thermal energy is needed to overcome the stronger electrostatic forces of attraction. Therefore, magnesium has a higher melting point than sodium.

Which metal has a stronger metallic bond — sodium (m.p. 97.79°C) or potassium (m.p. 63.5°C)? Why?

Sodium ions have an electronic configuration of 2,8 and potassium ions have an electronic configuration of 2,8,8. Sodium ions have one fewer electron shell than potassium ions, so sodium has a smaller ionic radius. In the giant metallic lattice structure, the delocalised valence electrons are closer to the positively charged nuclei of the sodium ions than the potassium ions. Hence, the electrostatic force of attraction between the nucleus of the positive ions and the valence electrons is stronger in sodium than in potassium. More thermal energy is therefore needed to overcome the stronger electrostatic forces of attraction in sodium, giving it a higher melting point.

ⓘ NOTE — EXCEPTIONS

Group 1 (alkali) metals have relatively low m.p./b.p. despite being solid at r.t.p.; mercury is an exception that is liquid at r.t.p.

PROPERTY 2 — HIGH DENSITY (EXCEPTIONS: LI, NA, K, WHICH FLOAT ON WATER)

Explain, in terms of structure, why iron has a high density while oxygen has a low density at r.t.p.

Iron is a solid while oxygen is a gas at room temperature. In a given volume, iron atoms are very closely packed together, while oxygen molecules are spaced far apart from each other. The relative atomic mass of iron (56) is also much greater than the relative molecular mass of oxygen (32). Thus, in a fixed volume, there is a much larger number of iron atoms, each with a higher mass, compared to a much smaller number of oxygen molecules with a lower mass. Hence, iron has a higher density than oxygen at room temperature.

PROPERTY 3 — GOOD ELECTRICAL CONDUCTIVITY

Explain, in terms of structure, why sodium is a good conductor of electricity.

Sodium has delocalised valence electrons which act as mobile charge carriers, and can move freely throughout the giant metallic lattice structure. When connected to an electrical circuit, these valence electrons are able to move from the negative terminal to the positive terminal, conducting electricity.

PROPERTY 4 — GOOD THERMAL CONDUCTIVITY

Explain, in terms of structure, why iron is a good thermal conductor.

Iron has delocalised valence electrons which can move freely throughout the giant metallic lattice structure. When heated, these electrons absorb thermal energy, gain kinetic energy, and move rapidly from the hotter end to the colder end, carrying the thermal energy with them. Hence, thermal energy is transferred easily by these delocalised electrons throughout the structure.

DEFINITIONS 5.2.2

Malleability — the ability of a material to be hammered into different shapes without breaking.
Ductility — the ability of a material to be drawn into wires.

PROPERTY 5 — MALLEABLE AND DUCTILE

Explain, in terms of structure, why aluminium is malleable and ductile.

In the giant metallic lattice structure of aluminium, aluminium atoms, which have the same atomic size, are very closely packed and orderly arranged in layers. When a force is applied, the layers of aluminium atoms can slide over one another easily without breaking the metallic bonds. Thus, aluminium is soft, and hence malleable and ductile.

Alloys

DEFINITION 5.2.3 — ALLOY

A mixture of a metal with one or more other elements.

Table 5.2.2 — Common alloys

Alloy	Composition	Properties	Uses
Brass	Cu 70%, Zn 30%	Harder/stronger than pure copper; corrosion-resistant; attractive gold-like colour	Ornaments, coins, musical instruments
Bronze	Cu 88%, Sn 12%	Harder/stronger; corrosion-resistant; red-brown colour	Ornaments, bells, medals
Mild steel	Fe 99.5%, C 0.5%	Harder/stronger; durable	Car bodies
Stainless steel	Fe 73%, Cr 18%, Ni 8%, C 1%	Harder/stronger; highly resistant to corrosion	Cooking utensils, surgical instruments

WHY ARE ALLOYS HARDER/STRONGER THAN PURE METALS?

In an alloy (e.g. bronze), atoms of different elements (Cu, Sn) have different atomic sizes, which disrupts the orderly layered arrangement of the pure metal. This prevents layers of atoms from sliding over each other easily when a force is applied — hence alloys are harder, stronger, but less malleable/ductile than the pure metal.

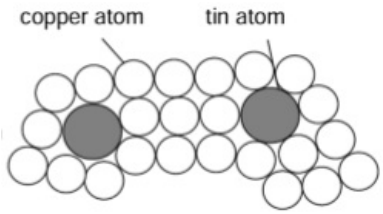


Fig 5.2.2. *Structure of bronze (an alloy of copper and tin) — tin atoms of a different size disrupt the orderly layered arrangement of copper atoms, making the alloy harder than pure copper.*

Since alloys are impure mixtures, they melt/boil over a **range of temperatures** (no fixed m.p./b.p.). As alloys still contain metal(s) with delocalised mobile electrons, they remain good electrical and thermal conductors.

CHAPTER 5 QUICK RECAP: Simple molecular (weak intermolecular forces → low m.p./b.p., poor conductors) vs giant covalent (strong covalent bonds throughout → high m.p./b.p., hard, usually poor conductors except graphite) vs macromolecular (long chains, weak intermolecular forces between chains → soften over a range, poor conductors). Metals: sea of delocalised electrons + metal cations → high m.p./b.p. (except Group 1, Hg), malleable/ductile, good conductors; alloys are harder/stronger due to disrupted atomic layering.

Chapter 6: Chemical Formulae, Equations and Stoichiometry

Writing formulae, balancing equations, and ionic equations

6.1 | Classifying Elements by Formula Type

DEFINITION 6.1.1 — MONATOMIC ELEMENTS

Pure substances that exist as single, uncombined atoms (e.g. noble gases He, Ne, Ar; most metals Na, Fe, Cu).

DEFINITION 6.1.2 — DIATOMIC MOLECULES

Particles made of two atoms covalently bonded together (e.g. H_2, N_2, O_2, Cl_2, HCl).

DEFINITION 6.1.3 — POLYATOMIC MOLECULES

Particles made of three or more atoms covalently bonded together (e.g. H_2O, CO_2, NH_3).

> **ⓘ NOTE — SULFUR EXCEPTION**
>
> Sulfur exists as S_8 molecules, but by convention we write "S" (not S_8) when writing chemical equations involving sulfur.

Elements with giant covalent structures (mostly Group 14 elements, e.g. carbon as diamond/graphite, silicon) are written simply using their chemical symbol (C, Si), since there is no discrete "molecule".

6.2 | Writing and Balancing Chemical Equations

To write a chemical formula for an ionic compound: identify the charges of the cation and anion, then combine in whatever ratio makes the overall charge zero (the "criss-cross" method). E.g. Al^{3+} and SO_4^{2-} combine as $Al_2(SO_4)_3$ (charges 3+ × 2 = 6+ balances 2− × 3 = 6−).

A balanced chemical equation has equal numbers of atoms of each element on both sides, and includes **state symbols**: (s) solid, (l) liquid, (g) gas, (aq) aqueous (dissolved in water).

6.3 | Writing Ionic Equations

An ionic equation shows only the species that actually take part in a reaction, omitting "spectator ions" that do not change.

WORKED EXAMPLE — WRITING AN IONIC EQUATION

Step 1: Write the balanced full chemical equation:

$$4HCl(aq) + MnO_2(s) \rightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$$

Step 2: Split/ionise all aqueous ionic compounds and acids into their constituent ions:

$$4H^+(aq) + 4Cl^-(aq) + MnO_2(s) \rightarrow Mn^{2+}(aq) + 2Cl^-(aq) + 2H_2O(l) + Cl_2(g)$$

Step 3: Cancel "spectator ions" — ions appearing unchanged (same physical state and charge) on both sides. Here, 2 of the 4 Cl^- ions cancel:

$$4H^+(aq) + 2Cl^-(aq) + MnO_2(s) \rightarrow Mn^{2+}(aq) + 2H_2O(l) + Cl_2(g)$$

ⓘ NOTE

Only species in aqueous solution are split into ions. Solids, liquids and gases (like $MnO_2(s)$, $H_2O(l)$, $Cl_2(g)$ above) are left as full formulae, since they are not dissociated into free ions.

CHAPTER 6 QUICK RECAP: Ionic compound formulae balance total charge to zero; equations must be balanced with correct state symbols. To write an ionic equation: write the full balanced equation → split aqueous ionic compounds/acids into ions → cancel spectator ions (unchanged species, same state, same charge, on both sides).

Chapter 7: The Mole Concept

Relative masses, moles, molar volume, concentration, and stoichiometric calculations

7.1 | Relative Masses and the Mole

DEFINITIONS 7.1.1 — RELATIVE MASSES

Relative atomic mass (A_r) — the average mass of one atom of an element relative to 1/12 the mass of one carbon-12 atom.
Relative molecular mass (M_r) — the average mass of one molecule of a covalent substance relative to 1/12 the mass of one carbon-12 atom.
Relative formula mass (M_r) — the average mass of one formula unit of an ionic compound relative to 1/12 the mass of one carbon-12 atom.

$$\text{Mass of element in compound} = (A_r \times \text{number of atoms of that element} \div M_r \text{ of compound}) \times \text{mass of compound}$$

DEFINITION 7.1.2 — THE MOLE

One mole of any substance always contains 6.02×10^{23} particles of that substance (the Avogadro constant, L or N_A).

$$\text{Number of moles} = \text{Number of particles} \div 6.02 \times 10^{23}$$

KEY QNS 7.1.1

Q. How many hydrogen atoms are there in 3 moles of water?
A. Each H_2O molecule has 2 H atoms, so 3 mol H_2O contains $3 \times 2 = 6$ mol H atoms = $6 \times 6.02 \times 10^{23} = 3.61 \times 10^{24}$ H atoms.

DEFINITION 7.1.3 — MOLAR MASS

The mass of one mole of a substance. Molar mass has the same numerical value as A_r/M_r, but carries units of g/mol.

$$\text{Molar mass} = \text{Mass of substance} \div \text{Number of moles}$$

KEY QNS 7.1.2 — AMMONIUM CHLORIDE

Calculate the number of moles in 10 kg of ammonium chloride, NH_4Cl ($M_r = 14+4+35.5 = 53.5$).
Moles = 10,000 g ÷ 53.5 g/mol = **187 mol** (3 s.f.)

Gases: Avogadro's Law and Molar Volume

DEFINITION 7.1.4 — AVOGADRO'S LAW

At the same temperature and pressure, equal volumes of gases contain the same number of particles (moles).

DEFINITION 7.1.5 — MOLAR VOLUME

1 mole of any gas occupies 24 dm^3 (24,000 cm^3) at room temperature and pressure (r.t.p.).

$$\text{Number of moles of gas} = \text{Volume } (dm^3) \div 24$$

KEY QNS 7.1.3 — CARBON DIOXIDE

Calculate the mass of 1500 cm^3 of CO_2 at r.t.p.

Moles of CO_2 = 1.5 dm^3 ÷ 24 = 0.0625 mol. Mass = 0.0625 × 44 = **2.75 g**

Concentration

DEFINITION 7.1.6 — CONCENTRATION

The amount of solute dissolved per unit volume of solution.

$$\text{Molar concentration } (\text{mol/dm}^3) = \text{Moles of solute} \div \text{Volume of solution } (\text{dm}^3)$$

$$\text{Concentration } (\text{g/dm}^3) = \text{Mass of solute} \div \text{Volume of solution } (\text{dm}^3)$$

$$\text{Concentration } (\text{mol/dm}^3) = \text{Concentration } (\text{g/dm}^3) \div \text{Molar mass of solute}$$

KEY QNS 7.1.4 — BARIUM CHLORIDE

Find the mass of $BaCl_2$ in 500 cm^3 of 0.250 mol/dm^3 solution.

Moles = 0.500 × 0.250 = 0.125 mol. Mass = 0.125 × (137+35.5×2) = **26.0 g**

7.2 | Empirical Formula, Molecular Formula & Stoichiometry

DEFINITIONS 7.2.1

Molecular formula — shows the actual types and numbers of atoms in one unit of a covalent substance.

Empirical formula — shows the simplest whole-number ratio of atoms of each element in a compound.

KEY QNS 7.2.1 — EMPIRICAL FORMULA FROM MASS

3.53 g of Fe reacts with chlorine to form 10.24 g of iron chloride. Find the empirical formula.

	Fe	Cl
Mass / g	3.53	10.24−3.53 = 6.71
Molar mass / g mol^{-1}	56	35.5
Moles / mol	0.0630	0.189
Ratio	1	3

Empirical formula: **$FeCl_3$**

KEY QNS 7.2.2 — EMPIRICAL FORMULA FROM PERCENTAGE COMPOSITION

An oxide of titanium contains 60% Ti by mass. Find its empirical formula.

	Ti	O
Mass / g (per 100 g)	60	40
Molar mass / g mol^{-1}	48	16
Moles / mol	1.25	2.5
Ratio	1	2

Empirical formula: **TiO_2**

Molecular formula = $(\text{Empirical formula})_n$, where n = Molecular mass ÷ Molar mass of empirical formula

KEY QNS 7.2.3 — MOLECULAR FORMULA FROM PERCENTAGE + M_R

A compound is 87.5% N and 12.5% H by mass, with M_r = 32. Find its molecular formula.
Moles N = 87.5/14 = 6.25; moles H = 12.5/1 = 12.5; ratio N:H = 1:2 ⇒ empirical formula NH_2 (mass 16).
n = 32/16 = 2 ⇒ molecular formula = **N_2H_4** (hydrazine).

KEY QNS 7.2.4 — WATER OF CRYSTALLISATION

A 124.8 g sample of hydrated copper(II) sulfate contains 31.8 g Cu^{2+} and 48.0 g SO_4^{2-}.
Mass of water = 124.8 − 31.8 − 48.0 = 45.0 g ⇒ moles H_2O = 45.0/18 = 2.5 mol.
Moles of $CuSO_4$ = (31.8+48.0)/(63.5+32+64) = 0.500 mol.
Ratio $CuSO_4$: H_2O = 1 : 5 ⇒ formula is **$CuSO_4.5H_2O$**.

Stoichiometry — Reacting Masses, Volumes and Concentrations

DEFINITION 7.2.2 — STOICHIOMETRY

The quantitative relationship between the number of moles of reactants and products in a balanced chemical equation.

KEY QNS 7.2.5 — REACTING MASS

$2H_2S(g) + 3O_2(g) \rightarrow 2H_2O(g) + 2SO_2(g)$. If 0.68 g H_2S burns completely, find the mass of SO_2 formed.
Moles H_2S = 0.68/34 = 0.02 mol. Ratio H_2S : SO_2 = 1:1 ⇒ moles SO_2 = 0.02 mol.
Mass SO_2 = 0.02 × 64 = **1.28 g**

KEY QNS 7.2.6 — FINDING WATER OF CRYSTALLISATION, X

$MgCl_2.xH_2O(s) \rightarrow MgCl_2(s) + xH_2O(g)$. Heating 20.3 g of hydrate gives 9.5 g anhydrous $MgCl_2$.
Mass H_2O = 20.3−9.5 = 10.8 g ⇒ moles H_2O = 10.8/18 = 0.6 mol. Moles $MgCl_2$ = 9.5/95 = 0.1 mol.
Ratio $MgCl_2$: H_2O = 1 : 6 ⇒ **x = 6**

KEY QNS 7.2.7 — GASEOUS VOLUME RATIOS (AVOGADRO'S LAW)

$CH_4(g) + H_2O(g) \rightarrow 3H_2(g) + CO(g)$. Find the volume of methane reacted when 100 cm^3 of H_2 is produced.
Mole ratio (= volume ratio, by Avogadro's Law) H_2:CH_4 = 3:1 ⇒ volume of CH_4 = (1/3) × 100 = **33.3 cm^3**

KEY QNS 7.2.8 — DEDUCING HYDROCARBON FORMULA FROM COMBUSTION DATA

30 cm^3 of C_nH_{2n} burns in excess O_2 to give 120 cm^3 CO_2. Find n.
Moles CO_2 = 0.120/24 = 0.005 mol = moles of C atoms. Moles hydrocarbon = 0.030/24 = 0.00125 mol.
Moles of C atoms in hydrocarbon = 0.00125n = 0.005 ⇒ n = 4 ⇒ **C_4H_8**

KEY QNS 7.2.9 — TITRATION-STYLE CONCENTRATION CALCULATION

2.65 g Na_2CO_3 dissolved and made up to 250 cm^3. 25.0 cm^3 of this solution requires 20.0 cm^3 of HNO_3 for complete neutralisation: $Na_2CO_3(aq) + 2HNO_3(aq) \rightarrow 2NaNO_3(aq) + CO_2(g) + H_2O(l)$. Find the concentration of HNO_3 in g/dm^3.
Moles Na_2CO_3 in 25 cm^3 = (2.65/106) × (25/250) = 0.0025 mol. Ratio Na_2CO_3:HNO_3 = 1:2 ⇒ moles HNO_3 = 0.005 mol.
Conc. = (0.005/0.020) × 63 = **15.8 g/dm^3**

7.3 | Limiting Reactants, Percentage Yield and Percentage Purity

DEFINITION 7.3.1 — LIMITING REACTANT

The reactant that is completely used up in a reaction, limiting the amount of product formed. Any other reactant remaining unreacted is "in excess."

KEY QNS 7.3.1 — IDENTIFYING THE LIMITING REACTANT (MASS)

Q. 6 g H_2 reacts with 32 g O_2: $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$. Find the mass of water formed.

Moles H_2 = 6 mol
Moles O_2 = 2 mol
Required ratio $H_2 : O_2$ = 2 : 1, so 2 mol O_2 needs 4 mol H_2
Since 6 mol H_2 is available, **H_2 is in excess and O_2 is limiting**

Ratio $O_2 : H_2O$ = 1 : 1, so moles H_2O = 2 mol
Mass = 2 × 18 = **36 g**

KEY QNS 7.3.2 — LIMITING REACTANT WITH GAS VOLUME

Q. 4.8 dm^3 H_2 (r.t.p.) is passed over 67.0 g PbO: $H_2(g) + PbO(s) \rightarrow Pb(s) + H_2O(g)$.

Moles H_2 = 4.8 ÷ 24 = 0.2 mol
Moles PbO = 67 ÷ 223 = 0.3 mol
Ratio is 1 : 1, so **H_2 is limiting**
PbO remaining = 0.3 − 0.2 = 0.1 mol = 22.3 g
Mass of Pb formed = 0.2 × 207 = **41.4 g**

KEY QNS 7.3.3 — LIMITING REACTANT WITH GAS VOLUMES ONLY

Q. 100 cm^3 NH_3 + 120 cm^3 O_2: $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$.

Volume ratio $NH_3 : O_2$ = 4 : 5
So 120 cm^3 O_2 needs (4/5) × 120 = 96 cm^3 NH_3
Since 100 cm^3 NH_3 is provided, **NH_3 is in excess (4 cm^3 remaining) and O_2 is limiting**

Volume of NO formed = (4/5) × 120 = **96 cm^3**

KEY QNS 7.3.4 — LIMITING REACTANT WITH CONCENTRATIONS

Q. 25.0 cm^3 of 3.00 mol/dm^3 $AlCl_3$ + 75.0 cm^3 of 2.00 mol/dm^3 $AgNO_3$: $AlCl_3(aq) + 3AgNO_3(aq) \rightarrow Al(NO_3)_3(aq) + 3AgCl(s)$.

Moles $AlCl_3$ = 0.075 mol
Moles $AgNO_3$ = 0.150 mol
Required ratio 1 : 3, so 0.150 mol $AgNO_3$ needs only 0.050 mol $AlCl_3$
$AlCl_3$ is in excess, $AgNO_3$ is limiting

Moles AgCl = 0.150 mol
Mass = 0.150 × 143.5 = **21.5 g**

DEFINITIONS 7.3.2

Percentage yield — a measure of reaction efficiency.

$$\text{Percentage yield} = (\text{Actual yield} \div \text{Theoretical yield}) \times 100\%$$

PERCENTAGE PURITY

A measure of how pure a sample/reactant is (relevant when reactants used are impure).

$$\text{Percentage purity} = (\text{Amount of pure substance} \div \text{Amount of sample}) \times 100\%$$

CHAPTER 7 QUICK RECAP: mol = mass/M_r = particles/6.02×10^{23} = volume(dm^3)/24 (gas at r.t.p.) = concentration × volume. Empirical formula from mass/% composition ratios; molecular formula = $(\text{empirical})_n$ using M_r. In any reaction, identify the limiting reactant by comparing mole ratios required vs provided; % yield and % purity both compare an actual/available amount against a theoretical/total amount.

Chapter 8: Acids and Bases

Properties, strength, pH, and types of oxides

8.1 | Properties of Acids

DEFINITION 8.1.1 — ACID

A substance which produces hydrogen ions, H^+, in aqueous solution.

DEFINITION 8.1.2 — BASICITY

The number of moles of H^+ ions produced when 1 mole of acid molecules ionises completely in aqueous solution. Basicity determines how many series of salts an acid can form.

Table 8.1.1 — Classification of acids by basicity

Basicity	Examples	Ionisation equation
Monobasic	Hydrochloric acid, nitric acid, ethanoic acid	$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$; $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$; $CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$
Dibasic	Sulfuric acid, carbonic acid	$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$; $H_2CO_3(aq) \rightleftharpoons 2H^+(aq) + CO_3^{2-}(aq)$
Tribasic	Phosphoric acid	$H_3PO_4(aq) \rightleftharpoons 3H^+(aq) + PO_4^{3-}(aq)$

DEFINITIONS 8.1.3 — STRENGTH OF ACIDS

Strength — the extent to which an acid ionises to produce H^+ ions in aqueous solution.
Strong acid — ionises completely in aqueous solution (one-way arrow, $\rightarrow$).
Weak acid — ionises only partially in aqueous solution (reversible arrow, $\rightleftharpoons$, typically only ~1% ionised).

Table 8.1.2 — Strong acids vs weak acids

Strong acids (e.g. HCl)	Weak acids (e.g. CH_3COOH)
Ionises completely; solution contains only H^+, Cl^- and H_2O — no HCl molecules remain	Ionises partially; solution contains H^+, CH_3COO^-, unionised CH_3COOH molecules, and H_2O
Higher $[H^+]$ for same basicity/conc./volume	Lower $[H^+]$ for same basicity/conc./volume
Reacts faster/more vigorously; feels hot	Reacts slower; feels only warm

Concentration of an acid can be increased by increasing the mass of acid dissolved, or decreasing the mass/volume of water used — note that concentration (dilute/concentrated) is independent of strength (strong/weak): a dilute strong acid can have lower $[H^+]$ than a concentrated weak acid.

Physical & Chemical Properties of Acids

- Sour taste
- Good electrical conductors (aqueous), as H^+ and the corresponding anion are mobile and act as charge carriers

Table 8.1.3 — Reactions of dilute acids

Reaction	General equation	Observation
+ reactive metal	metal + dilute acid → salt + hydrogen gas	Colourless, odourless gas (H_2, insoluble in water — effervescence observed)
+ carbonate	carbonate + dilute acid → salt + CO_2 + water	Colourless, odourless gas (CO_2, sparingly soluble)
+ hydrogen carbonate	hydrogencarbonate + dilute acid → salt + CO_2 + water	Effervescence of CO_2
+ metal oxide (base)	metal oxide + dilute acid → salt + water	Neutralisation; oxide dissolves
+ metal hydroxide (base)	metal hydroxide + dilute acid → salt + water	Neutralisation

ⓘ NOTE

Copper, silver, gold and platinum (below hydrogen in the reactivity series) do **not** react with dilute acids. Some reactions appear to stop early — e.g. lead reacting with dilute HCl produces insoluble $PbCl_2$, which coats the metal surface and prevents further reaction.

Table 8.1.4 — Solubility rules for salts

Salt type	Solubility
All Group 1 metal salts and ammonium salts	Soluble
All nitrates	Soluble
Sulfates	Soluble, except lead(II), calcium and barium sulfate
Carbonates	Insoluble, except Group 1 metal carbonates and ammonium carbonate
Halides (Cl^-, Br^-, I^-)	Soluble, except lead(II) and silver halides

ⓘ NOTE — IONIC EQUATIONS

Carbonate + acid: $2H^+(aq) + CO_3^{2-}(aq) \rightarrow CO_2(g) + H_2O(l)$
Neutralisation (alkali + acid): $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

8.2 | Properties of Bases

DEFINITION 8.2.1 — BASE

Any substance that reacts with an acid to produce salt and water. Types of bases: metal oxides, metal hydroxides, carbonates, hydrogen carbonates, aqueous ammonia.

PROPERTY 1 — HIGH MELTING/BOILING POINTS

Metal oxides/hydroxides are typically ionic with a giant ionic lattice structure; strong electrostatic forces between ions require a lot of energy to overcome, giving high m.p./b.p.

DEFINITION 8.2.2 — ALKALI

A base that is soluble in water and produces hydroxide ions, OH^-, in aqueous solution. All Group 1 metal oxides/hydroxides, the last four Group 2 metal oxides/hydroxides (Ca, Sr, Ba... down the group), and aqueous ammonia are alkalis.

Water-soluble metal oxides react with water to form the corresponding metal hydroxide: $K_2O(s) + H_2O(l) \rightarrow 2KOH(aq)$

PROPERTY 3 — GOOD CONDUCTOR WHEN MOLTEN/AQUEOUS

Why does aqueous $Ba(OH)_2$ conduct? Ba^{2+} and OH^- ions are mobile and act as charge carriers.
Why doesn't solid BaO conduct, but its aqueous solution does? In the solid, ions are fixed in the lattice by strong ionic bonds and cannot move; once dissolved, the ions become mobile.

DEFINITIONS 8.2.3 — STRENGTH OF ALKALIS

Strong alkali (e.g. NaOH, KOH) — ionises completely to give OH^-.
Weak alkali (e.g. NH_3, methylamine) — ionises only partially (~1%), leaving mostly unionised NH_3 molecules alongside NH_4^+, OH^- and H_2O.

Properties & Reactions of Alkalis

- Bitter taste; feels soapy; good electrical conductor when aqueous; soluble in water

Table 8.2.1 — Reactions of alkalis

Reaction	General equation	Observation
+ dilute acid	alkali + acid → salt + water	Neutralisation
+ ammonium salt (warmed)	alkali + ammonium salt → salt + ammonia gas + water	Pungent gas evolved (turns moist red litmus blue)
+ aqueous metal B salt	alkali + metal B salt → metal B hydroxide (ppt.) + salt	Precipitation — insoluble hydroxide forms

Table 8.2.2 — Applications of bases/alkalis

Base/Alkali	Application
Aqueous ammonia	Window cleaners; manufacturing nitrogenous fertilisers
Aqueous NaOH	Detergents (dissolves grease)
Calcium hydroxide (slaked lime)	Neutralising acidic soil / acid spills
Magnesium hydroxide	Toothpaste and antacids (neutralises acid)
Zinc oxide	Calamine solution (neutralises bee stings)

8.3 | The pH Scale

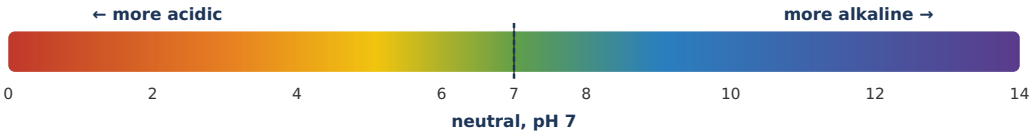


Fig 8.3.1. *The pH scale: lower pH = more acidic ($[H^+] > [OH^-]$); higher pH = more alkaline ($[OH^-] > [H^+]$); pH 7 is neutral.*

Water undergoes slight self-ionisation (a reversible reaction): $H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$. In neutral solutions, $[H^+] = [OH^-]$.

- **Lower pH** → more acidic → $[H^+] > [OH^-]$
- **Higher pH** → more alkaline → $[OH^-] > [H^+]$

Table 8.3.1 — Common indicators

Indicator	pH range for colour change	Colour change (acid → alkali)
Methyl orange	3.2–4.4	Red to yellow
Screened methyl orange	3.8–4.2	Violet to green
Litmus	4.5–8.3	Red to blue
Thymolphthalein	9.3–10.5	Colourless to blue

ⓘ NOTE

A pH meter/datalogger with pH sensor gives a more precise pH reading than an indicator (which only gives an approximate range), but must be calibrated before use.

substances	strong acids	weak acids	neutral aqueous solution & pure water	weak alkalis	strong alkalis
an example of solution	dilute hydrochloric acid	dilute ethanoic acid	salt solutions (e.g. sodium chloride solution)	aqueous ammonia	sodium hydroxide solution
extent of ionisation of acid/alkali	complete	partial	-	partial	complete
concentration of H^+ ions	high	lower than concentration of H^+ ions in strong acids but higher than concentration of H^+ ions in alkalis	concentration of H^+ ions is the same as concentration of OH^- ions	higher than concentration of H^+ ions in strong alkalis but lower than concentration of H^+ ions in acids	low
concentration of OH^- ions	low	higher than concentration of OH^- ions in strong acids but lower than concentration of OH^- ions in alkalis		lower than concentration of OH^- ions in strong alkalis but higher than concentration of OH^- ions in acids	high
presence of water molecules	yes	yes	yes	yes	yes
pH range of solution	0 ≤ pH ≤ 2	2 < pH < 7	pH 7 (no range)	7 < pH < 12	12 ≤ pH ≤ 14
suggested pH value of solution	pH 1	pH 4	pH 7	pH 10	pH 14
Observation made when a few drops of Universal Indicator solution was added to the solution.	Green Universal Indicator turned red.	Green Universal Indicator turned orange.	No visible change. Universal Indicator remained green.	Green Universal Indicator turned blue.	Green Universal Indicator turned violet.

Fig 8.3.2. *Colour of universal indicator across the pH scale, from strongly acidic (red) through neutral (green) to strongly alkaline (purple).*

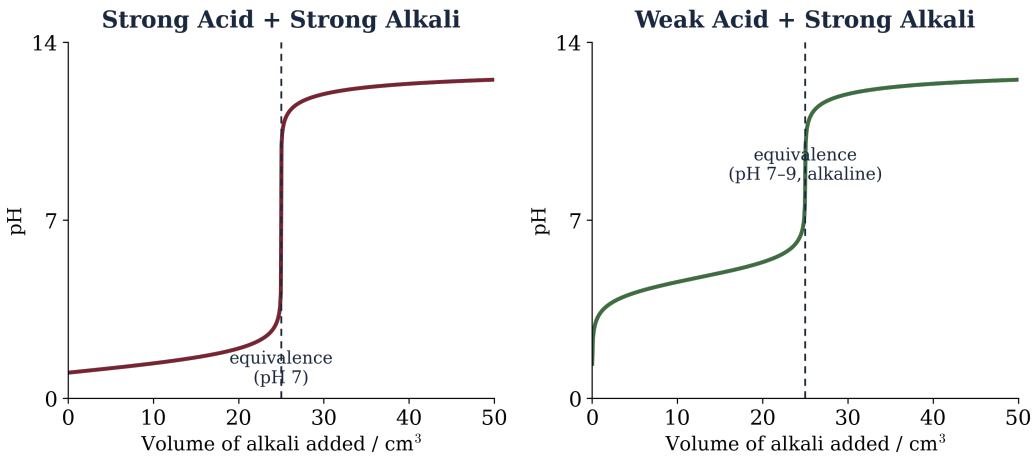


Fig 8.3.3. *Titration curves: a strong acid has a longer, steeper vertical jump centred exactly at pH 7; a weak acid starts at a higher initial pH and its equivalence point is slightly alkaline (above pH 7), since the conjugate base formed is itself weakly basic.*

KEY QNS 8.3.1 — TITRATION CURVE OF STRONG ACID + STRONG ALKALI (HCL + NAOH)

Initially, only HCl is present so pH is low. As NaOH is added drop by drop, pH rises only gradually, since excess H^+ is still present (pH stays below 7). Only at the exact equivalence point (equal moles reacted) does complete neutralisation occur and pH rises sharply through 7. Beyond this point, excess NaOH causes pH to rise further towards 13–14.

KEY QNS 8.3.2 — TITRATION CURVE OF WEAK ACID + STRONG ALKALI (ETHANOIC ACID + NAOH)

At 0 cm^3 NaOH, only ethanoic acid (weak, partially dissociated) is present, so initial pH (~2.5–3) is higher than for an equivalent strong acid. As NaOH is added, OH^- reacts with H^+, raising pH. At the equivalence point (all ethanoic acid converted to CH_3COONa), the solution is slightly alkaline (pH around 8–9, since ethanoate ion is a weak base) rather than exactly 7. Beyond this, excess NaOH pushes pH toward 13.

KEY QNS 8.3.3 — CHOOSING INDICATORS

Q. Why is methyl orange unsuitable for titrating ethanoic acid with NaOH?
A. The steep, sharp pH change (equivalence) for this titration occurs roughly between pH 7.5 and 11. Methyl orange changes colour at pH 3.2–4.4, well before this range, so it would give a false endpoint.

8.4 | Types of Oxides

DEFINITION 8.4.1 — OXIDE

A binary compound of oxygen and one other element.

Table 8.4.1 — Classification of oxides

	Basic oxides	Amphoteric oxides	Acidic oxides	Neutral oxides
Made from	Metal + O_2 (most metal oxides)	ZnO, Al_2O_3, PbO	Non-metal + O_2 (most non-metal oxides)	Non-metal + O_2 (e.g. CO, H_2O, NO)
Behaviour	Basic properties only	Reacts as both a base (with acid) and an acid (with base)	Acidic properties only	Generally unreactive with acids/bases
In water	Water-soluble ones form alkalis	Insoluble in water	Water-soluble ones form acidic solutions	No reaction (exception: water itself)

DEFINITIONS 8.4.2

Basic oxide — a compound of oxygen and a metal that reacts with acids to form salt and water.
Metal(s) + O_2(g) → basic oxide(s); basic oxide + dilute acid → salt + water; basic oxide + acidic oxide → salt.

DEFINITION 8.4.3 — AMPHOTERIC OXIDE

A metal oxide that reacts with both acids and bases to form salt and water.

Reacting as a base	amphoteric oxide + dilute acid → salt + water; amphoteric oxide + acidic oxide → salt
Reacting as an acid	amphoteric oxide + metal hydroxide → salt + water; amphoteric oxide + basic oxide → salt

WORKED EXAMPLE

$ZnO(s) + 2NaOH(aq) \rightarrow Na_2ZnO_2(aq) + H_2O(l)$ — zinc oxide acting as an acid with a strong alkali, forming aqueous sodium zincate.

DEFINITION 8.4.4 — ACIDIC OXIDE

A compound of oxygen and a non-metal that reacts with bases to form salt and water.

Non-metal(g) + O_2(g) → acidic oxide(s); acidic oxide + metal hydroxide → salt + water; acidic oxide + basic/amphoteric oxide → salt.

DEFINITION 8.4.5 — NEUTRAL OXIDE

A compound of oxygen and a non-metal that shows neither acidic nor basic properties (e.g. CO, H_2O, NO).

ⓘ NOTE

Water reacts with water-soluble metal oxides to form metal hydroxide solutions, and with acidic oxides to form acidic solutions (e.g. $SO_2 + H_2O \rightarrow H_2SO_3$).

CHAPTER 8 QUICK RECAP: Acids produce H^+ (aq); bases react with acids to give salt+water, with alkalis being the water-soluble bases producing OH^-. Strength (strong/weak = extent of ionisation) is independent of concentration (dilute/concentrated = amount dissolved). pH reflects $[H^+]$ vs $[OH^-]$; choose indicators whose colour-change range matches the titration's steep region. Oxides are basic (metal), acidic (non-metal), amphoteric (react with both acid and base, e.g. ZnO, Al_2O_3, PbO) or neutral (CO, H_2O, NO).

Chapter 9: Salts

Solubility and the three main preparation methods

9.1 | Solubility of Salts

DEFINITION 9.1.1 — SALT

A salt is an ionic compound consisting of a cation and an anion (any anion *except* oxide, O^{2-}, and hydroxide, OH^-, which instead classify the compound as an oxide/hydroxide, i.e. a base).

Table 9.1.1 — Solubility rules for salts

Salt type	Solubility
All Group 1 metal salts and ammonium salts	Soluble
All nitrates	Soluble
Chlorides	Soluble, except lead(II) chloride and silver chloride
Sulfates	Soluble, except lead(II) sulfate, calcium sulfate and barium sulfate
Carbonates	Insoluble, except Group 1 metal carbonates and ammonium carbonate
All other lead(II) and silver halides	Insoluble

9.2 | Preparation of Salts

The correct method depends on the **solubility** of the target salt.

Table 9.2.1 — Choosing a method of salt preparation

Method	Used for
Reaction of acid with excess insoluble substance (metal / insoluble base / insoluble carbonate)	Most soluble salts, EXCEPT Group 1 metal salts and ammonium salts
Titration	Soluble salts of Group 1 metals and ammonium salts
Precipitation (double decomposition)	Insoluble salts

Method 1: Reaction of Acid with Excess Insoluble Reactant

Used because the insoluble reactant (metal, insoluble base or insoluble carbonate) can simply be added in excess and later filtered off, without needing to judge an exact equivalence point.

PROCEDURE

1. Add excess insoluble reactant (e.g. copper(II) oxide) to a measured volume of warm dilute acid (e.g. dilute H_2SO_4), stirring until no more dissolves/reacts (effervescence stops, or excess solid visibly remains).
2. Filter to remove the unreacted excess solid, collecting the salt solution as filtrate.
3. Heat the filtrate gently to evaporate some water, until a hot saturated solution forms (test using a glass rod: solution crystallises on the rod when cool).
4. Allow the hot saturated solution to cool slowly, allowing crystals of the salt to form (crystallisation).
5. Filter to collect the crystals, wash with a little cold distilled water, and dry between sheets of filter paper / in a desiccator.

ⓘ NOTE

Copper, silver, gold and platinum cannot be used, as they do not react with dilute acids (below hydrogen in the reactivity series). Group 1 metal salts and ammonium salts cannot be prepared this way as Group 1 metals/carbonates react too vigorously (dangerously) with acids, and ammonia is a gas, not a solid reactant.

Method 2: Titration

Used for soluble salts of Group 1 metals and ammonium, since the "insoluble excess reactant" trick of Method 1 does not apply — both the alkali/soluble carbonate and the acid are soluble, so the exact reacting volumes must be found using an indicator.

PROCEDURE

1. Using a pipette, transfer a fixed volume of the alkali (e.g. NaOH) into a conical flask; add a few drops of indicator (e.g. phenolphthalein/methyl orange).
2. Fill a burette with the acid (e.g. dilute HCl) and titrate into the flask, swirling, until the indicator just changes colour (the endpoint) — record the volume of acid used.
3. Repeat *without indicator*, using the same volumes of alkali and (now known) acid, to obtain a pure salt solution free of indicator.
4. Heat the salt solution to evaporate to a hot, saturated solution; cool slowly to crystallise; filter, wash and dry the crystals, as in Method 1.

Applicable reactions: acid + alkali → salt + water; acid + soluble carbonate → salt + water + CO_2.

Method 3: Precipitation (Double Decomposition)

KEY REQUIREMENT

Both reactants must be soluble (aqueous) salt solutions; the reaction produces an insoluble salt (precipitate) plus a second product made up only of the remaining spectator ions.

WORKED EXAMPLE — PREPARING BARIUM SULFATE

$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$

PROCEDURE

1. Mix aqueous solutions of the two soluble salts (chosen so that the desired insoluble salt precipitates, and the other product remains soluble).
2. Stir; the insoluble salt forms immediately as a precipitate.
3. Filter to collect the precipitate as residue.
4. Wash the residue with distilled water (to remove soluble impurities/spectator ions), then leave the precipitate in a cool place to dry by standing in air.

COMMON EXAM PITFALL

When choosing reactants for precipitation, always check that *both* starting salts are soluble (using the solubility rules above) — otherwise, the reaction cannot proceed in aqueous solution, and you must instead use Method 1 or 2.

CHAPTER 9 QUICK RECAP: Salt solubility follows fixed rules (Group 1/NH_4^+/NO_3^- always soluble; most Cl^-/SO_4^{2-} soluble bar a few exceptions; most CO_3^{2-} insoluble bar Group 1/NH_4^+). Choose the method by solubility of the target salt: excess insoluble reactant + filter + crystallise (most soluble salts); titration (soluble Group 1/ammonium salts); precipitation + filter + wash + dry (insoluble salts).

Chapter 10: Ammonia

Reversible reactions and the Haber Process

10.1 | Reversible and Irreversible Reactions

DEFINITION 10.1.1 — IRREVERSIBLE REACTION

A chemical change in which reactants convert to products, and the products cannot convert back to the reactants. Represented with a one-way arrow, $\rightarrow$.

Table 10.1.1 — Characteristics and examples of irreversible reactions

Characteristics	Proceeds in one (forward) direction only; can proceed to completion (all reactants consumed)
Examples	Complete ionisation of strong acids/alkalis, e.g. $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$ Neutralisation of strong acid + strong base, e.g. $2HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2(aq) + 2H_2O(l)$ Precipitation reactions, e.g. $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$ Displacement reactions (metals and halogens), e.g. $Zn(s) + CuCl_2(aq) \rightarrow ZnCl_2(aq) + Cu(s)$ Combustion, e.g. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ Thermal decomposition (most compounds), e.g. $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$

KEY QNS 10.1.1 — EXPLAINING IRREVERSIBILITY OF A DISPLACEMENT REACTION

Q. Why is the reaction between zinc and copper(II) chloride solution irreversible?

A. $Zn(s) + CuCl_2(aq) \rightarrow ZnCl_2(aq) + Cu(s)$. Copper is less reactive than zinc (loses electrons less readily), so copper cannot displace zinc from zinc chloride solution — the backward reaction cannot occur, so the forward reaction is irreversible.

DEFINITION 10.1.2 — REVERSIBLE REACTION

A chemical change in which the reactants are not entirely converted to products; the reaction proceeds in both directions simultaneously. Represented with the double arrow, $\rightleftharpoons$ (forward reaction reads left $\rightarrow$ right; backward reaction reads right $\rightarrow$ left).

Table 10.1.2 — Characteristics of reversible reactions

Characteristic
Can be started from either side (reactants or products)
Forward and backward reactions occur simultaneously
Reactants and products are never fully consumed

Examples: dissociation of weak acids/alkalis in aqueous solution; thermal decomposition of ammonium chloride and hydrated copper(II) sulfate; and the Haber process (below).

10.2 | The Haber Process

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \text{(forward reaction is exothermic)}$$

Table 10.2.1 — Industrial conditions for the Haber Process

Condition	Value
Temperature	~450°C
Pressure	~250 atm
Catalyst	Finely-divided iron

Table 10.2.2 — Justifying the "compromise" conditions

Condition	Reasoning
Finely-divided iron catalyst	Increases the rate of the forward (and backward) reaction, allowing equilibrium (and hence a usable yield) to be reached faster, without affecting the position of equilibrium/yield
Compromise (moderate) temperature, 450°C	Since the forward reaction is exothermic, a *lower* temperature would favour a *higher* yield of NH_3 (less decomposition of ammonia). However, too low a temperature makes the reaction rate too slow to be economical. 450°C is chosen as a compromise between a reasonable rate and a reasonable yield.
Compromise (high) pressure, 250 atm	A *higher* pressure favours the forward reaction (which produces fewer gas molecules: 4 mol reactants → 2 mol product), increasing yield. However, very high pressures are costly to generate and maintain, and pose safety/engineering challenges. 250 atm is chosen as an economic compromise.

Sourcing the Raw Materials

Nitrogen	Obtained by fractional distillation of liquefied air (b.p. of N_2 = −196°C, lower than O_2's −183°C, so nitrogen distills off first)
Hydrogen	Obtained from cracking of crude oil fractions, or from the electrolysis of certain aqueous solutions/water

Uses and Environmental Impact of Ammonia

Ammonia's main use is manufacturing **nitrogenous fertilisers** (e.g. ammonium nitrate, ammonium sulfate) to supply nitrogen for plant growth (needed to make proteins/chlorophyll).

> ⓘ **NOTE — PROBLEMS CAUSED BY NITROGENOUS FERTILISERS**
>
> Loss of nitrogen from soil occurs if ammonium fertilisers are mixed with an alkali such as slaked lime (calcium hydroxide) or quicklime (calcium oxide), since alkali + ammonium salt → salt + **ammonia gas** (lost to air) + water — wasting the fertiliser's nitrogen content. Excess fertiliser can also be washed into waterways (leaching/run-off), causing **eutrophication**: excessive algal growth blocks sunlight, and when the algae die and decompose, bacteria deplete dissolved oxygen, killing aquatic life.

CHAPTER 10 QUICK RECAP: Irreversible reactions (→) go to completion in one direction; reversible reactions (⇌) proceed both ways simultaneously and never fully consume reactants/products. The Haber process ($N_2+3H_2 \rightleftharpoons 2NH_3$, exothermic forward reaction) uses a compromise of 450°C, 250 atm and an iron catalyst to balance rate and yield economically.

Chapter 11: Qualitative Analysis

Tests for cations, anions, and gases

11.1 | Tests for Cations

> ⓘ **NOTE**
>
> Both alkalis provide OH^- ions, which react with metal cations to form an insoluble metal hydroxide precipitate: $M^{n+}(aq) + nOH^-(aq) \rightarrow M(OH)_n(s)$

Pathway 1: Aqueous Sodium Hydroxide, NaOH(aq)

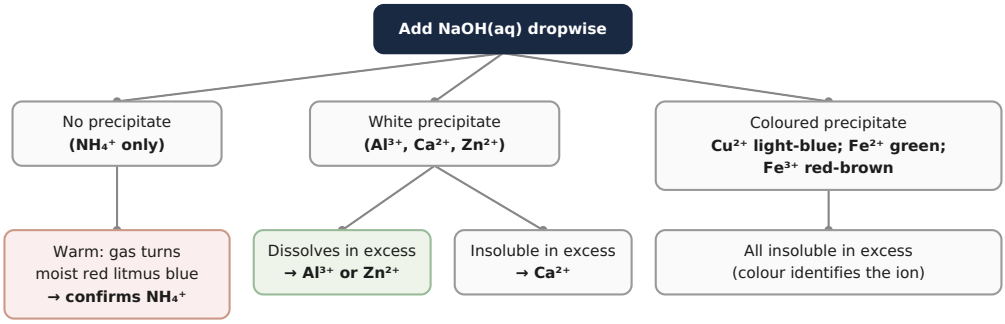


Fig 11.1.1. *Cation-testing decision flow using aqueous sodium hydroxide, dropwise then in excess.*

Pathway 2: Aqueous Ammonia, $NH_3(aq)$

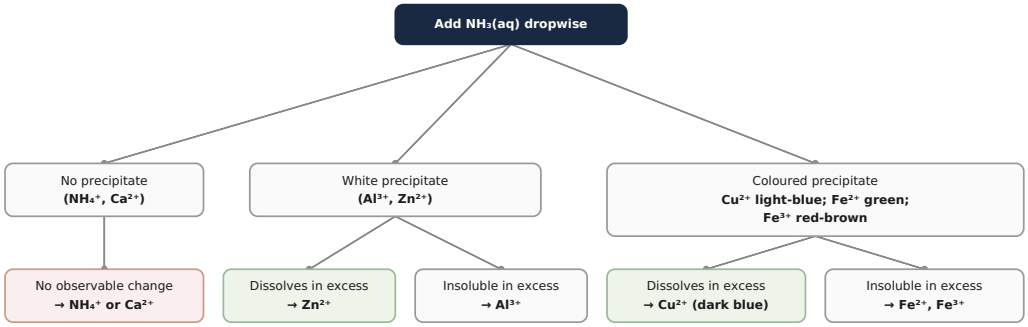


Fig 11.1.2. *Cation-testing decision flow using aqueous ammonia, dropwise then in excess.*

Table 11.1.1 — Cation tests and observations

Cation	+ NaOH(aq), dropwise then excess	+ NH_3(aq), dropwise then excess
Aluminium, Al^{3+}	White ppt.; **dissolves** in excess to colourless solution	White ppt.; **insoluble** in excess
Ammonium, NH_4^+	No visible ppt.; on warming, pungent gas (NH_3) evolved, turns moist red litmus blue	No observable change
Calcium, Ca^{2+}	White ppt.; **insoluble** in excess	No observable change
Copper(II), Cu^{2+}	Light-blue ppt.; **insoluble** in excess	Light-blue ppt.; **dissolves** in excess to a dark blue solution
Iron(II), Fe^{2+}	Green ppt.; **insoluble** in excess (turns red-brown on standing in air, as $Fe(OH)_2$ is oxidised to $Fe(OH)_3$)	Green ppt.; **insoluble** in excess
Iron(III), Fe^{3+}	Red-brown ppt.; **insoluble** in excess	Red-brown ppt.; **insoluble** in excess
Zinc, Zn^{2+}	White ppt.; **dissolves** in excess to colourless solution	White ppt.; **dissolves** in excess to colourless solution

Table 11.1.2 — Colours of common metal hydroxide precipitates

Hydroxide	Colour
Aluminium hydroxide	White
Calcium hydroxide	White
Copper(II) hydroxide	Light blue
Iron(II) hydroxide	Green
Iron(III) hydroxide	Red-brown
Zinc hydroxide	White

> **COMMON EXAM PITFALL**
>
> Al^{3+} and Zn^{2+} give white precipitates that dissolve in *excess* NaOH, but only Zn^{2+} (not Al^{3+}) also dissolves in *excess* NH_3. This difference (behaviour in excess NH_3) is the key distinguishing test between the two.

11.2 | Tests for Anions

Table 11.2.1 — Anion tests and observations

Anion	Reagents / Procedure	Observation
Carbonate, CO_3^{2-}	Add any dilute acid; bubble gas into limewater	Effervescence of a colourless, odourless gas; white ppt. forms in limewater — gas is CO_2
Chloride, Cl^-	Add dilute HNO_3, then aqueous $AgNO_3$	No change with HNO_3; white precipitate with $AgNO_3$, which dissolves in excess aqueous NH_3 to a colourless solution
Iodide, I^-	Add dilute HNO_3, then aqueous $AgNO_3$	No change with HNO_3; yellow precipitate with $AgNO_3$ (insoluble in excess NH_3)
Nitrate, NO_3^-	Add aqueous NaOH and a piece of aluminium foil; warm carefully	Effervescence of a colourless, pungent gas that turns moist red litmus blue — gas is NH_3 (aluminium reduces NO_3^- to NH_3 in alkaline conditions)
Sulfate, SO_4^{2-}	Add dilute HNO_3, then aqueous $Ba(NO_3)_2$	No change with HNO_3; white precipitate with barium nitrate

11.3 | Tests for Gases

Table 11.3.1 — Colour and odour of common gases

Gas	Colour	Odour
Ammonia, NH_3	Colourless	Pungent
Carbon dioxide, CO_2	Colourless	Odourless
Chlorine, Cl_2	Pale yellowish-green	Pungent
Hydrogen, H_2	Colourless	Odourless
Oxygen, O_2	Colourless	Odourless
Sulfur dioxide, SO_2	Colourless	Pungent
Hydrogen chloride, HCl	Colourless	Pungent
Nitrogen dioxide, NO_2	Brown	Pungent

Table 11.3.2 — Gas identification tests

Gas	Test	Positive observation
Ammonia, NH_3	Hold moist red litmus paper near the mouth of the tube	Turns moist red litmus paper blue (NH_3 is alkaline)
Carbon dioxide, CO_2	Bubble gas through limewater	White ppt. forms in limewater ($CaCO_3$); with excess CO_2, precipitate redissolves to a colourless solution (forms soluble $Ca(HCO_3)_2$)
Chlorine, Cl_2	Hold moist blue litmus paper near the mouth of the tube	Litmus turns red, then is bleached white ($Cl_2 + H_2O \rightarrow HCl + HClO$; HClO bleaches)
Hydrogen, H_2	Hold a lighted splint near the mouth of the tube	Extinguishes with a "pop" sound
Oxygen, O_2	Insert a glowing splint into the tube	Splint re-ignites (relights)
Sulfur dioxide, SO_2	Hold filter paper soaked in acidified potassium manganate(VII) near the mouth of the tube (or bubble gas through the solution)	Purple acidified $KMnO_4$ is decolourised (turns colourless)

ⓘ NOTE — CHEMISTRY OF THE SO_2 TEST

SO_2 is the reducing agent: it reduces purple MnO_4^- to colourless Mn^{2+} (Mn oxidation state falls from +7 to +2), while SO_2 itself is oxidised to SO_4^{2-} (S oxidation state rises from +4 to +6). See Chapter 12 for oxidation states.

ⓘ NOTE — APPEARANCE OF COMMON COMPOUNDS

Group 1, Group 2, Group 13 and ammonium compounds are typically white solids that form colourless solutions. Zn^{2+} compounds are white solid when cold, but turn yellow when hot (reversibly), and form colourless solutions.

CHAPTER 11 QUICK RECAP: Cations are identified by precipitate colour with $NaOH/NH_3$ and solubility behaviour in excess (Al^{3+} dissolves only in excess NaOH; Zn^{2+} dissolves in both; Cu^{2+} dissolves only in excess NH_3). Anions are identified via characteristic precipitates ($AgNO_3$ for halides, $Ba(NO_3)_2$ for sulfate) or gas evolution (acid for carbonate, NaOH+Al foil for nitrate). Gases are identified by distinct tests: litmus (NH_3, Cl_2), limewater (CO_2), lit/glowing splint (H_2, O_2), and acidified $KMnO_4$ (SO_2).

Chapter 12: Oxidation and Reduction

Four ways to define redox, oxidation states, and testing for oxidising/reducing agents

---

DEFINITION 12.1 — REDOX REACTION

A chemical reaction in which both oxidation and reduction occur simultaneously.

Table 12.1 — Four equivalent ways to define oxidation and reduction

In terms of...	Oxidation	Reduction
Oxygen	Gain of oxygen	Loss of oxygen
Hydrogen	Loss of hydrogen	Gain of hydrogen
Electrons	Loss of electrons	Gain of electrons
Oxidation state	Increase in oxidation state	Decrease in oxidation state

12.1 | Oxidation and Reduction in Terms of Oxygen

DEFINITIONS

Oxidation — a reaction in which oxygen combines with an element/compound.
Reduction — a reaction in which oxygen is removed from a compound.
Oxidising agent — a substance that *loses* oxygen to another substance, causing it to be oxidised (itself being reduced).
Reducing agent — a substance that *gains* oxygen from another substance, causing it to be reduced (itself being oxidised).

WORKED EXAMPLE — $CUO + H_2$

$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$. H_2 gains oxygen from CuO to form H_2O — H_2 is oxidised. CuO loses oxygen to H_2 to form Cu — CuO is reduced. Since both occur simultaneously, this is a redox reaction.
Oxidising agent: CuO (loses O to H_2, causing H_2 to be oxidised).
Reducing agent: H_2 (gains O from CuO, causing CuO to be reduced).

12.2 | Oxidation and Reduction in Terms of Hydrogen

DEFINITIONS

Oxidation — a reaction in which hydrogen is removed from a compound.
Reduction — a reaction in which hydrogen combines with an element/compound.
Oxidising agent — gains hydrogen from another substance, causing it to be oxidised.
Reducing agent — loses hydrogen to another substance, causing it to be reduced.

WORKED EXAMPLE — $H_2S + CL_2$

$H_2S + Cl_2 \rightarrow S + 2HCl$. H_2S loses hydrogen to Cl_2 to form S — H_2S is oxidised. Cl_2 gains hydrogen from H_2S to form HCl — Cl_2 is reduced.
Oxidising agent: Cl_2. **Reducing agent:** H_2S.

12.3 | Oxidation and Reduction in Terms of Electron Transfer

DEFINITIONS

Oxidation — loss of electrons.
Reduction — gain of electrons.
Oxidising agent — gains electrons from another substance, causing it to be oxidised.
Reduction agent — loses electrons to another substance, causing it to be reduced.

WORKED EXAMPLE — NABR + CL_2

$2Br^-(aq) \rightarrow Br_2(aq) + 2e^-$ — Br^- loses electrons to Cl_2, so NaBr is oxidised.
$Cl_2(aq) + 2e^- \rightarrow 2Cl^-(aq)$ — Cl_2 gains electrons from Br^-, so Cl_2 is reduced.
Oxidising agent: Cl_2 (gains electrons from Br^-, oxidising it to Br_2). **Reducing agent:** NaBr (loses electrons to Cl_2, reducing it to Cl^-).

12.4 | Oxidation States

DEFINITION 12.4.1 — OXIDATION STATE

The charge an atom of an element would have if it existed as an ion in a compound.

Table 12.2 — Rules for assigning oxidation states

Species	Oxidation state
Any element (uncombined)	0
Group 1 metals (in compounds)	+1
Group 2 metals (in compounds)	+2
Fluorine (in compounds)	−1
Hydrogen (usually, in compounds)	+1
Oxygen (usually, in compounds)	−2
Other elements in a compound	Deduced so the sum of oxidation states = 0 (neutral compound) or = overall ionic charge

DEFINITIONS USING OXIDATION STATE

Oxidation — a reaction in which a reactant's oxidation state increases (loses electrons).
Reduction — a reaction in which a reactant's oxidation state decreases (gains electrons).
Oxidising agent — causes another reactant's oxidation state to increase.
Reducing agent — causes another reactant's oxidation state to decrease.

WORKED EXAMPLE — CU + O_2

$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$. Cu's oxidation state increases from 0 to +2 — Cu is oxidised. O's oxidation state decreases from 0 to −2 — O_2 is reduced.
Oxidising agent: O_2. **Reducing agent:** Cu.

Table 12.3 — Redox vs non-redox reactions, and common oxidising/reducing agents

Redox reactions	**Non-redox reactions**
Combustion (e.g. Cu + O_2), respiration, photosynthesis, displacement reactions	Neutralisation, precipitation
Common oxidising agents	**Common reducing agents**
Halogens (Cl_2, Br_2), oxygen gas, hydrogen peroxide, acidified potassium manganate(VII)	Reactive metals, hydrogen gas, carbon, carbon monoxide, hydrogen peroxide, potassium iodide, sulfur dioxide

> ⓘ **NOTE**
>
> Hydrogen peroxide can act as either an oxidising or a reducing agent depending on what it reacts with — this dual behaviour is a common exam talking point.

12.5 | Testing for Oxidising and Reducing Agents

Table 12.4 — Chemical tests

	Test for an oxidising agent	**Test for a reducing agent**
Reagent	Aqueous potassium iodide, KI (colourless)	Acidified potassium manganate(VII) (purple)
Procedure	Add a few drops of KI(aq) to the solution being tested (for a gas: use starch-iodide paper — see note below)	Add a few drops of acidified $KMnO_4$ to the solution; for a gas, place filter paper soaked in acidified $KMnO_4$ at the mouth of the test tube
Positive observation	Colourless KI(aq) turns brown	Purple acidified $KMnO_4$ is decolourised
Explanation	$2I^-(aq) \rightarrow I_2(aq) + 2e^-$: the oxidising agent oxidises colourless I^- to brown I_2, increasing iodine's oxidation state from −1 to 0	The reducing agent reduces purple MnO_4^- to colourless Mn^{2+}, decreasing manganese's oxidation state from +7 to +2

> ⓘ **NOTE**
>
> Starch-iodide paper (starch + KI) can also test for oxidising agents/gases: an oxidising agent turns moist starch-iodide paper from white to blue-black.

CHAPTER 12 QUICK RECAP: Oxidation = gain O / loss H / loss e^- / rise in oxidation state; reduction is the reverse. An oxidising agent gets reduced while oxidising something else; a reducing agent gets oxidised while reducing something else. Test for oxidising agents with KI(aq) (colourless → brown); test for reducing agents with acidified $KMnO_4$ (purple → colourless).

Chapter 13: Electrochemistry

Electrolysis, selective discharge, purification/electroplating, simple cells and fuel cells

13.1 | Electrolysis

DEFINITION 13.1.1 — ELECTROLYSIS

The process of using electricity to break down (decompose) a compound called an **electrolyte**, which must be molten or dissolved in water. This provides evidence that ions exist, held in fixed positions in the solid lattice but free to move when molten/aqueous.

DEFINITION 13.1.2 — ELECTROLYTE

A compound that conducts electricity when molten or in aqueous solution, and decomposes as it does so, due to the presence of mobile cations and anions that act as charge carriers.

Electrolysis takes place in an **electrolytic cell**, which uses electrical energy to drive a *non-spontaneous* chemical (redox) reaction. Energy conversion: **electrical energy → chemical energy**.

Parts of an electrolytic cell: battery; anode (positive electrode, connected to + terminal); cathode (negative electrode, connected to − terminal); connecting wires; electrolyte; (optional: bulb/ammeter).

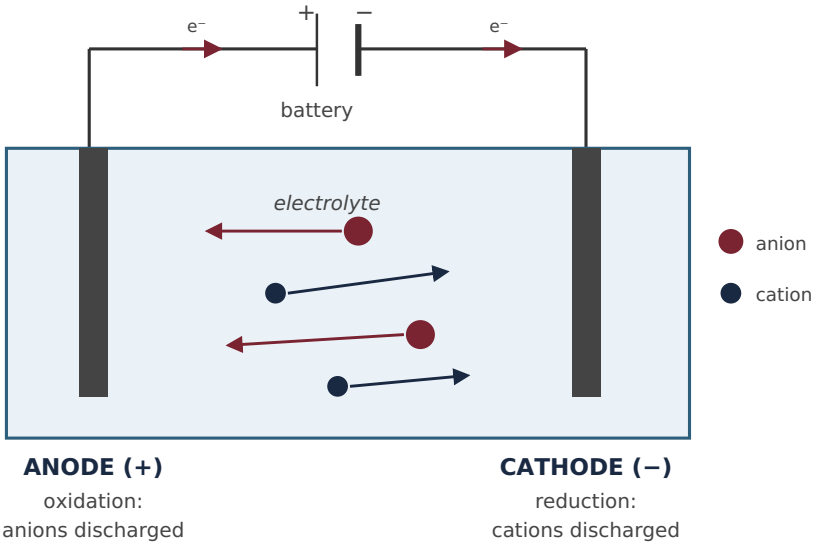


Fig 13.1.1. *A basic electrolytic cell: battery drives current through the electrolyte via two electrodes; cations migrate to the cathode, anions to the anode.*

Table 13.1.1 — Inert vs reactive electrodes

	Inert electrodes	Reactive electrodes
Material	Graphite (an allotrope of carbon), platinum	Any metal more reactive than platinum
Behaviour at anode	Remains chemically unchanged; does not interfere with the discharge of ions	Undergoes oxidation itself (dissolves into the electrolyte); interferes with/prevents normal anion discharge
Behaviour at cathode	Remains chemically unchanged in both cases; does not interfere with ion discharge	

Electrolysis of Molten Binary Compounds (Inert Electrodes)

> **WORKED EXAMPLE — MOLTEN SODIUM CHLORIDE**
>
> Molten NaCl contains only two types of mobile ions: Na^+ and Cl^-.
> At the cathode (reduction): $Na^+(l) + e^- \rightarrow Na(l)$
> At the anode (oxidation): $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$
> Overall: $2NaCl(l) \rightarrow 2Na(l) + Cl_2(g)$

> **ⓘ NOTE**
>
> For a molten binary compound, there is no "choice" of which ion is discharged — the metal cation is always reduced at the cathode and the non-metal anion always oxidised at the anode, since these are the only ions present. This method extracts highly reactive metals (Group 1, Group 2, Al) that cannot be obtained by reduction with carbon.

13.2 | Selective Discharge in Aqueous Electrolysis (Inert Electrodes)

In an aqueous electrolyte, extra ions H^+ and OH^- (from the self-ionisation of water) are present alongside the electrolyte's own ions, so a choice ("selective discharge") must be made at each electrode.

Table 13.2.1 — Rules for selective discharge

Electrode	Rule
Cathode (cations)	The cation **lower** in the reactivity series is preferentially discharged (reduced), as it gains electrons more readily. If the metal is above hydrogen in reactivity, H^+ is discharged instead (producing H_2 gas) and the metal ion remains in solution.
Anode (anions), dilute solution	OH^- is preferentially discharged (oxidised) over other anions (except halides at high concentration), as it loses electrons more readily, producing O_2 gas.
Anode (anions), concentrated halide solution	The halide ion (Cl^-, Br^-, I^-) is preferentially discharged instead of OH^-, since [halide] $\gg$ [OH^-], producing the halogen gas.
Anode, sulfate/nitrate solutions	SO_4^{2-} and NO_3^- are never discharged (too stable); OH^- is discharged instead, producing O_2 gas.

> **WORKED EXAMPLE — AQUEOUS COPPER(II) SULFATE, INERT ELECTRODES**
>
> Ions present: Cu^{2+}, H^+ (cations); SO_4^{2-}, OH^- (anions).
> Cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ — Cu is below hydrogen, so Cu^{2+} is preferentially discharged (pink/brown deposit).
> Anode: $4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$ — SO_4^{2-} is never discharged, so OH^- is (colourless gas, relights glowing splint).
> As electrolysis proceeds, the blue colour of the solution fades (Cu^{2+} is removed).

> **WORKED EXAMPLE — DILUTE SODIUM CHLORIDE SOLUTION (ESSENTIALLY ELECTROLYSIS OF WATER)**
>
> Ions present: Na^+, H^+ (cations); Cl^-, OH^- (anions).
> Cathode: $2H^+(aq) + 2e^- \rightarrow H_2(g)$ — Na is more reactive than hydrogen, so H^+ is discharged instead.
> Anode: $4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$ — dilute solution, so OH^- > Cl^- in relative concentration, and OH^- is discharged.
> Overall this is effectively the electrolysis of water: $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$, producing H_2 and O_2 in a 2:1 volume ratio.

> **WORKED EXAMPLE — CONCENTRATED SODIUM CHLORIDE (BRINE)**
>
> Cathode: $2H^+(aq) + 2e^- \rightarrow H_2(g)$ (same as above).
> Anode: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$ — now [Cl^-] $\gg$ [OH^-], so Cl^- is discharged instead (used industrially to manufacture chlorine gas, hydrogen gas and sodium hydroxide — the chlor-alkali process).

13.3 | Electrolysis with Reactive Electrodes

If the anode is reactive, the electrode itself is oxidised (dissolves) in preference to any anion in solution: **Anode(s) → Anode^{n+}(aq) + ne$^-$**

Table 13.3.1 — Applications of reactive-electrode electrolysis

Application	Anode	Cathode	Electrolyte
Purification of copper	Impure copper	Pure copper (thin sheet)	Aqueous copper(II) sulfate
Electroplating (e.g. copper-plating)	The plating metal (e.g. copper)	Object to be plated	Aqueous solution containing the plating metal's cations

Purification of copper: the impure copper anode dissolves (Cu(s) → Cu^{2+}(aq) + 2e$^-$); pure Cu^{2+} ions are deposited on the pure copper cathode (Cu^{2+}(aq) + 2e$^-$ → Cu(s)); insoluble impurities fall to the bottom as "anode sludge" (often containing valuable metals like silver/gold).

Electroplating coats a (often cheaper/less resistant) object with a thin, even layer of a more attractive or corrosion-resistant metal — e.g. silver-plating cutlery, chromium-plating car parts — for appearance and/or protection against corrosion.

13.4 | Simple Cells (Voltaic/Galvanic Cells)

DEFINITION 13.4.1 — SIMPLE CELL

A device that uses spontaneous chemical (redox) reactions to generate electricity. Energy conversion: **chemical energy → electrical energy** (the opposite of an electrolytic cell).

Parts: two metal electrodes of **different reactivity**, an electrolyte, connecting wires, and a bulb/voltmeter (no battery is present, unlike an electrolytic cell).

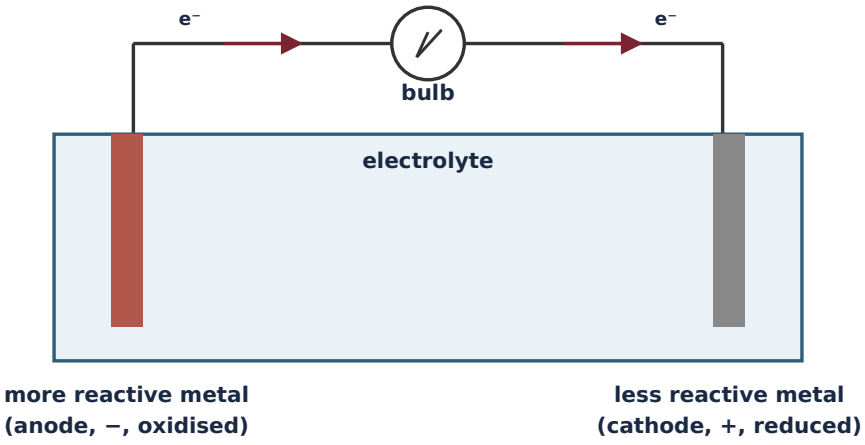


Fig 13.4.1. *A simple (voltaic) cell: no battery is needed — the spontaneous reaction between two different metals in an electrolyte itself drives electron flow through the external circuit.*

Table 13.4.1 — Anode and cathode in a simple cell

Anode (negative electrode)	The **more reactive** metal; loses electrons more readily, so undergoes **oxidation**: M(s) → M^{n+}(aq) + ne$^-$. It is "negative" because electrons are produced/released here.
Cathode (positive electrode)	The **less reactive** metal; site of **reduction** of cations from the electrolyte, e.g. $2H^+$(aq) + 2e$^-$ → H_2(g). Electrons flow through the external circuit from anode to cathode.

ⓘ NOTE

The **greater the difference in reactivity** between the two electrode metals, the greater the potential difference (voltage) produced, and the brighter the bulb glows. Selective discharge of cations at the cathode still applies; selective discharge of anions does not apply, as the reactive anode itself is oxidised instead of any anion.

WORKED EXAMPLE — MG & AG ELECTRODES IN DILUTE H_2SO_4

Mg is more reactive than Ag, so Mg is the anode (negative) and Ag is the cathode (positive).
Anode: $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$ (Mg electrode gradually dissolves/loses mass)
Cathode: $2H^+(aq) + 2e^- \rightarrow H_2(g)$ (effervescence at the silver electrode)
Overall: $Mg(s) + 2H^+(aq) \rightarrow H_2(g) + Mg^{2+}(aq)$

ⓘ NOTE — CHOICE OF ELECTROLYTE AFFECTS CELL LIFETIME

Using a salt solution (e.g. KCl) instead of a dilute acid means the reactive metal electrode corrodes/reacts more slowly (a salt reacts with the metal much less readily than an acid does), so the simple cell generates electricity for a longer period of time before the reactive electrode is used up.

13.5 | The Hydrogen–Oxygen Fuel Cell

A fuel cell is similar to a simple cell, but the reactants (fuel and oxidiser) are continuously supplied rather than being part of the electrode itself.

Anode (negative electrode)	H_2 gas supplied; oxidised: $2H_2(g) + 4OH^-(aq) \rightarrow 4H_2O(l) + 4e^-$
Cathode (positive electrode)	O_2 gas supplied; reduced: $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$
Overall	$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ (H_2:O_2 consumed in a 2:1 ratio)

Electrons released at the anode travel through the external circuit to the cathode, generating electricity directly (used to power electric vehicles and spacecraft, also supplying drinking water for astronauts).

ⓘ NOTE — ENVIRONMENTAL ADVANTAGE

Although the overall reaction is the same as combustion of hydrogen, no combustion (and hence no direct CO_2/pollutant emissions) actually occurs, since the electrodes keep H_2 and O_2 from directly contacting each other. Water is the only product, making hydrogen fuel cells attractive as a clean energy source (though producing the hydrogen fuel itself may still have an environmental footprint).

CHAPTER 13 QUICK RECAP: Electrolytic cells use electricity to force non-spontaneous redox (electrical → chemical energy); simple/fuel cells use spontaneous redox to generate electricity (chemical → electrical energy). In aqueous electrolysis with inert electrodes: cathode discharges the less reactive cation (H^+ if metal is above H); anode discharges OH^- unless a concentrated halide is present, in which case the halide is discharged (SO_4^{2-}/NO_3^- are never discharged). Reactive anodes dissolve instead (used in copper purification/electroplating). In a simple cell, the more reactive metal is the anode (oxidised, negative electrode) and the less reactive metal is the cathode (site of cation reduction, positive electrode).

Chapter 14: The Periodic Table

Periodic trends, Groups 1, 17, 18, and transition elements

14.1 | Periodic Trends

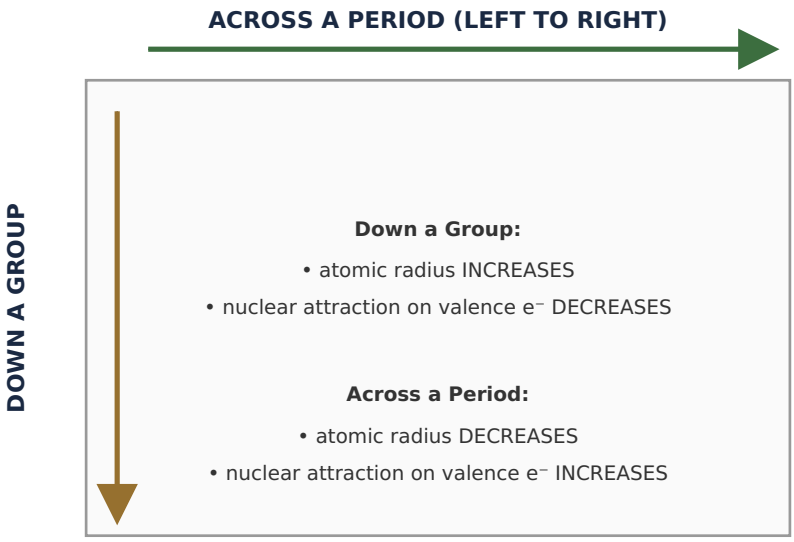


Fig 14.1.1. *Down a group, atomic radius increases and nuclear attraction on valence electrons weakens; across a period, the reverse occurs.*

Table 14.1.1 — Trends down a group vs across a period

	Down a group	Across a period (left to right)
Atomic (proton) number	Increases	Increases
Number of electron shells	Increases	Stays the same
Electrostatic attraction (nucleus–valence e^-)	Decreases	Increases
Atomic radius	Increases	Decreases
Ease of losing electrons (metals)	Increases	Decreases
Ease of gaining electrons / electronegativity (non-metals)	Decreases	Increases

14.2 | Group 1 — The Alkali Metals

Table 14.2.1 — Physical properties of Group 1 metals

Physical properties
Shiny, silvery; soft (easily cut); good electrical/thermal conductors; malleable & ductile; low densities (Li, Na, K float on water); low melting/boiling points (relative to other metals)

> **TREND — M.P./B.P. DECREASES DOWN GROUP 1**
>
> Down the group, more electron shells means valence electrons are further from the nucleus, so the electrostatic attraction between nucleus and delocalised valence electrons weakens. Less thermal energy is needed to overcome these weaker metallic bonds, so m.p./b.p. decreases down the group.

Density increases down Group 1 due to increasing atomic mass (more protons/neutrons per atom).

KEY QNS 14.2.1 — WHY ARE ALKALI METALS HIGHLY REACTIVE?

Group 1 atoms have just 1 valence electron. On reacting, this single electron is lost very readily to form a stable, singly-charged positive ion with noble gas configuration — hence Group 1 metals are highly reactive (and must be stored under oil to prevent reaction with O_2/H_2O in air).

Table 14.2.2 — Chemical properties of alkali metals

Reaction	Equation	Observations
+ water	$2M(s) + 2H_2O(l) \rightarrow 2MOH(aq) + H_2(g)$ (exothermic)	• Li: reacts quickly, floats, no flame • Na: reacts violently, melts into a silvery globule, darts about, catches fire with a yellow flame • K: reacts explosively, melts, catches fire with a lilac flame
+ dilute acid	$2M(s) + 2HCl(aq) \rightarrow 2MCl(aq) + H_2(g)$	Even more vigorous than the reaction with water
+ oxygen	$4M(s) + O_2(g) \rightarrow 2M_2O(s)$	• Li: red flame, white solid • Na: bright yellow flame, white solid • K: violent, lilac flame, white solid
Displacement	Group 1 metals (very low ionisation energy) readily displace less reactive metals from their compounds	

Reactivity increases down Group 1: larger atomic size (more shells) weakens the attraction holding the valence electron, so it is lost more readily by larger atoms.

14.3 | Group 17 — The Halogens

Halogens exist as diatomic, non-metal molecules.

Table 14.3.1 — Physical states and colours at r.t.p.

Element	State @ r.t.p.	Colour
Fluorine	Gas	Pale yellow
Chlorine	Gas	Yellow-green
Bromine	Liquid	Red-brown
Iodine	Solid	Purple-black
Astatine	Solid	Black

EXPLAINING LOW M.P./B.P. AND THE TREND DOWN THE GROUP

Halogens are simple covalent molecules held by weak intermolecular forces — little thermal energy is needed to overcome them, giving low m.p./b.p. Down the group, molecules become larger (more electron shells), increasing intermolecular forces of attraction, so m.p./b.p. **increases** down the group (matching the state change gas → liquid → solid) and colour becomes progressively darker.

Other physical properties: poor electrical conductors; solubility in water decreases down the group.

DEFINITION 14.3.1 — REACTIVITY OF NON-METALS

The tendency of a non-metal atom to gain one or more electrons to form a negative ion. Halogens have a high tendency to gain electrons, making them powerful oxidising agents.

Table 14.3.2 — Halogen displacement reactions

General rule
A more reactive (higher) halogen displaces a less reactive (lower) halogen from its halide salt, since the more reactive halogen has a greater tendency to gain electrons/form negative ions.

> **WORKED EXAMPLE — BROMINE DISPLACING IODIDE**
>
> $Br_2(aq) + 2KI(aq) \rightarrow 2KBr(aq) + I_2(aq)$. Observation: the colourless KI solution turns dark brown (with a black solid deposit, as I_2 is only sparingly soluble in water). Since Br is more reactive than I, Br_2 displaces I^- from solution.

Halogens also react with most metals to form metal halides, and with hydrogen gas to form hydrogen halides.

> **TREND — REACTIVITY DECREASES DOWN GROUP 17**
>
> Larger atomic size (more shells) down the group weakens the nuclear attraction for an incoming electron, so it becomes harder for larger halogen atoms to gain an electron and form a stable negative ion — hence reactivity decreases down the group.

14.4 | Group 18 — The Noble Gases

Monatomic non-metals; low m.p./b.p.; colourless gases at r.t.p.; insoluble in water; chemically unreactive (full valence shell — see Chapter 3).

Table 14.4.1 — Uses of noble gases

Gas	Use
Argon	Fills light bulbs (unreactive with white-hot tungsten filament)
Helium	Fills weather/advertising balloons (low density, unreactive/non-flammable)
Neon	Advertising strip lights ("neon lights")

14.5 | Transition Elements

Found in the central "d-block" of the Periodic Table, between Groups 2 and 13.

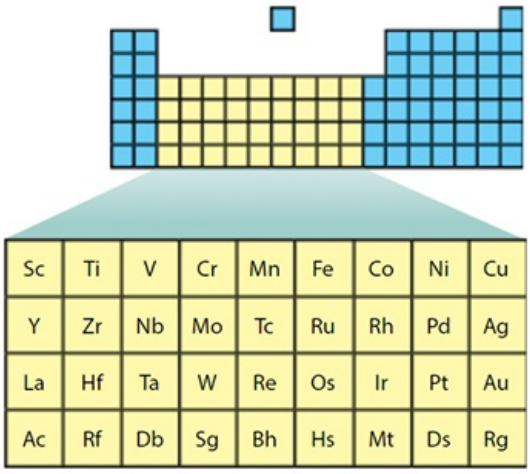


Fig 14.5.1. *Transition elements occupy the central "d-block" of the Periodic Table, between Group 2 and Group 13.*

Table 14.5.1 — Physical properties of transition metals

Property	Explanation
Shiny (lustrous), except copper (red-brown)	—
Good electrical & thermal conductors	Giant metallic lattice structure; delocalised electrons act as mobile charge/heat carriers
Malleable and ductile	Atoms of the same size, orderly arranged in layers that can slide over each other
Harder and stronger than Group 1/2 metals	Transition metal atoms can lose more than 2 valence electrons, forming more highly charged ions with more delocalised electrons — stronger metallic bonds
High density and high melting points	Same reasoning: stronger metallic bonds (from higher ionic charge and more delocalised electrons) require more energy/force to overcome, and atoms are closely packed with high A_r

Table 14.5.2 — Colours of transition metal compounds

Compound	Colour / state @ r.t.p.
Iron(II) hydroxide	Dark-green solid
Aqueous iron(II) compounds	Green solution
Iron(III) hydroxide / oxide	Red-brown solid
Aqueous iron(III) compounds	Yellow solution
Copper(II) carbonate	Green powdery solid
Copper(II) hydroxide	Light blue solid
Copper(II) oxide	Black powdery solid
Other copper(II) compounds	Blue solid/solution
Silver nitrate	White solid / colourless solution
Silver chloride	White solid
Manganese(IV) oxide	Black powdery solid
Potassium manganate(VII)	Purple solid
Acidified potassium manganate(VII)	Purple solution
Manganese(II) compounds	White solid / colourless solution

> ⓘ **NOTE — HYDRATED VS ANHYDROUS COMPOUNDS**
>
> Water of crystallisation can change a compound's colour: anhydrous $CuSO_4$ is white, but $CuSO_4.5H_2O$ is blue; anhydrous $CoCl_2$ is blue, but $CoCl_2.6H_2O$ is pink. (This colour change is used as a test for the presence of water.)

Chemical Properties of Transition Elements

- **Form coloured compounds** (see table above)
- **Form ions with variable oxidation states**, e.g. iron forms Fe^{2+} and Fe^{3+} — this variability lets haemoglobin's iron bind and release oxygen as it circulates the body.
- **Transition elements and their compounds act as good catalysts**: iron catalyses the Haber process; platinum and rhodium (in catalytic converters) remove pollutants from vehicle exhaust; nickel catalyses the manufacture of margarine (hydrogenation of vegetable oils).

CHAPTER 14 QUICK RECAP: Down a group, atomic radius increases and nuclear attraction on valence electrons weakens (metals more reactive, non-metals less reactive); across a period, the reverse happens. Group 1: very reactive, soft, low-density metals, reactivity increases down the group. Group 17: reactive diatomic non-metals, reactivity decreases down the group; more reactive halogens displace less reactive ones. Group 18: unreactive monatomic gases. Transition metals: harder, denser, higher-melting than Group 1/2, form coloured compounds with variable oxidation states, and are useful catalysts.

Chapter 15: The Reactivity Series

Reactions of metals, extraction, and rust prevention

DEFINITION 15.1 — REACTIVITY OF METALS

The tendency of an atom of a metal to lose one or more electrons to form its positive ion.

ⓘ NOTE — REACTIVITY AND REDUCING POWER

The more reactive a metal, the greater its tendency to lose electrons (be oxidised), and so the **stronger a reducing agent** it is. The less reactive a metal, the weaker a reducing agent it is.

15.1 | The Reactivity Series

MNEMONIC: *"Please Stop Calling Me A Careless Zebra Instead Try Learning How Copper Saves Gold"*

Table 15.1.1 — The reactivity series (most to least reactive)

Reactivity series (decreasing reactivity, left to right, top to bottom)			
Potassium (K)	Sodium (Na)	Calcium (Ca)	Magnesium (Mg)
Aluminium (Al)	(Carbon)	Zinc (Zn)	Iron (Fe)
Tin (Sn)	Lead (Pb)	(Hydrogen)	Copper (Cu)
Silver (Ag)	Gold (Au)		

(Carbon and hydrogen are non-metals included as reference points for extraction/displacement.)

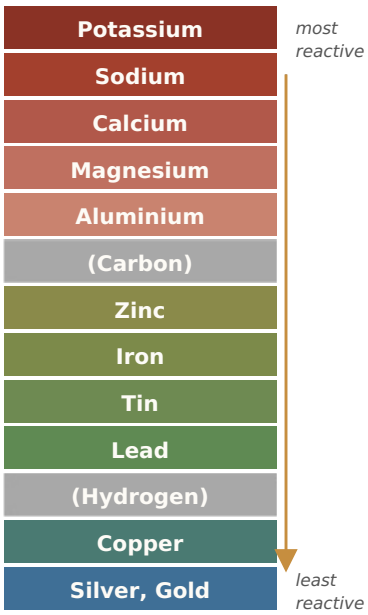


Fig 15.1.1. *The reactivity series, from most reactive (potassium) to least reactive (gold), with carbon and hydrogen shown as non-metal reference points.*

15.2 | Reactions of Metals

Table 15.2.1 — General equations

General equation
metal(s) + water(l) → metal hydroxide(aq) + hydrogen gas(g)
metal(s) + steam(g) → metal oxide(s) + hydrogen gas(g)
metal(s) + dilute acid(aq) → salt(aq) + hydrogen gas(g)

Table 15.2.2 — Observations for individual metals

Metal	+ Water	+ Steam	+ Dilute HCl
Potassium	Explosive reaction; darts on surface; lilac flame	Highly explosive	Highly explosive
Sodium	Violent reaction; darts on surface; orange flame	Highly explosive	Highly explosive
Calcium	Reacts readily; vigorous bubbling	Highly explosive	Reacts violently
Magnesium	Reacts very slowly	Reacts violently; bright white glow	Reacts rapidly; vigorous bubbling
Zinc	No observable change	Reacts readily (ZnO: yellow hot, white cold)	Reacts moderately fast (oxide layer may delay start); effervescence
Iron	No observable change	Reacts slowly → Fe_3O_4 (magnetite, mixed Fe(II)/Fe(III) oxide)	Reacts slowly
Lead, Copper, Silver, Gold, Platinum	No observable change	No observable change	No observable change

> **ⓘ NOTE**
>
> Potassium, sodium and calcium already react so violently with cold water that their reactions with steam or dilute acid are far too dangerous to carry out in a school laboratory — the general reactivity trend simply predicts these would be even more vigorous (highly explosive) than their reaction with water.

Displacement Reactions

> **DEFINITION 15.2.1 — DISPLACEMENT REACTION OF A METAL**
>
> A chemical change where a more reactive metal displaces a less reactive metal from a compound of the less reactive metal.

Table 15.2.3 — General equations for displacement reactions

General equation
more reactive metal(s) + salt of less reactive metal(aq) → less reactive metal(s) + salt of more reactive metal(aq)
more reactive metal(s) + oxide of less reactive metal(s) → less reactive metal(s) + oxide of more reactive metal(s)

> **WORKED EXAMPLE — MG + $CUSO_4$(AQ)**
>
> Magnesium is more reactive than copper (loses electrons more readily), so Mg displaces Cu from blue $CuSO_4$ solution, forming colourless $MgSO_4$ solution and a pink/brown deposit of copper. The blue colour fades as $[Cu^{2+}]$ decreases. Since displacement is exothermic, the mixture warms up.

WORKED EXAMPLE — ZN + CUO (THERMITE-TYPE REACTION)

On heating, zinc displaces copper from copper(II) oxide, forming yellow ZnO (turns white on cooling) and reddish-brown Cu. Highly exothermic, releasing heat and light.

15.3 | Extraction of Metals

Table 15.3.1 — Extraction method by position in the reactivity series

Metal(s)	Extraction method
K, Na, Ca, Mg, Al	Electrolysis of the molten ore (too reactive for chemical reduction)
Zn	Reduction by carbon (heated with carbon/carbon monoxide, e.g. in a blast furnace)
Fe, Sn, Pb, Cu	Reduction by carbon, **or** by hydrogen (heated with carbon/carbon monoxide, e.g. in a blast furnace; or heated in a stream of hydrogen gas)
Silver, Gold	Found naturally uncombined (unreactive)

Table 15.3.2 — Reduction equations

Equation
metal oxide(s) → metal(s) + oxygen(g) — via electrolysis
metal oxide(s) + carbon(s) → metal(s) + carbon dioxide(g)
metal oxide(s) + hydrogen(g) → metal(s) + steam(g) — possible for Fe, Sn, Pb (and Cu) oxides

Thermal Decomposition of Metal Carbonates

$$\text{metal carbonate(s)} \rightarrow \text{metal oxide(s)} + \text{carbon dioxide(g)}$$

ⓘ NOTE

Potassium and sodium carbonates do **not** decompose on heating (too stable/thermally stable). As reactivity of the metal decreases, the thermal stability of its carbonate decreases, so ease of thermal decomposition **increases** down the reactivity series. Silver carbonate decomposes unusually ($2Ag_2CO_3(s) \rightarrow 4Ag(s) + 2CO_2(g) + O_2(g)$) — a redox reaction, unlike the others (which are not redox).

15.4 | Rusting and Its Prevention

DEFINITION 15.4.1 — RUSTING

The slow oxidation of iron to form hydrated iron(III) oxide.

$$4Fe(s) + 3O_2(g) + 2xH_2O(l) \rightarrow 2Fe_2O_3.xH_2O(s)$$

Conditions needed for rusting: water AND oxygen (both must be present). **Factors that speed up rusting:** presence of salt (electrolyte, increases conductivity); presence of acidic substances.

Table 15.4.1 — Methods of rust prevention

Method	Description
Barrier methods	Form a protective coating around iron/steel to exclude water and oxygen: painting, oiling/greasing, plastic coating, galvanising (zinc-plating), tin-plating, chromium-plating
Sacrificial protection	A more reactive metal (e.g. magnesium, zinc) is attached to iron/steel; it corrodes ("sacrifices itself") in place of the iron

KEY QNS 15.4.1 — HOW DOES ZINC PREVENT IRON FROM RUSTING?

Zinc is more reactive than iron and has a greater tendency to lose electrons to form positive ions. Zinc atoms therefore lose electrons (react with oxygen/water) in preference to iron atoms, corroding in place of iron and protecting it from rusting.

ⓘ NOTE — GALVANISING VS SIMPLE ZINC COATING

Galvanising is special because even if the zinc coating is scratched (exposing the iron beneath), the iron is still protected via sacrificial protection, since zinc continues to react preferentially. A simple barrier coating (e.g. paint) offers no protection once scratched/broken.

CHAPTER 15 QUICK RECAP: Reactivity = tendency to lose electrons; more reactive metals react more vigorously with water/steam/acid, and can displace less reactive metals from their compounds. Extraction method depends on reactivity: electrolysis (most reactive), carbon/hydrogen reduction (moderately reactive), or found native (least reactive). Rusting needs both water and oxygen; prevented by barrier methods or sacrificial protection using a more reactive metal.

Chapter 16: Chemical Energetics

Exothermic/endothermic changes, bond energy, and enthalpy

16.1 | Exothermic and Endothermic Changes

DEFINITION 16.1.1 — EXOTHERMIC CHANGE

A physical or chemical change that transfers (releases) energy to the surroundings.

Table 16.1.1 — Examples of exothermic changes

Physical	Dissolving alkalis/sodium carbonate in water; dissolving concentrated acids in water; condensation; freezing; deposition
Chemical (reactions)	Neutralisation; reaction of reactive metals with water/steam/acid; displacement of metals from solution (incl. thermite reactions); corrosion of metals; combustion; respiration

KEY QNS 16.1.1 — TEMPERATURE CHANGE DURING AN EXOTHERMIC CHANGE

Q. Explain the temperature rise observed when NaOH pellets dissolve in water.
A. As the exothermic change occurs, the system releases thermal energy to the surroundings. Since the thermometer/solution is part of the surroundings, temperature rises (e.g. from 28.0°C to a maximum of 34.0°C) until dissolving is complete.

KEY QNS 16.1.2 — WHY DOES TEMPERATURE FALL BACK TO ROOM TEMPERATURE AFTERWARD?

Once the NaOH has completely dissolved, no more thermal energy is released to the surroundings. The (now-warmer) solution loses heat to the (cooler) surroundings over time, so temperature gradually falls back to room temperature.

KEY QNS 16.1.3 — WHY IS MAXIMUM TEMPERATURE REACHED WHEN THE REACTION IS COMPLETE?

The greater the number of moles reacted, the greater the total energy released, so the higher the temperature rise. Once the reaction is complete, the maximum possible number of moles has reacted, so the maximum possible amount of energy has been released — giving the maximum temperature.

DEFINITION 16.1.2 — ENDOTHERMIC CHANGE

A physical or chemical change that takes in (absorbs) energy from the surroundings.

Table 16.1.2 — Examples of endothermic changes

Physical	Dissolving certain salts (e.g. NH_4Cl, $NaNO_3$) in water; evaporation; boiling; sublimation
Chemical (reactions)	Electrolysis of molten ore; thermal decomposition of metal carbonates; photosynthesis; formation of nitrogen oxides (e.g. in a car engine or lightning: $N_2(g)+O_2(g) \rightarrow 2NO(g)$); decomposition of CFCs by sunlight (e.g. $CCl_2F_2 \rightarrow CClF_2 + Cl$)

16.2 | Bond Energy, Activation Energy and Enthalpy Change

DEFINITIONS

Bond energy — the energy absorbed to break one mole of a particular chemical bond (equal to the energy released when one mole of that bond forms).
Activation energy — the minimum energy that colliding reacting particles must possess in order to react (undergo an effective collision).
Enthalpy change (ΔH) — the difference in energy between products and reactants.

$$\Delta H = E_{products} - E_{reactants}$$

A negative ΔH ($E_{products} < E_{reactants}$) indicates an **exothermic** change; a positive ΔH ($E_{products} > E_{reactants}$) indicates an **endothermic** change.

Table 16.2.1 — Exothermic vs endothermic reactions

	Exothermic	Endothermic
Heat transfer	Given out to surroundings	Taken in from surroundings
Energy of products vs reactants	Products lower than reactants	Products higher than reactants
Sign of ΔH	Negative	Positive
Temperature of surroundings	Increases	Decreases

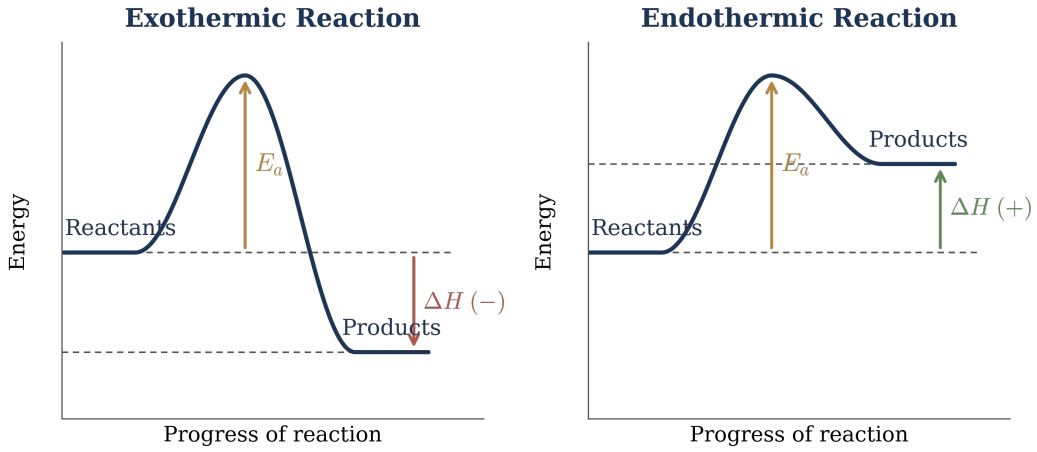


Fig 16.2.1. *Energy profile diagrams: the peak above the reactants marks the activation energy (E_a); the vertical gap between the reactants' and products' final energy levels is ΔH.*

> ⓘ **NOTE — ENERGY LEVEL / ENERGY PROFILE DIAGRAMS**
>
> An **energy level diagram** shows reactants and products as horizontal lines at their respective energy levels, with ΔH marked as the vertical gap between them (below the reactants' level for exothermic; above for endothermic).
> An **energy profile diagram** additionally shows the reaction pathway as a curve rising from reactants to a peak (the activation energy, E_a) before falling to products — for exothermic reactions the products' curve ends below the reactants' level; for endothermic reactions it ends above.
> When comparing two energy diagrams, always address: (i) direction of heat transfer, (ii) relative energy content of products vs reactants, (iii) relative magnitude of the energy change, and (iv) resulting change in surroundings' temperature.

Calculating Enthalpy Change from Bond Energies

$$\Delta H = \Delta H_{bonds\ broken} + \Delta H_{bonds\ formed}$$

(Bond breaking always requires energy input, so $\Delta H_{bonds\ broken}$ is positive; bond formation always releases energy, so $\Delta H_{bonds\ formed}$ is negative.)

> **WORKED EXAMPLE — H_2 + BR_2 → 2HBR (EXOTHERMIC)**
>
> **Q.** Use bond breaking/making to explain why this reaction is exothermic.
> **A.** Bond breaking is endothermic; bond making is exothermic. The energy absorbed to break the bonds in 1 mol H_2 and 1 mol Br_2 is *less* than the energy released when forming the bonds in 2 mol HBr. Since more energy is released than absorbed, the overall reaction is exothermic (ΔH is negative), meaning the energy level of the products (HBr) is lower than that of the reactants (H_2 and Br_2).

CHAPTER 16 QUICK RECAP: Exothermic changes release heat (ΔH negative, temperature of surroundings rises); endothermic changes absorb heat (ΔH positive, temperature falls). $\mathbf{\Delta H = E_{products} - E_{reactants}}$ = energy needed to break bonds + energy released forming new bonds. A reaction is exothermic overall when more energy is released forming new bonds than was needed to break the old ones.

Chapter 17: Rate of Reaction

Collision theory and the factors affecting reaction rate

17.1 | Measuring Rate of Reaction

Table 17.1.1 — Common methods to measure rate of reaction

Method
Measure the volume of gas produced at regular time intervals (e.g. using a gas syringe)
Measure the change in mass (e.g. of the reaction flask, if a gas escapes) at regular time intervals
Measure the time taken for a fixed, observable change to occur (e.g. a precipitate to obscure a mark, or a colour change)

$$\text{Average rate of reaction} = \text{Amount of reactant used (or product formed)} \div \text{Time taken}$$

17.2 | Collision Theory

DEFINITION 17.2.1 — ACTIVATION ENERGY

The minimum amount of energy that colliding reacting particles must possess for an effective collision (one that results in a reaction) to occur.

For a reaction to occur, particles must (i) collide, and (ii) collide with energy ≥ the activation energy, and (iii) with the correct orientation. Only such "effective collisions" result in reaction. Increasing the **frequency of effective collisions** increases the rate of reaction.

17.3 | Factors Affecting Rate of Reaction

Table 17.3.1 — Factors and their explanations

Factor	Effect	Explanation
Concentration (solutions)	Rate increases with concentration	More reacting particles per unit volume → more collisions per unit time → more frequent effective collisions → faster rate
Pressure (gases)	Rate increases with pressure	Particles are pushed closer together, increasing particles per unit volume → more frequent effective collisions → faster rate
Particle size (solids)	Rate increases as particle size decreases	Smaller particles → greater total exposed surface area for other reactants to collide with → more frequent effective collisions → faster rate
Temperature	Rate increases with temperature	Particles gain kinetic energy, move faster, and collide more often per unit time; more particles possess energy ≥ activation energy, so the *frequency* of effective collisions increases sharply → faster rate
Catalyst	Rate increases with a catalyst present	Provides an alternative reaction pathway with a *lower* activation energy, so more particles have sufficient energy to react → more frequent effective collisions → faster rate

ⓘ NOTE — MEMORY AID

For every factor: identify what increases (concentration/pressure/temperature/particle collisions) → state its effect on collision frequency and/or the proportion of particles meeting the activation energy → conclude "more frequent effective collisions → faster rate of reaction, as more product is formed per unit time."

17.4 | Catalysts

DEFINITION 17.4.1 — CATALYST

A substance that increases the rate of a reaction by providing an alternative reaction pathway with a lower activation energy, while remaining chemically unchanged at the end of the reaction (though it may take part in intermediate steps during the reaction).

ⓘ NOTE — ENZYMES AS BIOLOGICAL CATALYSTS

Enzymes are proteins that act as **biological catalysts**, speeding up biochemical reactions in living organisms (e.g. the enzyme zymase in yeast catalyses the fermentation of glucose into ethanol — see Chapter 20). Like other catalysts, enzymes lower the activation energy of a reaction and remain unchanged at the end. However, enzymes are highly sensitive to temperature and pH: at too high a temperature (typically above ~40°C), an enzyme is **denatured** (its structure is permanently altered) and permanently loses its catalytic activity, unlike most inorganic catalysts.

Table 17.4.1 — Characteristics of catalysts

Characteristic
Provides an alternative pathway with lower activation energy
Does not affect the enthalpy change (ΔH) of the reaction
Does not affect the yield of the reaction (only the rate at which equilibrium/completion is reached)
Remains chemically unchanged in mass and composition at the end of the reaction
Highly selective in action (specific to particular reactions)
Industrial catalysts are typically transition metals or their compounds (Chapter 14)
Can be "poisoned" (deactivated) by impurities

COMMON EXAM PITFALL

A catalyst changes the *rate* of reaction, never the *yield* or the *enthalpy change*. Do not confuse "increases rate of reaction" with "increases amount of product formed" — a catalyst gets you to the same final amount of product faster, it does not create more product.

CHAPTER 17 QUICK RECAP: Rate of reaction increases with concentration, pressure, temperature, and decreasing particle size, and with a catalyst present — all essentially by increasing the frequency of effective collisions (temperature also increases the proportion of particles with energy ≥ activation energy; catalysts lower the activation energy itself). Catalysts are unchanged in mass/composition at the end, do not affect yield or ΔH, and are highly selective.

Chapter 18: Fuels and Crude Oil

Fossil fuels, fractional distillation of crude oil, and biofuels

18.1 | Organic Compounds, Hydrocarbons and Fossil Fuels

DEFINITIONS 18.1.1

Organic compound — a pure substance containing carbon atoms covalently bonded to atoms of other elements, usually hydrogen, oxygen or chlorine (e.g. methane, ethanoic acid, ethanol).
Hydrocarbon — an organic compound made up of carbon and hydrogen only (e.g. methane, CH_4).
Fossil fuels — mixtures consisting mainly of hydrocarbons, formed from the decayed remains of plants and animals compressed over millions of years. Coal, crude oil (petroleum) and natural gas are the three common fossil fuels.

> **ⓘ NOTE — WHY FOSSIL FUELS ARE NON-RENEWABLE**
>
> Fossil fuels form over millions of years, but are consumed by human activity far faster than they can be replenished — hence they are considered finite, non-renewable resources.

Natural Gas

A colourless, odourless gaseous mixture consisting mainly of methane (~70–90%) and other short-chain alkanes (ethane, propane, butane).

$$\mathrm{CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g) \quad (complete\ combustion)}$$

> **ⓘ NOTE — WHY NATURAL GAS IS THE "CLEANEST" FOSSIL FUEL**
>
> Natural gas has the lowest percentage by mass of carbon of the fossil fuels, and contains negligible sulfur, so its combustion produces the least CO_2 and no SO_2, compared to coal or crude oil.

Natural gas is liquefied (as LNG) before transport/storage: liquefying reduces the volume needed for the same mass by roughly 600 times (particles in a liquid are packed far more closely than in a gas), and LNG is also safer to handle (less flammable than the pressurised gas). A pungent-smelling additive (e.g. tetrahydrothiophene, THT) is added, since natural gas itself is odourless, so leaks can be detected by smell.

Crude Oil (Petroleum)

A dark, sticky liquid at r.t.p.; a mixture of many different hydrocarbons with 1 to 70 carbon atoms per molecule, and therefore has limited direct uses on its own — it must be separated by fractional distillation into useful fractions.

> **ⓘ NOTE — IMPORTANT DISTINCTIONS**
>
> Crude oil is a **mixture**, not a compound. Petroleum (crude oil) ≠ petrol (petrol is one fraction of petroleum). Petroleum gas ≠ natural gas (petroleum gas is one fraction of petroleum).

Table 18.1.1 — Comparing crude oil and natural gas as fuels

	Crude oil	Natural gas
State @ r.t.p.	Liquid	Gas
Uses	Fractionally distilled into many fractions used as fuels/feedstock	Cooking fuel (LNG), vehicle fuel (CNG)
Advantages	Easily transported/stored in tanks; economical, convenient	Cleanest fossil fuel (no soot/SO_2); lower CO_2 per unit energy; burns efficiently; economical
Disadvantages	Higher carbon content and sulfur impurities cause incomplete combustion (soot, CO, SO_2); CO_2 emissions contribute to global warming	Still produces CO_2 on combustion, contributing to global warming

$$\text{crude oil (limiting } O_2\text{)} + \text{oxygen} \rightarrow$$
$$\text{carbon (soot)} + \text{carbon monoxide} + \text{carbon dioxide} + \text{steam} + \text{sulfur dioxide}$$

18.2 | Fractional Distillation of Crude Oil

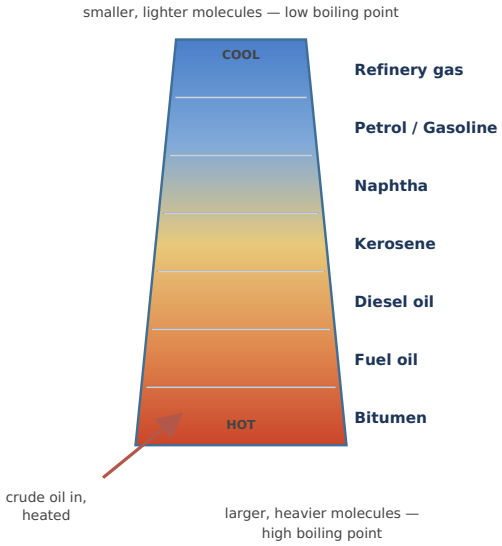


Fig 18.2.1. *Fractional distillation column: crude oil vapour rises, and fractions condense out at different heights according to their boiling point range — lightest/smallest hydrocarbons at the top, heaviest/largest at the bottom.*

Crude oil is separated by fractional distillation, exploiting the different boiling points of its many hydrocarbon components (Chapter 1). The oil is heated in a furnace, vaporises, and hot vapour rises up a fractionating column, which is hotter at the bottom and progressively cooler towards the top.

> **EXPLAINING THE SEPARATION PROCESS**
>
> 1. Crude oil is heated in a furnace; boiling takes place and hot vapour enters the fractionating column.
> 2. Heavier fractions (larger hydrocarbons, more carbon atoms) have higher boiling point ranges, so they condense at the higher temperatures found lower down the column, and are collected there.
> 3. Lighter fractions (smaller hydrocarbons, fewer carbon atoms) have lower boiling point ranges, so they rise further up the (cooler) column before condensing, and are collected at higher levels.

Table 18.2.1 — Properties of lighter vs heavier fractions

	Lighter fractions (top of column)	Heavier fractions (bottom of column)
Carbon atoms per molecule	Fewer (as few as 1)	More (up to ~70)
Molecule size	Smaller	Larger
Boiling point range	Low	High
Viscosity	Low (flows easily)	High (flows poorly; solids or highly viscous liquids)
Volatility	Volatile	Non-volatile
Ease of ignition	Ignites easily	Does not ignite easily
Uses	Fuels for cooking, heating, motor vehicles, electricity generation; petrochemical feedstock (detergents, medicines, fertilisers, pesticides, synthetic rubber)	Lubricating oil, road surfacing (bitumen/asphalt)
Demand	High (supply often cannot meet demand)	Low

ⓘ NOTE — MEETING DEMAND: CRACKING

Since demand for lighter fractions exceeds supply, while heavier fractions are in low demand, oil refineries use **catalytic cracking** to break down excess heavy fractions into smaller, more useful lighter hydrocarbons (see Chapter 19).

18.3 | Biofuels

DEFINITION 18.3.1 — BIOFUEL

A fuel derived from plant or animal matter, offering a renewable alternative to fossil fuels. Examples: bioethanol (from sugarcane/corn fermentation), biodiesel (vegetable oils/animal fats), green diesel (algae), biogas (methane from decomposing organic waste).

DEFINITION 18.3.2 — ENVIRONMENTALLY SUSTAINABLE

Able to be maintained at a steady level without exhausting natural resources or causing ecological damage. A sustainable energy source is renewable, causes little/no pollution, and is carbon neutral (does not contribute to net CO_2 emissions over its life cycle).

WORKED EXAMPLE — WHY IS BIOETHANOL MORE ENVIRONMENTALLY SUSTAINABLE?

Photosynthesis (as sugarcane grows): $6H_2O(l) + 6CO_2(g) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$

Combustion (when bioethanol burns): $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$

As sugarcane plants can be regrown and replanted in a short period of time, the ethanol generated from sugarcane is considered a **renewable** source, unlike fossil fuels which take millions of years to form.

As sugarcane grows, it absorbs CO_2 from the atmosphere during photosynthesis, which offsets the CO_2 released when the bioethanol derived from it is later burned — giving a lower *net* CO_2 emission than burning fossil fuels (which release CO_2 that was locked away for millions of years).

COMMON EXAM PITFALL — BIOETHANOL IS NOT 100% CARBON NEUTRAL

Producing and transporting biofuels (planting, harvesting, processing sugarcane) itself requires energy, often from burning fossil fuels, which releases additional CO_2 not offset by the crop's own photosynthesis. Other drawbacks: clearing land (sometimes by burning forests) reduces biodiversity and releases CO_2; growing fuel crops competes with food production; fertiliser use can pollute waterways.

Table 18.3.1 — Key considerations for a good alternative energy source

Consideration
Renewability — replenished at least as fast as it is consumed
Ease of transport and storage
Environmental sustainability — minimal pollution, carbon neutral over its life cycle
Energy efficiency — releases a large amount of energy per unit of fuel
Safety — easily ignited, but safe to handle

CHAPTER 18 QUICK RECAP: Fossil fuels (coal, crude oil, natural gas) are non-renewable hydrocarbon mixtures. Crude oil is separated by fractional distillation into fractions of similar chain length: lighter fractions (fewer C atoms, low b.p., low viscosity, volatile, ignite easily, high demand) collect near the top; heavier fractions (more C atoms, high b.p., viscous, non-volatile) collect near the bottom. Biofuels (e.g. bioethanol) are renewable and lower in net CO_2 emissions than fossil fuels, but production/transport still carries an environmental cost, so they are not fully carbon neutral.

Chapter 19: Hydrocarbons

Homologous series, alkanes, alkenes, isomerism and cracking

19.1 | Homologous Series

DEFINITION 19.1.1 — HOMOLOGOUS SERIES

A family of organic compounds with the same general formula and similar chemical properties, because they share the same functional group.

DEFINITION 19.1.2 — FUNCTIONAL GROUP

An atom or group of atoms that gives an organic molecule its characteristic chemical properties/reactions.

Table 19.1.1 — General characteristics of a homologous series

Characteristic
Members share the same general formula
Members share the same functional group, and hence similar chemical properties
Each member differs from the next by an additional $-CH_2-$ unit (so M_r increases by 14 between consecutive members)
Physical properties (m.p., b.p., density, viscosity) show a gradual trend down the series

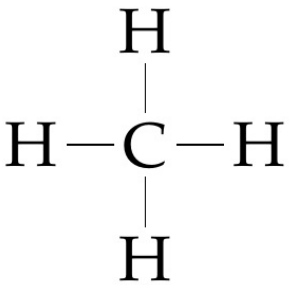
Table 19.1.2 — Key homologous series

Series	General formula	Functional group	Suffix / prefix example
Alkanes	C_nH_{2n+2}	None (saturated)	–ane (e.g. methane, CH_4)
Alkenes	C_nH_{2n}	C=C double bond	–ene (e.g. propene, C_3H_6)
Alcohols	$C_nH_{2n+1}OH$	Hydroxyl, –OH	–ol (e.g. ethanol, C_2H_5OH)
Carboxylic acids	$C_nH_{2n+1}COOH$	Carboxyl, –COOH	–oic acid (e.g. butanoic acid)
Esters	—	Ester linkage, –COO–	–anoate (Chapter 20)

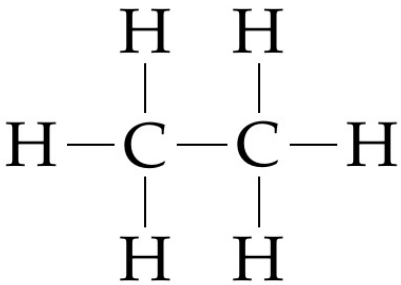
ⓘ NOTE — NAMING ORGANIC COMPOUNDS

Name = (prefix: number of carbon atoms) + (suffix: homologous series).
Prefixes: meth-(1), eth-(2), prop-(3), but-(4), pent-(5), hex-(6), hept-(7), oct-(8).
Suffixes: alkanes –ane, alkenes –ene, alcohols –ol, carboxylic acids –oic acid, esters –anoate.

19.2 | Alkanes



Methane, CH_4



Ethane, C_2H_6

Fig 19.1.1. *Displayed (full structural) formulae of methane and ethane, showing only single C–C and C–H bonds (saturated).*

DEFINITION 19.2.1 — ALKANES

A homologous series of **saturated** hydrocarbons with general formula C_nH_{2n+2}, containing only C–C and C–H single bonds.

Table 19.2.1 — First five alkanes

n	Name	Formula	M_r	B.p. / °C	State @ r.t.p.
1	Methane	CH_4	16	−162	Gas
2	Ethane	C_2H_6	30	−89	Gas
3	Propane	C_3H_8	44	−42	Gas
4	Butane	C_4H_{10}	58	−0.5	Gas
5	Pentane	C_5H_{12}	72	36	Liquid

Branched vs Straight-Chain Alkanes

A **straight-chain (unbranched)** alkane has all its carbon atoms joined in a single row (e.g. butane, $CH_3CH_2CH_2CH_3$). A **branched** alkane has at least one carbon atom outside the main chain, forming a side chain (e.g. 2-methylpropane, $CH_3CH(CH_3)CH_3$, has a 3-carbon main chain with a $-CH_3$ branch).

ⓘ NOTE — EFFECT OF BRANCHING ON BOILING POINT

For alkanes with the same number of carbon atoms, **increased branching lowers the boiling point** (and density) — e.g. among the C_5H_{12} isomers, unbranched pentane (36°C) > 2-methylbutane (one branch, 28°C) > 2,2-dimethylpropane (two branches, 9°C). Branched molecules pack together less efficiently and have a smaller surface area of contact, so intermolecular forces are weaker.

Physical Properties and Trends

Alkanes are simple covalent molecules: low m.p./b.p.; insoluble in water but soluble in organic solvents; poor conductors of electricity/heat.

Table 19.2.2 — Trends down the alkane series (increasing carbon number)

Property	Trend	Reason
Melting/boiling point	Increases	Larger molecules → greater surface area of contact → stronger intermolecular forces → more energy needed to overcome them
Viscosity	Increases	Same reasoning: stronger intermolecular forces between larger molecules make the liquid resist flow more
Solubility in organic solvents	Decreases	—
Density	Increases	Larger M_r and (for liquids) more closely packed molecules per unit volume

WORKED EXAMPLE — WHY DOES BUTANE HAVE A HIGHER DENSITY THAN ETHANE?

Both butane (M_r=58) and ethane (M_r=30) are gases at r.t.p. Butane molecules have a greater molecular mass than ethane molecules. Thus, in a fixed unit volume of gas, there is a greater mass of butane present compared to the same unit volume of ethane. Hence, butane has a higher density than ethane.

Chemical Properties of Alkanes

Alkanes are generally unreactive (all bonds are strong C–C/C–H single covalent bonds), except for combustion and substitution reactions.

Reactions of Alkanes (exemplified by methane)

Table 19.2.3 — Reactions of alkanes

Reaction	Equation	Conditions
Complete combustion	$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$	Excess oxygen
Incomplete combustion	$2CH_4(g) + 3O_2(g) \rightarrow 2CO(g) + 4H_2O(l)$	Limited oxygen
Incomplete combustion (sooting)	$CH_4(g) + O_2(g) \rightarrow C(s) + 2H_2O(l)$	Very limited oxygen
Substitution by chlorine (step 1)	$CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g)$	UV light/sunlight
Substitution by chlorine (step 2, excess Cl_2)	$CH_3Cl(g) + Cl_2(g) \rightarrow CH_2Cl_2(g) + HCl(g)$	UV light/sunlight

Further substitution can continue ($CH_2Cl_2 \rightarrow CHCl_3 \rightarrow CCl_4$) with excess chlorine, giving a mixture of chlorinated products — one H atom is replaced by one Cl atom at each step.

19.3 | Alkenes

DEFINITION 19.3.1 — ALKENES

A homologous series of **unsaturated** hydrocarbons with general formula C_nH_{2n}, containing at least one carbon–carbon double bond (C=C).

Table 19.3.1 — First few alkenes

n	Name	Formula
2	Ethene	C_2H_4
3	Propene	C_3H_6
4	Butene	C_4H_8

Testing for Saturation — Bromine Water

Table 19.3.2 — Saturated vs unsaturated hydrocarbons

	Saturated (alkanes)	Unsaturated (alkenes)
Bonding	Only C–C and C–H single bonds	Contains at least one C=C double bond
+ aqueous bromine (orange/brown)	No reaction; bromine water stays orange/brown	Bromine water is decolourised (orange/brown → colourless)

WORKED EXAMPLE — ADDITION OF BROMINE TO ETHENE

$C_2H_4(g) + Br_2(aq) \rightarrow C_2H_4Br_2(l)$ — the C=C double bond opens up and each carbon bonds to one Br atom (an **addition reaction**), decolourising the bromine water. This is the standard chemical test to distinguish alkenes from alkanes.

Reactions of Alkenes (exemplified by ethene)

Table 19.3.3 — Reactions of alkenes

Reaction	Equation	Conditions
Combustion	$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$	Excess oxygen
Addition of bromine	$C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$	r.t.p.
Addition of steam (hydration)	$C_2H_4(g) + H_2O(g) \rightarrow C_2H_5OH(g)$	Catalyst (phosphoric acid), ~300°C, 60–70 atm
Addition of hydrogen (hydrogenation)	$C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$	Nickel catalyst, 150°C
Polymerisation	$n\ C_2H_4 \rightarrow -(CH_2CH_2)_n-$ (poly(ethene))	High pressure, catalyst (Chapter 21)

ⓘ NOTE — POLYUNSATURATED FATS AND MARGARINE

"Polyunsaturated" describes a fat/oil molecule containing **more than one** C=C double bond. Vegetable oils (liquid, unsaturated) are converted to margarine (solid) by **hydrogenation**: adding hydrogen gas across the C=C double bonds (nickel catalyst), which reduces the number of double bonds and raises the melting point above r.t.p., producing a solid/semi-solid product.

19.4 | Isomerism

DEFINITION 19.4.1 — ISOMERS

Compounds with the same molecular formula but different structural formulae (different arrangement of atoms). Isomers can differ in: straight-chain vs branched-chain structure; type/position of functional group; type/number of bonds; or position of a substituent/side-chain.

WORKED EXAMPLE — BUTANE AND 2-METHYLPROPANE

Both have molecular formula C_4H_{10} (same type and number of atoms), but different structural formulae: butane, $CH_3CH_2CH_2CH_3$, is a straight-chain alkane (all 4 C in one row); 2-methylpropane, $CH_3CH(CH_3)CH_3$, is a branched alkane (main chain of 3 C, with a $-CH_3$ side chain). Since they share a molecular formula but differ in structure, they are isomers.

19.5 | Cracking

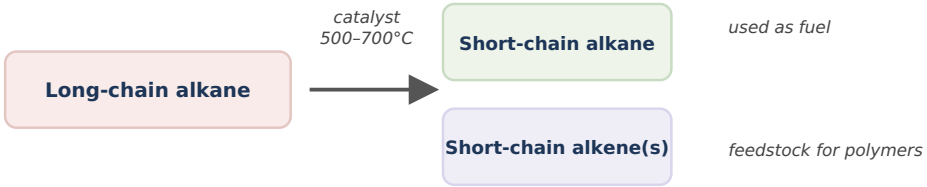


Fig 19.5.1. *Catalytic cracking breaks a long-chain (heavy, low-demand) alkane into a shorter alkane and alkene(s), matching supply to demand for lighter fractions.*

DEFINITION 19.5.1 — CRACKING

The breakdown of large, heavy (long-chain) hydrocarbon molecules into smaller, more useful molecules (shorter-chain alkanes and alkenes), by breaking C–C bonds.

long-chain alkane → shorter-chain alkane + shorter-chain alkene(s)
(+ hydrogen gas, in some cases)

Table 19.5.1 — Conditions for (catalytic) cracking

Condition
Catalyst: aluminium oxide (Al_2O_3) and silicon(IV) oxide (SiO_2)
High temperature: 500–700°C
Pressure: ~1 atm

> **ⓘ NOTE — WHY CRACKING MATTERS**
>
> Fractional distillation of crude oil yields far more heavy (low-demand) fractions than the market needs, while demand for light fractions (e.g. petrol, used as fuel) exceeds supply. Cracking heavy fractions into lighter alkanes (fuel) and alkenes (feedstock for plastics/polymers, Chapter 21) helps match supply to demand. Cracking is also used to break down waste plastics (e.g. poly(ethene)) into useful short-chain alkanes/alkenes as a chemical recycling method.

CHAPTER 19 QUICK RECAP: A homologous series shares a general formula, functional group and trend in physical properties. Alkanes (C_nH_{2n+2}, saturated, unreactive except combustion/substitution) vs alkenes (C_nH_{2n}, unsaturated, reactive via addition reactions with Br_2/steam/H_2, decolourise bromine water). Isomers share a molecular formula but differ in structure. Cracking breaks long-chain hydrocarbons into shorter, more useful alkanes/alkenes using a catalyst, heat, and moderate pressure.

Chapter 20: Alcohols, Carboxylic Acids and Esters

Fermentation, oxidation, and esterification

20.1 | Alcohols

DEFINITION 20.1.1 — ALCOHOLS

A homologous series with general formula $C_nH_{2n+1}OH$, containing the hydroxyl functional group, –OH.

First members: methanol (CH_3OH), ethanol (C_2H_5OH), propanol (C_3H_7OH), butanol (C_4H_9OH). Ethanol is the most important industrially and biologically.

COMBUSTION OF ALCOHOLS

Like other flammable organic compounds, alcohols undergo complete combustion in excess oxygen to form carbon dioxide and water, releasing energy:

$$\mathbf{C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)}$$

This reaction is exothermic, which is why alcohols (e.g. ethanol, methanol) are commonly used as fuels.

Manufacture of Ethanol by Fermentation

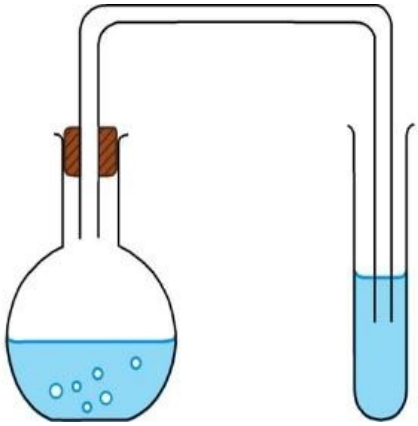


Fig 20.1.1. *Fermentation set-up: glucose solution with yeast connected to limewater. White ppt. forms as CO_2 is produced.*

DEFINITION 20.1.2 — FERMENTATION

A chemical process in which microorganisms (yeast) act on carbohydrates like glucose, in the absence of oxygen, to produce ethanol and carbon dioxide.

$$\mathbf{C_6H_{12}O_6(aq) \xrightarrow{\text{yeast}} 2C_2H_5OH(aq) + 2CO_2(g)}$$

Table 20.1.1 — Conditions for fermentation

Condition
Anaerobic (no oxygen) environment — apparatus must be airtight
Presence of yeast (containing the enzyme zymase, which catalyses the breakdown of glucose)
Optimum temperature ~18–37°C (yeast is inactive below 18°C and denatured above ~40°C)

ⓘ NOTE

Limewater is used to confirm CO_2 production (formation of white ppt.). Fermentation alone produces only dilute ethanol (up to ~20%), since higher ethanol concentrations denature the yeast and halt the process. **Fractional distillation** is then used to concentrate the ethanol, since ethanol and water are miscible liquids with different boiling points.

COMMON EXAM PITFALL

If oxygen is allowed into the fermentation vessel, the ethanol formed is oxidised by atmospheric oxygen to ethanoic acid (vinegar), turning the mixture sour: $C_2H_5OH + O_2 \rightarrow CH_3COOH + H_2O$. This is why fermentation vessels/wine bottles must be kept airtight/corked.

Oxidation of Ethanol

Ethanol can be oxidised to ethanoic acid in two main ways: slowly by atmospheric oxygen (as described above), or more rapidly in the laboratory using an oxidising agent such as acidified potassium manganate(VII), $KMnO_4(aq)$, or acidified potassium dichromate(VI).

EQUATION 1 — SLOW ATMOSPHERIC OXIDATION

Occurs whenever ethanol is left exposed to air (e.g. an uncorked bottle of wine turning sour):

$$C_2H_5OH(aq) + O_2(g) \rightarrow CH_3COOH(aq) + H_2O(l)$$

EQUATION 2 — RAPID CATALYTIC/LABORATORY OXIDATION

Ethanol is heated **under reflux** with acidified potassium manganate(VII). By convention, the oxygen supplied by the oxidising agent is written as **2[O]** (oxygen atoms from the oxidising agent), not O_2:

$$C_2H_5OH(aq) + 2[O] \rightarrow CH_3COOH(aq) + H_2O(l) \quad (\text{heat, under reflux})$$

Ethanol gains oxygen from acidified $KMnO_4(aq)$ — ethanol is **oxidised** (and is the reducing agent), while acidified $KMnO_4(aq)$ is **reduced** (and is the oxidising agent): this is a redox reaction. The purple acidified $KMnO_4(aq)$ is decolourised as Mn is reduced from +7 (in MnO_4^-) to +2 (in Mn^{2+}).

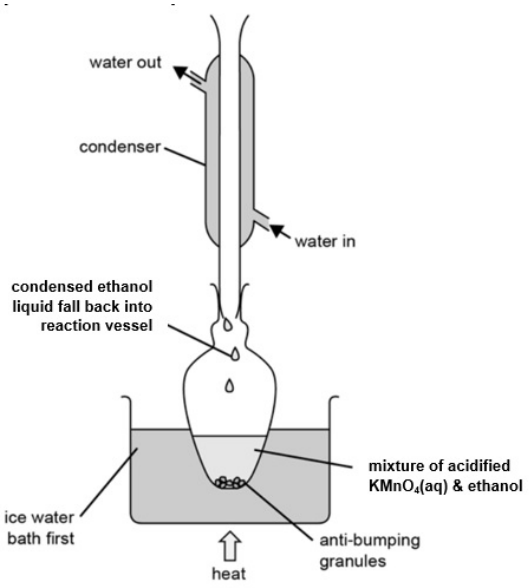


Fig 20.1.2. *Experimental set-up for the oxidation of ethanol to ethanoic acid under reflux.*

ⓘ PURPOSE OF THE CONDENSER

Ethanol is a highly volatile liquid. The condenser cools and condenses ethanol vapour so that it is not lost from the reaction vessel — the condensed ethanol liquid falls back into the flask and continues to be oxidised to carboxylic acid.

ⓘ WHY USE A WATER BATH INSTEAD OF DIRECT HEATING?

Ethanol is highly flammable. To prevent an explosion or fire, the reaction mixture is heated gently using a water bath rather than being heated directly with a Bunsen flame.

20.2 | Carboxylic Acids

DEFINITION 20.2.1 — CARBOXYLIC ACIDS

A homologous series with general formula $C_nH_{2n+1}COOH$, containing the carboxyl functional group, –COOH. They are weak, monobasic (see Chapter 8) acids.

Table 20.2.1 — First members and ionisation

Acid	Formula	Ionisation (partial, weak acid)
Methanoic acid	HCOOH	$HCOOH(aq) \rightleftharpoons HCOO^-(aq) + H^+(aq)$
Ethanoic acid	CH_3COOH	$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$
Propanoic acid	C_2H_5COOH	$C_2H_5COOH(aq) \rightleftharpoons C_2H_5COO^-(aq) + H^+(aq)$

General acidic properties: sour taste; turns blue litmus red; turns green universal indicator orange/yellow (weak acid, so pH is not extremely low).

Table 20.2.2 — Chemical reactions of carboxylic acids

Reaction	General equation
+ reactive metal	carboxylic acid + reactive metal → carboxylate salt + hydrogen gas
+ base/alkali (neutralisation)	carboxylic acid + base → carboxylate salt + water
+ alcohol (esterification)	carboxylic acid + alcohol → ester + water

20.3 | Esters and Esterification

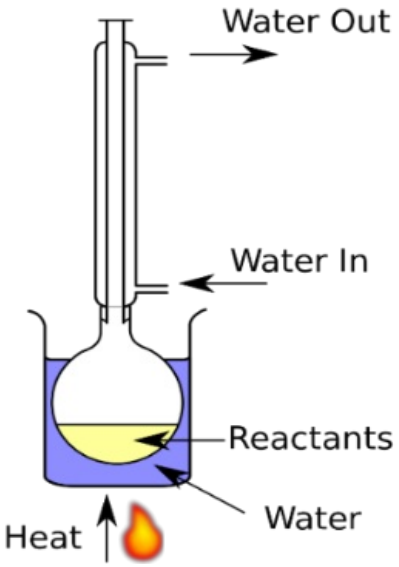


Fig 20.3.1. *Esterification under reflux: carboxylic acid + alcohol with a few drops of concentrated H_2SO_4; the condenser returns evaporated reactants/products back into the flask, preventing loss of volatile substances during prolonged heating.*

DEFINITION 20.3.1 — ESTERIFICATION

The reaction between a carboxylic acid and an alcohol, warmed with a few drops of concentrated sulfuric acid as catalyst under reflux, forming an ester and water.

Table 20.3.1 — Esterification summary

Summary point
Reactants: carboxylic acid + alcohol
Condition: concentrated H_2SO_4 catalyst; reflux (to prevent loss of volatile reactants/products)
General equation: carboxylic acid + alcohol ⇌ ester + water

WORKED EXAMPLE — PROPYL ETHANOATE

Propanol + ethanoic acid ⇌ propyl ethanoate + water

$C_3H_7OH(aq) + CH_3COOH(aq) \rightleftharpoons CH_3COOC_3H_7(aq) + H_2O(l)$

ⓘ NOTE — NAMING ESTERS

An ester name has two parts: the **alcohol** part becomes the first word (e.g. "propyl", from propanol) and the **acid** part becomes the second word with "–oate" (e.g. "ethanoate", from ethanoic acid) — giving "propyl ethanoate."

COMMON EXAM PITFALL

Esterification is also a **condensation reaction** (two molecules combine, releasing a small molecule — water), but it is *not* a neutralisation (acid–base) reaction: the alcohol's –OH group does not ionise to form OH^- in water, so alcohol does not behave as an alkali here.

Isolating the Ester

Esters of small molecular size (up to ~5 carbon atoms) are soluble in water and are separated from the reaction mixture by **fractional distillation** (miscible liquids). Larger esters, being insoluble in water, can instead be separated using a separating funnel.

Fats and Oils as Esters

Natural fats and oils (triglycerides) are esters, since they contain the ester linkage, –COO–. A fat molecule is a **triester** — it contains three ester linkages, formed from one molecule of glycerol (an alcohol with 3 –OH groups) and three molecules of (usually long-chain) carboxylic acids ("fatty acids").

CHAPTER 20 QUICK RECAP: Ethanol is manufactured by anaerobic fermentation of glucose using yeast (~18–37°C), then concentrated by fractional distillation; exposure to air oxidises ethanol to ethanoic acid (vinegar). Carboxylic acids (–COOH) are weak acids that react with metals/bases like any acid, and react with alcohols (concentrated H_2SO_4 catalyst, reflux) in esterification to form an ester + water — a condensation, not a neutralisation, reaction.

Chapter 21: Polymers

Addition and condensation polymers, and plastic waste management

DEFINITIONS 21.1 — POLYMERS AND MONOMERS

A **polymer** is a large molecule built up from many small, repeating units called **monomers**, covalently joined together into long chains. **Polymerisation** is the process of joining monomers together to form a polymer. Different polymers are built from different monomers.

ⓘ NOTE

Polymers do not have a fixed melting/boiling point, since chain length (number of monomers joined) is not fixed — see Chapter 5 for the full structural explanation.

21.1 | Addition Polymerisation

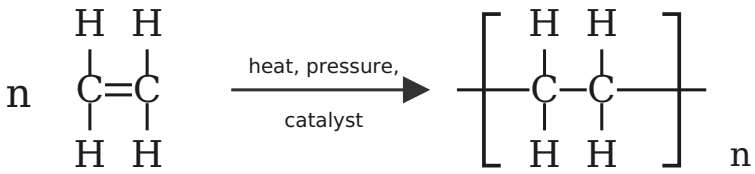


Fig 21.1.1. *Addition polymerisation of ethene: the C=C double bond opens; new single bonds link many monomer units into poly(ethene).*

DEFINITION 21.1.1 — ADDITION POLYMERISATION

A reaction in which many unsaturated monomers (each containing a C=C double bond) join together, without losing any atoms/molecules, to form one large addition polymer as the only product.

Table 21.1.1 — Conditions and examples

Detail
Conditions: high temperature, high pressure, catalyst
Suitable monomers: unsaturated (containing C=C), e.g. ethene, propene, chloroethene, tetrafluoroethene

WORKED EXAMPLE — POLY(ETHENE)

n CH_2=CH_2 → –(CH_2–CH_2)$_n$–

The C=C double bond in each ethene monomer opens up; the resulting single bonds join thousands of monomers together into one long polymer chain. Uses of poly(ethene): plastic bags, clingfilm.

ⓘ NOTE — NAMING ADDITION POLYMERS

Name = "poly" + (name of monomer in brackets), e.g. poly(1,1-dichloroethene) is formed from the monomer 1,1-dichloroethene.

21.2 | Condensation Polymerisation

DEFINITION 21.2.1 — CONDENSATION POLYMERISATION

A reaction in which monomers (usually with two different functional groups) join together with the loss of a small molecule (usually water) at each linkage, forming a condensation polymer.

Nylon (a Polyamide)

Nylon is formed from a **dicarboxylic acid** (a molecule with two −COOH groups) and a **diamine** (a molecule with two $-NH_2$ groups). The −OH of a carboxyl group and one −H from an amino group combine to form water at each linkage, joining the monomers via an **amide linkage** (−CONH−).

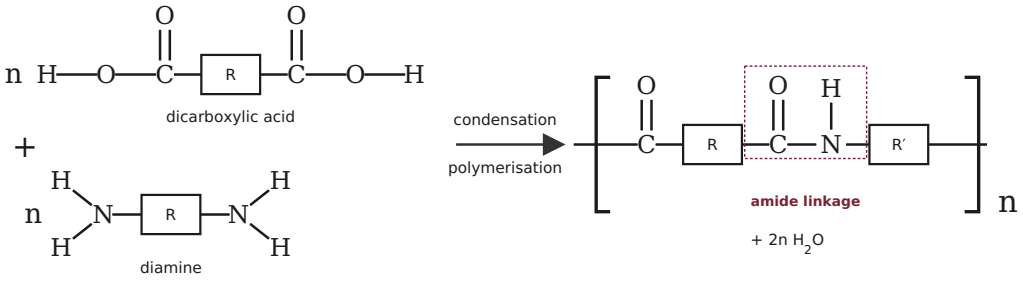


Fig 21.2.1. *Formation of nylon (a polyamide) from a dicarboxylic acid and a diamine monomer, joined by amide linkages with the loss of water.*

Terylene (a Polyester)

Terylene is formed from a **dicarboxylic acid** and a **diol** (a molecule with two −OH groups). The −OH of a carboxyl group and the −H from the diol's hydroxyl group combine to form water at each linkage, joining the monomers via an **ester linkage** (−COO−).

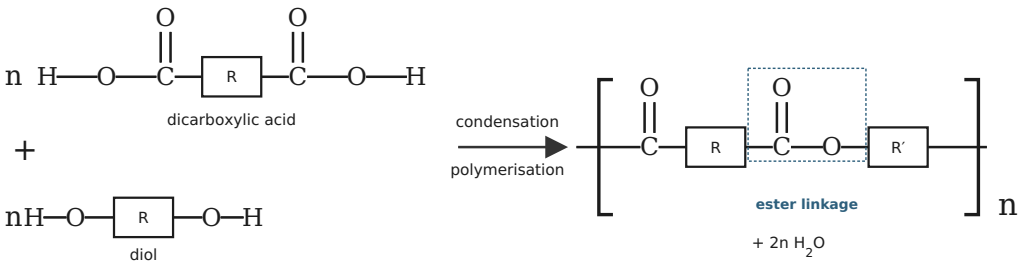


Fig 21.2.2. *Formation of Terylene (a polyester) from a dicarboxylic acid and a diol monomer, joined by ester linkages with the loss of water.*

Table 21.2.1 — Nylon and Terylene

	Nylon	Terylene
Type of polymer	Polyamide (contains the amide linkage, −CONH−)	Polyester (contains the ester linkage, −COO−)
Monomers	Dicarboxylic acid + diamine	Dicarboxylic acid + diol
Typical uses	Clothing, fishing line, parachutes	Clothing, curtain materials, sleeping bags

Table 21.2.2 — Differences Between Addition Polymerisation and Condensation Polymerisation

Differences	Addition Polymerisation	Condensation Polymerisation
Functional group(s) in monomers which are involved in polymerisation	• Carbon-carbon double bond	• Carboxyl group • Hydroxyl group • Amino group
Products of reaction	• 1 molecule of addition polymer	• 1 molecule of condensation polymer and many small molecules such as water and hydrogen chloride (depending on the type of monomers used for the reaction)
Empirical formula of monomers and polymers	• Same	• Different
How repeat units are joined together in the structure of polymer	• Adjacent repeat units are joined to each other by carbon-carbon single bonds	• Polyamide: amide linkage • Polyester: ester linkage

ⓘ NOTE

Deriving the full synthesis mechanism of nylon/Terylene from scratch is not required — focus on recognising the monomer functional groups, the linkage formed (amide/ester), and that water is eliminated at each linkage.

21.3 | Plastic Waste and Its Management

ⓘ NOTE — THE PROBLEM WITH PLASTICS

Most synthetic polymers (plastics) are **non-biodegradable**: they are not broken down by microorganisms, so they persist in the environment for a very long time, accumulating in landfills, polluting oceans, and harming wildlife (e.g. through ingestion or entanglement).

Table 21.3.1 — Recycling methods for plastic waste

Method	Description
Physical method	Melting small pieces of plastic waste (e.g. poly(ethene)) and reforming it into pellets, which can be remoulded into new plastic products
Chemical method — cracking	Long-chain polymer waste is broken down (high temperature, catalyst) into shorter-chain alkanes (used as fuel) and alkenes (used as chemical feedstock)
Chemical method — depolymerisation	Polymers are broken down back into their original monomers, e.g. hydrolysis of polyesters using acid as a catalyst

DEFINITION 21.3.1 — DEPOLYMERISATION

A chemical process in which a polymer is broken down into its constituent monomers (e.g. acid-catalysed hydrolysis of a polyester back into its original carboxylic acid and alcohol monomers).

CHAPTER 21 QUICK RECAP: Addition polymerisation joins unsaturated monomers (C=C) with no by-product (e.g. ethene → poly(ethene)); condensation polymerisation joins monomers with the loss of a small molecule, usually water (e.g. nylon = polyamide, Terylene = polyester). Most plastics are non-biodegradable; waste can be managed by physical recycling (melting/remoulding) or chemical methods (cracking into fuel/feedstock, or depolymerisation back into monomers).

Chapter 22: Maintaining Air Quality

Composition of air, pollutants, acid rain, the ozone layer, carbon cycle and global warming

22.1 | Composition and Uses of Clean, Dry Air

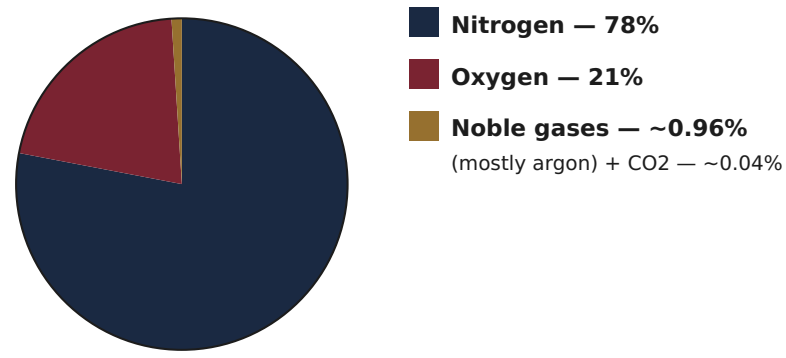


Fig 22.1.1. *Approximate composition by volume of clean, dry air.*

Table 22.1.1 — Approximate composition of clean, dry air

Gas	% by volume
Nitrogen, N_2	78%
Oxygen, O_2	21%
Noble gases (mainly argon) + carbon dioxide	~1% (CO_2 ~0.04%)

> **WORKED EXAMPLE — DETERMINING % OXYGEN IN AIR (HEATED COPPER METHOD)**
>
> A fixed volume of air (100 cm^3) is passed repeatedly over excess heated copper: $2Cu(s) + O_2(g) \rightarrow 2CuO(s)$. Oxygen is consumed and removed from the air, so the gas volume shrinks (e.g. from 100 cm^3 to 79 cm^3). Volume of O_2 reacted = 100−79 = 21 cm^3, so % O_2 = 21/100 × 100% = **21%**. (An equivalent method uses rusting of iron wool in a sealed tube over water, with water rising to occupy the volume vacated by the consumed oxygen.)

Table 22.1.2 — Uses/importance of gases in air

Gas	Importance / uses
Nitrogen	Very unreactive (strong N ≡ N triple bond), so dilutes O_2 and prevents spontaneous combustion; essential component of proteins/DNA in living cells; raw material for manufacturing ammonia
Oxygen	Essential for combustion of fuels; essential for respiration; medical oxygen therapy; oxidiser in welding
Argon (and other noble gases)	Chemically unreactive; provides an inert atmosphere for light bulb filaments and high-temperature metal purification

Fractional distillation of liquid air: air is compressed, then cooled to very low temperatures; water vapour and CO_2 freeze out first (solid/dry ice) and are removed, then the remaining liquid air (N_2, O_2, noble gases) is passed up a fractionating column (hottest at bottom, coolest at top). Nitrogen (b.p. −196°C) distils off first at the top; oxygen (b.p. −183°C) is collected as liquid at the bottom.

22.2 | Air Pollutants

DEFINITION 22.2.1 — AIR POLLUTION

The condition where the atmosphere contains a high concentration of chemicals (pollutants) that may harm living things or damage non-living things. Pollutants arise from natural causes (lightning, volcanoes, wildfires) and human activity (industry, vehicle engines, fires).

Table 22.2.1 — The six common air pollutants

Pollutant	Main source	Harmful effect
Carbon monoxide, CO	Incomplete combustion of carbon-containing fuels, e.g. $C_8H_{18}(g) + 17/2\ O_2(g) \rightarrow 8CO(g) + 9H_2O(g)$	Toxic: binds irreversibly to haemoglobin (forming stable carboxyhaemoglobin), preventing oxygen transport — causes headaches, fatigue, breathing difficulty, death
Oxides of nitrogen, NO & NO_2	High-temperature reaction of N_2 and O_2 in car engines/factories, or from lightning: $N_2(g)+O_2(g) \rightarrow 2NO(g)$; $2NO(g)+O_2(g) \rightarrow 2NO_2(g)$	Contributes to acid rain; irritates eyes/lungs, causes breathing difficulty and bronchitis
Sulfur dioxide, SO_2	Combustion of sulfur impurities in fossil fuels; volcanic eruptions: $S(s)+O_2(g) \rightarrow SO_2(g)$	Contributes to acid rain; irritates eyes/lungs, causes breathing difficulty and bronchitis
Unburnt hydrocarbons (e.g. octane)	Vehicle exhaust, chemical plants	Carcinogenic; reacts with NO_x to form ozone/photochemical smog at low altitude
Methane, CH_4	Bacterial decay of organic matter (incl. landfills); livestock digestion	Greenhouse gas — contributes to global warming
Ozone, O_3 (at low altitude)	NO_2 reacting with unburnt hydrocarbons in sunlight	Component of photochemical smog; irritates eyes/lungs, damages crops

Acid Rain

FORMATION OF ACID RAIN

$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$; $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$ (strong acid)
$4NO_2(g) + 2H_2O(l) + O_2(g) \rightarrow 4HNO_3(aq)$ (strong acid)
Resulting rainwater has pH ≈ 4 or below.

ⓘ NOTE — UNPOLLUTED RAIN IS ALREADY SLIGHTLY ACIDIC

Unpolluted rainwater has pH ≈ 5–5.5 due to dissolved CO_2 forming weak carbonic acid: $CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$. Acid rain (pH ≤ 4) is substantially more acidic than this natural background level (pH 4 is roughly 30 times more acidic than pH 5.5 on the logarithmic pH scale).

Table 22.2.2 — Harmful effects of acid rain

Effect
Corrodes metal structures and limestone/marble buildings (reacts with carbonates and metals)
Lowers pH of lakes/streams (below pH 4), killing fish and aquatic life
Lowers soil pH beyond what many plants can tolerate; leaches essential nutrients (Mg, Ca) from soil
Releases toxic Al^{3+} ions from soil aluminium compounds, harming plants

Control of Air Pollution

Table 22.2.3 — Three main methods of controlling air pollution

Method	Details
Catalytic converters	Fitted in vehicle exhaust systems; platinum/palladium/rhodium catalysts convert pollutants via redox reactions
Flue gas desulfurisation	Wet scrubbing of waste combustion gases using a calcium carbonate (or calcium oxide) slurry, removing SO_2
Liming	Adding calcium oxide/hydroxide directly to acidified soil or water to neutralise excess acidity

REDOX REACTIONS IN A CATALYTIC CONVERTER

$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$ — CO oxidised to less harmful CO_2
$2NO(g) + 2CO(g) \rightarrow N_2(g) + 2CO_2(g)$ — NO reduced to harmless N_2; CO oxidised to CO_2
$2NO_2(g) \rightarrow N_2(g) + 2O_2(g)$ — decomposition of NO_2
$2C_8H_{18}(g) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$ — unburnt hydrocarbons oxidised

FLUE GAS DESULFURISATION REACTIONS

$CaCO_3(s) + SO_2(g) \rightarrow CaSO_3(s) + CO_2(g)$
$2CaSO_3(s) + O_2(g) \rightarrow 2CaSO_4(s)$ (further oxidation)
$CaSO_4(s) + 2H_2O(l) \rightarrow CaSO_4.2H_2O(s)$ (hydrated to gypsum, a useful by-product)
Alternative using lime (CaO): $CaO(s) + SO_2(g) \rightarrow CaSO_3(s)$ (this variant is called "liming")

22.3 | The Ozone Layer

The ozone layer lies roughly 15–30 km above the Earth (in the stratosphere) and absorbs harmful ultraviolet (UV) radiation, acting as a natural "shield" against UV-induced skin cancer, genetic mutations, cataracts, and harm to marine life.

$$3O_2(g) \rightleftharpoons 2O_3(g) \quad \text{(natural formation/breakdown cycle, absorbing UV energy)}$$

ⓘ NOTE — OZONE: HELPFUL VS HARMFUL, DEPENDING ON ALTITUDE

At **low altitude** (ground level), ozone is a harmful pollutant contributing to photochemical smog. In the **stratosphere**, ozone is beneficial, shielding the Earth from UV radiation.

DEFINITION 22.3.1 — CFCS

Chlorofluorocarbons: stable compounds containing carbon, fluorine and chlorine, formerly used widely as aerosol propellants, refrigerant coolants, and in foam packaging.

CFCs are very stable and persist in the atmosphere for a long time, slowly diffusing up into the stratosphere where they react with and destroy ozone, thinning the ozone layer and allowing more UV radiation to reach Earth's surface. Even though many countries have now banned CFCs, ozone depletion continues for years afterward because of the large amount of stable CFCs already present in the atmosphere.

22.4 | The Carbon Cycle

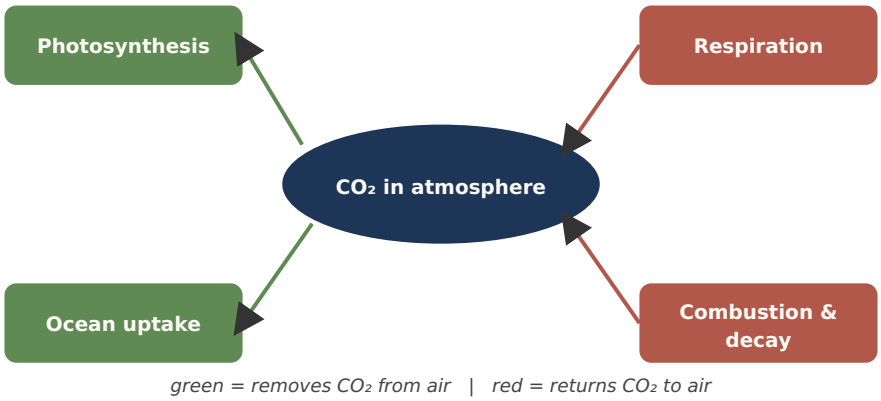


Fig 22.4.1. *The carbon cycle balances processes that remove CO_2 from the atmosphere (photosynthesis, ocean uptake) against processes that return it (respiration, combustion, decay).*

DEFINITION 22.4.1 — CARBON CYCLE

The set of processes that regulate the amount of carbon dioxide in the atmosphere, converting carbon between different forms/compounds.

$$\text{Rate of removal of atmospheric } CO_2 = \text{Rate of return of } CO_2 \text{ to atmosphere}$$
$$\text{(to keep } CO_2 \text{ level constant)}$$

Table 22.4.1 — Processes that remove vs return atmospheric CO_2

Removes CO_2	Returns CO_2
Photosynthesis: $6CO_2(g)+6H_2O(l) \rightarrow C_6H_{12}O_6(aq)+6O_2(g)$	**Respiration**: $C_6H_{12}O_6(aq)+6O_2(g) \rightarrow 6CO_2(g)+6H_2O(l)$
Ocean uptake: dissolved CO_2 used by marine organisms in photosynthesis; converted into calcium carbonate (shells/skeletons), eventually forming limestone deposits	**Combustion of fuels**: e.g. $CH_4(g)+2O_2(g) \rightarrow CO_2(g)+2H_2O(g)$
	Decay/bacterial decomposition of dead organisms

22.5 | Global Warming and Climate Change

DEFINITION 22.5.1 — GLOBAL WARMING

The increase in the average temperature of the Earth's surface due to increasing amounts of greenhouse gases in the atmosphere.

Main greenhouse gases: carbon dioxide (CO_2), methane (CH_4), water vapour (H_2O).

DEFINITION 22.5.2 — GREENHOUSE EFFECT

Greenhouse gases allow visible radiation from the Sun to reach and warm the Earth's surface, but absorb/trap infrared radiation re-emitted by the Earth, reducing the amount of heat that escapes to space. This retains thermal energy in the atmosphere, warming it.

ⓘ NOTE — THE GREENHOUSE EFFECT IS NATURALLY NECESSARY

Without any natural greenhouse effect, Earth's surface temperature would be around −40°C and permanently ice-covered. The *problem* is the **enhanced** greenhouse effect: human activity (burning fossil fuels, deforestation, agriculture) has raised greenhouse gas concentrations beyond natural levels, intensifying warming.

Table 22.5.1 — Potential effects of increased global warming

Potential effect
More extreme/frequent weather events (storms, heatwaves, droughts)
Melting of polar ice caps and glaciers
Rising sea levels, causing flooding of low-lying coastal areas
Disruption to ecosystems and agriculture

CHAPTER 22 QUICK RECAP: Clean dry air is ~78% N_2, 21% O_2, ~1% noble gases/CO_2. Six key pollutants (CO, NO_x, SO_2, unburnt hydrocarbons, CH_4, low-altitude O_3) arise mainly from combustion; NO_x/SO_2 cause acid rain, CO is toxic, CH_4/CO_2 are greenhouse gases. Pollution is controlled via catalytic converters (redox), flue gas desulfurisation/liming ($CaCO_3$/CaO + SO_2). Stratospheric ozone protects against UV but is depleted by CFCs; the carbon cycle balances CO_2 removal (photosynthesis, ocean uptake) against CO_2 return (respiration, combustion, decay); excess greenhouse gases enhance the natural greenhouse effect, driving global warming and climate change.

Appendix A: Key Definitions

A consolidated glossary of every defined term across all 22 chapters

---

Key Definitions, by Chapter

Chapter 1 — Experimental Chemistry

DEFINITION 1.3.1 — MIXTURE

A mixture is a substance made up of two or more substances that are *not* chemically combined (i.e. physically combined, in no fixed ratio).

DEFINITION 1.3.2 — PURE SUBSTANCE

A pure substance is made up of only one element or compound; it is not mixed with any other substance.

DEFINITION 1.3.3 — R_F VALUE

R_f = (distance travelled by solute) ÷ (distance travelled by solvent). R_f values are always between 0 and 1, and are unique to a substance in a given solvent, allowing identification by comparison with known/reference substances.

Chapter 2 — Kinetic Particle Theory

DEFINITION 2.1.1 — KINETIC PARTICLE THEORY

All matter is made up of tiny particles (atoms, molecules or ions) that are in constant, random motion. These particles have mass and occupy space — this is why matter has mass and occupies space.

DEFINITION 2.1.2 — TEMPERATURE

Temperature is a measure of the average kinetic energy of the particles in a substance.

DEFINITION 2.2.1 — DIFFUSION

Diffusion is the net movement of particles from a region of higher concentration to a region of lower concentration.

Chapter 3 — Atomic Structure

DEFINITION 3.1.1 — ELEMENT

An element is a pure substance that cannot be broken down into two or more simpler substances by chemical methods. There are 118 known elements, each made of a unique type of atom.

DEFINITION 3.1.2 — ATOM

An atom is the smallest particle that can still have the chemical characteristics of an element.

DEFINITION 3.1.3 — COMPOUND

A compound is a pure substance containing two or more types of elements chemically combined in a fixed ratio.

DEFINITIONS 3.2.1

Proton (atomic) number, Z — the number of protons in an atom's nucleus (= number of electrons in a neutral atom). Unique to each element.
Nucleon (mass) number, A — the total number of protons and neutrons in the nucleus.
Number of neutrons = nucleon number (A) − proton number (Z)

DEFINITION 3.4.1 — ION

An ion is a charged particle formed when an atom (or group of atoms) gains or loses electron(s); the number of protons and neutrons remains unchanged.

DEFINITION 3.5.1 — ISOTOPES

Isotopes are atoms of the same element with the same number of protons (and electrons) but different numbers of neutrons.

Chapter 4 — Chemical Bonding – Ionic Bonding

DEFINITION 4.1.1 — MIXTURE

A mixture is an impure substance containing two or more substances (elements and/or compounds) physically mixed in any ratio.

DEFINITIONS 4.2.1

Cation — a positively charged particle formed when an atom loses electrons.
Anion — a negatively charged particle formed when an atom gains electrons.
Ionic compound — a pure substance made up of cations and anions held together by strong ionic bonds.
Ionic bonding — the strong electrostatic force of attraction between cations and anions.

DEFINITION 4.3.1 — MOLECULE

A molecule is an electrically neutral particle made of two or more atoms chemically combined by sharing electrons.

DEFINITIONS 4.3.2

Molecule of an element — atoms of the same kind chemically combined by electron sharing.
Molecule of a compound — atoms of different elements chemically combined by electron sharing.

DEFINITION 4.3.3 — COVALENT BOND

The strong electrostatic force of attraction between the positively charged nuclei of two atoms and the electron pair(s) shared between them.

DEFINITION 4.3.4 — VALENCY

The number of electron(s) an atom must lose, gain or share to attain the stable noble gas electronic configuration.

Chapter 5 — Structure and Properties of Materials

DEFINITION 5.1.1 — ALLOTROPES

Allotropes are different structural forms of the same element in the same physical state (e.g. diamond and graphite are both allotropes of carbon).

DEFINITION 5.2.1 — METALLIC BONDING

The strong electrostatic force of attraction between positively charged metal ions and a sea of delocalised, mobile, negatively charged valence electrons.

DEFINITIONS 5.2.2

Malleability — the ability of a material to be hammered into different shapes without breaking.
Ductility — the ability of a material to be drawn into wires.

DEFINITION 5.2.3 — ALLOY

A mixture of a metal with one or more other elements.

Chapter 6 — Chemical Formulae, Equations & Stoichiometry

DEFINITION 6.1.1 — MONATOMIC ELEMENTS

Pure substances that exist as single, uncombined atoms (e.g. noble gases He, Ne, Ar; most metals Na, Fe, Cu).

DEFINITION 6.1.2 — DIATOMIC MOLECULES

Particles made of two atoms covalently bonded together (e.g. H_2, N_2, O_2, Cl_2, HCl).

DEFINITION 6.1.3 — POLYATOMIC MOLECULES

Particles made of three or more atoms covalently bonded together (e.g. H_2O, CO_2, NH_3).

Chapter 7 — The Mole Concept

DEFINITIONS 7.1.1 — RELATIVE MASSES

Relative atomic mass (A_r) — the average mass of one atom of an element relative to 1/12 the mass of one carbon-12 atom.
Relative molecular mass (M_r) — the average mass of one molecule of a covalent substance relative to 1/12 the mass of one carbon-12 atom.
Relative formula mass (M_r) — the average mass of one formula unit of an ionic compound relative to 1/12 the mass of one carbon-12 atom.

DEFINITION 7.1.2 — THE MOLE

One mole of any substance always contains 6.02×10^{23} particles of that substance (the Avogadro constant, L or N_A).

DEFINITION 7.1.3 — MOLAR MASS

The mass of one mole of a substance. Molar mass has the same numerical value as A_r/M_r, but carries units of g/mol.

DEFINITION 7.1.4 — AVOGADRO'S LAW

At the same temperature and pressure, equal volumes of gases contain the same number of particles (moles).

DEFINITION 7.1.5 — MOLAR VOLUME

1 mole of any gas occupies 24 dm^3 (24,000 cm^3) at room temperature and pressure (r.t.p.).

DEFINITION 7.1.6 — CONCENTRATION

The amount of solute dissolved per unit volume of solution.

DEFINITIONS 7.2.1

Molecular formula — shows the actual types and numbers of atoms in one unit of a covalent substance.
Empirical formula — shows the simplest whole-number ratio of atoms of each element in a compound.

DEFINITION 7.2.2 — STOICHIOMETRY

The quantitative relationship between the number of moles of reactants and products in a balanced chemical equation.

DEFINITION 7.3.1 — LIMITING REACTANT

The reactant that is completely used up in a reaction, limiting the amount of product formed. Any other reactant remaining unreacted is "in excess."

DEFINITIONS 7.3.2

Percentage yield — a measure of reaction efficiency.

PERCENTAGE PURITY

A measure of how pure a sample/reactant is (relevant when reactants used are impure).

Chapter 8 — Acids and Bases

DEFINITION 8.1.1 — ACID

A substance which produces hydrogen ions, H^+, in aqueous solution.

DEFINITION 8.1.2 — BASICITY

The number of moles of H^+ ions produced when 1 mole of acid molecules ionises completely in aqueous solution. Basicity determines how many series of salts an acid can form.

DEFINITIONS 8.1.3 — STRENGTH OF ACIDS

Strength — the extent to which an acid ionises to produce H^+ ions in aqueous solution.
Strong acid — ionises completely in aqueous solution (one-way arrow, $\rightarrow$).
Weak acid — ionises only partially in aqueous solution (reversible arrow, $\rightleftharpoons$, typically only ~1% ionised).

DEFINITION 8.2.1 — BASE

Any substance that reacts with an acid to produce salt and water. Types of bases: metal oxides, metal hydroxides, carbonates, hydrogen carbonates, aqueous ammonia.

DEFINITION 8.2.2 — ALKALI

A base that is soluble in water and produces hydroxide ions, OH^-, in aqueous solution. All Group 1 metal oxides/hydroxides, the last four Group 2 metal oxides/hydroxides (Ca, Sr, Ba... down the group), and aqueous ammonia are alkalis.

DEFINITIONS 8.2.3 — STRENGTH OF ALKALIS

Strong alkali (e.g. NaOH, KOH) — ionises completely to give OH^-.
Weak alkali (e.g. NH_3, methylamine) — ionises only partially (~1%), leaving mostly unionised NH_3 molecules alongside NH_4^+, OH^- and H_2O.

DEFINITION 8.4.1 — OXIDE

A binary compound of oxygen and one other element.

DEFINITIONS 8.4.2

Basic oxide — a compound of oxygen and a metal that reacts with acids to form salt and water.
Metal(s) + O_2(g) $\rightarrow$ basic oxide(s); basic oxide + dilute acid $\rightarrow$ salt + water; basic oxide + acidic oxide $\rightarrow$ salt.

DEFINITION 8.4.3 — AMPHOTERIC OXIDE

A metal oxide that reacts with both acids and bases to form salt and water.

DEFINITION 8.4.4 — ACIDIC OXIDE

A compound of oxygen and a non-metal that reacts with bases to form salt and water.

DEFINITION 8.4.5 — NEUTRAL OXIDE

A compound of oxygen and a non-metal that shows neither acidic nor basic properties (e.g. CO, H_2O, NO).

Chapter 9 — Salts

DEFINITION 9.1.1 — SALT

A salt is an ionic compound consisting of a cation and an anion (any anion *except* oxide, O^{2-}, and hydroxide, OH^-, which instead classify the compound as an oxide/hydroxide, i.e. a base).

KEY REQUIREMENT

Both reactants must be soluble (aqueous) salt solutions; the reaction produces an insoluble salt (precipitate) plus a second product made up only of the remaining spectator ions.

Chapter 10 — Ammonia

DEFINITION 10.1.1 — IRREVERSIBLE REACTION

A chemical change in which reactants convert to products, and the products cannot convert back to the reactants. Represented with a one-way arrow, $\rightarrow$.

DEFINITION 10.1.2 — REVERSIBLE REACTION

A chemical change in which the reactants are not entirely converted to products; the reaction proceeds in both directions simultaneously. Represented with the double arrow, $\rightleftharpoons$ (forward reaction reads left $\rightarrow$ right; backward reaction reads right $\rightarrow$ left).

Chapter 12 — Oxidation and Reduction

DEFINITION 12.1 — REDOX REACTION

A chemical reaction in which both oxidation and reduction occur simultaneously.

DEFINITIONS

Oxidation — a reaction in which oxygen combines with an element/compound.
Reduction — a reaction in which oxygen is removed from a compound.
Oxidising agent — a substance that *loses* oxygen to another substance, causing it to be oxidised (itself being reduced).
Reducing agent — a substance that *gains* oxygen from another substance, causing it to be reduced (itself being oxidised).

DEFINITIONS

Oxidation — a reaction in which hydrogen is removed from a compound.
Reduction — a reaction in which hydrogen combines with an element/compound.
Oxidising agent — gains hydrogen from another substance, causing it to be oxidised.
Reducing agent — loses hydrogen to another substance, causing it to be reduced.

DEFINITIONS

Oxidation — loss of electrons.
Reduction — gain of electrons.
Oxidising agent — gains electrons from another substance, causing it to be oxidised.
Reducing agent — loses electrons to another substance, causing it to be reduced.

DEFINITION 12.4.1 — OXIDATION STATE

The charge an atom of an element would have if it existed as an ion in a compound.

DEFINITIONS USING OXIDATION STATE

Oxidation — a reaction in which a reactant's oxidation state increases (loses electrons).
Reduction — a reaction in which a reactant's oxidation state decreases (gains electrons).
Oxidising agent — causes another reactant's oxidation state to increase.
Reducing agent — causes another reactant's oxidation state to decrease.

Chapter 13 — Electrochemistry

DEFINITION 13.1.1 — ELECTROLYSIS

The process of using electricity to break down (decompose) a compound called an **electrolyte**, which must be molten or dissolved in water. This provides evidence that ions exist, held in fixed positions in the solid lattice but free to move when molten/aqueous.

DEFINITION 13.1.2 — ELECTROLYTE

A compound that conducts electricity when molten or in aqueous solution, and decomposes as it does so, due to the presence of mobile cations and anions that act as charge carriers.

DEFINITION 13.4.1 — SIMPLE CELL

A device that uses spontaneous chemical (redox) reactions to generate electricity. Energy conversion: **chemical energy → electrical energy** (the opposite of an electrolytic cell).

Chapter 14 — The Periodic Table

DEFINITION 14.3.1 — REACTIVITY OF NON-METALS

The tendency of a non-metal atom to gain one or more electrons to form a negative ion. Halogens have a high tendency to gain electrons, making them powerful oxidising agents.

Chapter 15 — The Reactivity Series

DEFINITION 15.1 — REACTIVITY OF METALS

The tendency of an atom of a metal to lose one or more electrons to form its positive ion.

DEFINITION 15.2.1 — DISPLACEMENT REACTION OF A METAL

A chemical change where a more reactive metal displaces a less reactive metal from a compound of the less reactive metal.

DEFINITION 15.4.1 — RUSTING

The slow oxidation of iron to form hydrated iron(III) oxide.

Chapter 16 — Chemical Energetics

DEFINITION 16.1.1 — EXOTHERMIC CHANGE

A physical or chemical change that transfers (releases) energy to the surroundings.

DEFINITION 16.1.2 — ENDOTHERMIC CHANGE

A physical or chemical change that takes in (absorbs) energy from the surroundings.

DEFINITIONS

Bond energy — the energy absorbed to break one mole of a particular chemical bond (equal to the energy released when one mole of that bond forms).
Activation energy — the minimum energy that colliding reacting particles must possess in order to react (undergo an effective collision).
Enthalpy change (ΔH) — the difference in energy between products and reactants.

Chapter 17 — Rate of Reaction

DEFINITION 17.2.1 — ACTIVATION ENERGY

The minimum amount of energy that colliding reacting particles must possess for an effective collision (one that results in a reaction) to occur.

DEFINITION 17.4.1 — CATALYST

A substance that increases the rate of a reaction by providing an alternative reaction pathway with a lower activation energy, while remaining chemically unchanged at the end of the reaction (though it may take part in intermediate steps during the reaction).

Chapter 18 — Fuels and Crude Oil

DEFINITIONS 18.1.1

Organic compound — a pure substance containing carbon atoms covalently bonded to atoms of other elements, usually hydrogen, oxygen or chlorine (e.g. methane, ethanoic acid, ethanol).
Hydrocarbon — an organic compound made up of carbon and hydrogen only (e.g. methane, CH_4).
Fossil fuels — mixtures consisting mainly of hydrocarbons, formed from the decayed remains of plants and animals compressed over millions of years. Coal, crude oil (petroleum) and natural gas are the three common fossil fuels.

DEFINITION 18.3.1 — BIOFUEL

A fuel derived from plant or animal matter, offering a renewable alternative to fossil fuels. Examples: bioethanol (from sugarcane/corn fermentation), biodiesel (vegetable oils/animal fats), green diesel (algae), biogas (methane from decomposing organic waste).

DEFINITION 18.3.2 — ENVIRONMENTALLY SUSTAINABLE

Able to be maintained at a steady level without exhausting natural resources or causing ecological damage. A sustainable energy source is renewable, causes little/no pollution, and is carbon neutral (does not contribute to net CO_2 emissions over its life cycle).

Chapter 19 — Hydrocarbons

DEFINITION 19.1.1 — HOMOLOGOUS SERIES

A family of organic compounds with the same general formula and similar chemical properties, because they share the same functional group.

DEFINITION 19.1.2 — FUNCTIONAL GROUP

An atom or group of atoms that gives an organic molecule its characteristic chemical properties/reactions.

DEFINITION 19.2.1 — ALKANES

A homologous series of **saturated** hydrocarbons with general formula C_nH_{2n+2}, containing only C–C and C–H single bonds.

DEFINITION 19.3.1 — ALKENES

A homologous series of **unsaturated** hydrocarbons with general formula C_nH_{2n}, containing at least one carbon–carbon double bond (C=C).

DEFINITION 19.4.1 — ISOMERS

Compounds with the same molecular formula but different structural formulae (different arrangement of atoms). Isomers can differ in: straight-chain vs branched-chain structure; type/position of functional group; type/number of bonds; or position of a substituent/side-chain.

DEFINITION 19.5.1 — CRACKING

The breakdown of large, heavy (long-chain) hydrocarbon molecules into smaller, more useful molecules (shorter-chain alkanes and alkenes), by breaking C–C bonds.

Chapter 20 — Alcohols, Carboxylic Acids & Esters

DEFINITION 20.1.1 — ALCOHOLS

A homologous series with general formula $C_nH_{2n+1}OH$, containing the hydroxyl functional group, –OH.

DEFINITION 20.1.2 — FERMENTATION

A chemical process in which microorganisms (yeast) act on carbohydrates like glucose, in the absence of oxygen, to produce ethanol and carbon dioxide.

DEFINITION 20.2.1 — CARBOXYLIC ACIDS

A homologous series with general formula $C_nH_{2n+1}COOH$, containing the carboxyl functional group, –COOH. They are weak, monobasic (see Chapter 8) acids.

DEFINITION 20.3.1 — ESTERIFICATION

The reaction between a carboxylic acid and an alcohol, warmed with a few drops of concentrated sulfuric acid as catalyst under reflux, forming an ester and water.

Chapter 21 — Polymers

DEFINITIONS 21.1 — POLYMERS AND MONOMERS

A **polymer** is a large molecule built up from many small, repeating units called **monomers**, covalently joined together into long chains. **Polymerisation** is the process of joining monomers together to form a polymer. Different polymers are built from different monomers.

DEFINITION 21.1.1 — ADDITION POLYMERISATION

A reaction in which many unsaturated monomers (each containing a C=C double bond) join together, without losing any atoms/molecules, to form one large addition polymer as the only product.

DEFINITION 21.2.1 — CONDENSATION POLYMERISATION

A reaction in which monomers (usually with two different functional groups) join together with the loss of a small molecule (usually water) at each linkage, forming a condensation polymer.

DEFINITION 21.3.1 — DEPOLYMERISATION

A chemical process in which a polymer is broken down into its constituent monomers (e.g. acid-catalysed hydrolysis of a polyester back into its original carboxylic acid and alcohol monomers).

Chapter 22 — Maintaining Air Quality

DEFINITION 22.2.1 — AIR POLLUTION

The condition where the atmosphere contains a high concentration of chemicals (pollutants) that may harm living things or damage non-living things. Pollutants arise from natural causes (lightning, volcanoes, wildfires) and human activity (industry, vehicle engines, fires).

DEFINITION 22.3.1 — CFCS

Chlorofluorocarbons: stable compounds containing carbon, fluorine and chlorine, formerly used widely as aerosol propellants, refrigerant coolants, and in foam packaging.

DEFINITION 22.4.1 — CARBON CYCLE

The set of processes that regulate the amount of carbon dioxide in the atmosphere, converting carbon between different forms/compounds.

DEFINITION 22.5.1 — GLOBAL WARMING

The increase in the average temperature of the Earth's surface due to increasing amounts of greenhouse gases in the atmosphere.

DEFINITION 22.5.2 — GREENHOUSE EFFECT

Greenhouse gases allow visible radiation from the Sun to reach and warm the Earth's surface, but absorb/trap infrared radiation re-emitted by the Earth, reducing the amount of heat that escapes to space. This retains thermal energy in the atmosphere, warming it.

Appendix B: Complete Equation Reference

Every chemical equation across all 22 chapters, with reaction conditions where applicable

---

This appendix consolidates every balanced chemical equation, half-equation, and general reaction pattern appearing in the notes, organised by chapter. Reaction conditions (heat, catalyst, UV light, etc.) are shown alongside the equation where specified.

Chapter 3 — Atomic Structure

$$^{A}_{Z}\text{X} \text{ where A = nucleon number (top left), Z = proton number (bottom left)}$$

$$A_r = \Sigma \text{ (\% abundance of isotope} \times \text{isotopic mass)} \div 100$$

Chapter 6 — Chemical Formulae, Equations & Stoichiometry

$$4HCl(aq) + MnO_2(s) \rightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$$

$$4H^+(aq) + 4Cl^-(aq) + MnO_2(s) \rightarrow Mn^{2+}(aq) + 2Cl^-(aq) + 2H_2O(l) + Cl_2(g)$$

$$4H^+(aq) + 2Cl^-(aq) + MnO_2(s) \rightarrow Mn^{2+}(aq) + 2H_2O(l) + Cl_2(g)$$

Chapter 7 — The Mole Concept

Mass of element in compound = (A_r × number of atoms of that element ÷ M_r of compound) × mass of compound

Number of moles = Number of particles ÷ 6.02×10^{23}

Molar mass = Mass of substance ÷ Number of moles

Number of moles of gas = Volume (dm^3) ÷ 24

Molar concentration (mol/dm^3) = Moles of solute ÷ Volume of solution (dm^3)

Concentration (g/dm^3) = Mass of solute ÷ Volume of solution (dm^3)

Concentration (mol/dm^3) = Concentration (g/dm^3) ÷ Molar mass of solute

Molecular formula = $(\text{Empirical formula})_n$, where n = Molecular mass ÷ Molar mass of empirical formula

Percentage yield = (Actual yield ÷ Theoretical yield) × 100%

Percentage purity = (Amount of pure substance ÷ Amount of sample) × 100%

$2H_2S(g) + 3O_2(g) \rightarrow 2H_2O(g) + 2SO_2(g)$. If 0.68 g H_2S burns completely, find the mass of SO_2 formed.

$MgCl_2.xH_2O(s) \rightarrow MgCl_2(s) + xH_2O(g)$. Heating 20.3 g of hydrate gives 9.5 g anhydrous $MgCl_2$.

$CH_4(g) + H_2O(g) \rightarrow 3H_2(g) + CO(g)$. Find the volume of methane reacted when 100 cm^3 of H_2 is produced.

6 g H_2 reacts with 32 g O_2: $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$. Find the mass of water formed.

4.8 dm^3 H_2 (r.t.p.) is passed over 67.0 g PbO: $H_2(g) + PbO(s) \rightarrow Pb(s) + H_2O(g)$.

100 cm^3 NH_3 + 120 cm^3 O_2: $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$.

25.0 cm^3 of 3.00 mol/dm^3 $AlCl_3$ + 75.0 cm^3 of 2.00 mol/dm^3 $AgNO_3$: $AlCl_3(aq) + 3AgNO_3(aq) \rightarrow Al(NO_3)_3(aq) + 3AgCl(s)$.

Chapter 8 — Acids and Bases

$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$; $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$; $CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$

$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$; $H_2CO_3(aq) \rightleftharpoons 2H^+(aq) + CO_3^{2-}(aq)$

$H_3PO_4(aq) \rightleftharpoons 3H^+(aq) + PO_4^{3-}(aq)$

metal + dilute acid → salt + hydrogen gas

carbonate + dilute acid → salt + CO_2 + water

hydrogencarbonate + dilute acid → salt + CO_2 + water

metal oxide + dilute acid → salt + water

metal hydroxide + dilute acid → salt + water

alkali + acid → salt + water

alkali + ammonium salt → salt + ammonia gas + water

alkali + metal B salt → metal B hydroxide (ppt.) + salt

$ZnO(s) + 2NaOH(aq) \rightarrow Na_2ZnO_2(aq) + H_2O(l)$ — zinc oxide acting as an acid with a strong alkali, forming aqueous sodium zincate.

Chapter 9 — Salts

$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$

Chapter 10 — Ammonia

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ (forward reaction is exothermic)

Complete ionisation of strong acids/alkalis, e.g. $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$
Neutralisation of strong acid + strong base, e.g. $2HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2(aq) + 2H_2O(l)$
Precipitation reactions, e.g. $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$
Displacement reactions (metals and halogens), e.g. $Zn(s) + CuCl_2(aq) \rightarrow ZnCl_2(aq) + Cu(s)$
Combustion, e.g. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$
Thermal decomposition (most compounds), e.g. $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$

$2I^-(aq) \rightarrow I_2(aq) + 2e^-$: the oxidising agent oxidises colourless I^- to brown I_2, increasing iodine's oxidation state from -1 to 0

$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$. H_2 gains oxygen from CuO to form H_2O — H_2 is oxidised. CuO loses oxygen to H_2 to form Cu — CuO is reduced. Since both occur simultaneously, this is a redox reaction.

$H_2S + Cl_2 \rightarrow S + 2HCl$. H_2S loses hydrogen to Cl_2 to form S — H_2S is oxidised. Cl_2 gains hydrogen from H_2S to form HCl — Cl_2 is reduced.

$2Br^-(aq) \rightarrow Br_2(aq) + 2e^-$ — Br^- loses electrons to Cl_2, so NaBr is oxidised.

$Cl_2(aq) + 2e^- \rightarrow 2Cl^-(aq)$ — Cl_2 gains electrons from Br^-, so Cl_2 is reduced.

$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$. Cu's oxidation state increases from 0 to $+2$ — Cu is oxidised. O's oxidation state decreases from 0 to -2 — O_2 is reduced.

The less reactive metal; site of reduction of cations from the electrolyte, e.g. $2H^+(aq) + 2e^- \rightarrow H_2(g)$. Electrons flow through the external circuit from anode to cathode.

O_2 gas supplied; reduced: $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ (H_2:O_2 consumed in a 2:1 ratio)

At the cathode (reduction): $Na^+(l) + e^- \rightarrow Na(l)$

At the anode (oxidation): $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$

Overall: $2NaCl(l) \rightarrow 2Na(l) + Cl_2(g)$

Cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ — Cu is below hydrogen, so Cu^{2+} is preferentially discharged (pink/brown deposit).

Anode: $4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$ — SO_4^{2-} is never discharged, so OH^- is (colourless gas, relights glowing splint).

Cathode: $2H^+(aq) + 2e^- \rightarrow H_2(g)$ — Na is more reactive than hydrogen, so H^+ is discharged instead.

Anode: $4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$ — dilute solution, so $OH^- > Cl^-$ in relative concentration, and OH^- is discharged.

Overall this is effectively the electrolysis of water: $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$, producing H_2 and O_2 in a 2:1 volume ratio.

Cathode: $2H^+(aq) + 2e^- \rightarrow H_2(g)$ (same as above).

Anode: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$ — now $[Cl^-] \gg [OH^-]$, so Cl^- is discharged instead (used industrially to manufacture chlorine gas, hydrogen gas and sodium hydroxide — the chlor-alkali process).

Anode: $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$ (Mg electrode gradually dissolves/loses mass)

Cathode: $2H^+(aq) + 2e^- \rightarrow H_2(g)$ (effervescence at the silver electrode)

Overall: $Mg(s) + 2H^+(aq) \rightarrow H_2(g) + Mg^{2+}(aq)$

Chapter 14 — The Periodic Table

$$2M(s) + 2H_2O(l) \rightarrow 2MOH(aq) + H_2(g)$$
(exothermic)

$$2M(s) + 2HCl(aq) \rightarrow 2MCl(aq) + H_2(g)$$

$$4M(s) + O_2(g) \rightarrow 2M_2O(s)$$

Chapter 15 — The Reactivity Series

$$\text{metal carbonate(s)} \rightarrow \text{metal oxide(s)} + \text{carbon dioxide(g)}$$

$$4Fe(s) + 3O_2(g) + 2xH_2O(l) \rightarrow 2Fe_2O_3.xH_2O(s)$$

$$\text{metal(s)} + \text{water(l)} \rightarrow \text{metal hydroxide(aq)} + \text{hydrogen gas(g)}$$

$$\text{metal(s)} + \text{steam(g)} \rightarrow \text{metal oxide(s)} + \text{hydrogen gas(g)}$$

$$\text{metal(s)} + \text{dilute acid(aq)} \rightarrow \text{salt(aq)} + \text{hydrogen gas(g)}$$

$$\text{metal oxide(s)} \rightarrow \text{metal(s)} + \text{oxygen(g)} \text{ — via electrolysis}$$

$$\text{metal oxide(s)} + \text{carbon(s)} \rightarrow \text{metal(s)} + \text{carbon dioxide(g)}$$

$$\text{metal oxide(s)} + \text{hydrogen(g)} \rightarrow \text{metal(s)} + \text{steam(g)}$$

Chapter 16 — Chemical Energetics

$$\Delta H = E_{products} - E_{reactants}$$

$$\Delta H = \Delta H_{bonds\ broken} + \Delta H_{bonds\ formed}$$

Electrolysis of molten ore; thermal decomposition of metal carbonates; photosynthesis; formation of nitrogen oxides (e.g. in a car engine or lightning: $N_2(g) + O_2(g) \rightarrow 2NO(g)$); decomposition of CFCs by sunlight (e.g. $CCl_2F_2 \rightarrow CClF_2 + Cl$)

Chapter 17 — Rate of Reaction

$$\text{Average rate of reaction} = \text{Amount of reactant used (or product formed)} \div \text{Time taken}$$

Chapter 18 — Fuels and Crude Oil

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g) \quad \text{(complete combustion)}$$

crude oil (limiting O_2) + oxygen →
carbon (soot) + carbon monoxide + carbon dioxide + steam + sulfur dioxide

Photosynthesis (as sugarcane grows): $6H_2O(l) + 6CO_2(g) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$

Combustion (when bioethanol burns): $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$

Chapter 19 — Hydrocarbons

long-chain alkane → shorter-chain alkane + shorter-chain alkene(s)
(+ hydrogen gas, in some cases)

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

Conditions: Excess oxygen

$$2CH_4(g) + 3O_2(g) \rightarrow 2CO(g) + 4H_2O(l)$$

Conditions: Limited oxygen

$$CH_4(g) + O_2(g) \rightarrow C(s) + 2H_2O(l)$$

Conditions: Very limited oxygen

$$CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g)$$

Conditions: UV light/sunlight

$$CH_3Cl(g) + Cl_2(g) \rightarrow CH_2Cl_2(g) + HCl(g)$$

Conditions: UV light/sunlight

$$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$$

Conditions: Excess oxygen

$$C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$$

Conditions: r.t.p.

$$C_2H_4(g) + H_2O(g) \rightarrow C_2H_5OH(g)$$

Conditions: Catalyst (phosphoric acid), ~300°C, 60–70 atm

$$C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$$

Conditions: Nickel catalyst, 150°C

$$n\ C_2H_4 \rightarrow -(CH_2CH_2)_n- \text{ (poly(ethene))}$$

Conditions: High pressure, catalyst (Chapter 21)

Chapter 20 — Alcohols, Carboxylic Acids & Esters

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$

$$C_6H_{12}O_6(aq) \xrightarrow{\text{yeast}} 2C_2H_5OH(aq) + 2CO_2(g)$$

$$C_2H_5OH + O_2 \rightarrow CH_3COOH + H_2O$$

$$HCOOH(aq) \rightleftharpoons HCOO^-(aq) + H^+(aq)$$

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

$$C_2H_5COOH(aq) \rightleftharpoons C_2H_5COO^-(aq) + H^+(aq)$$

$$\text{carboxylic acid} + \text{reactive metal} \rightarrow \text{carboxylate salt} + \text{hydrogen gas}$$

$$\text{carboxylic acid} + \text{base} \rightarrow \text{carboxylate salt} + \text{water}$$

$$\text{carboxylic acid} + \text{alcohol} \rightarrow \text{ester} + \text{water}$$

$$\text{General equation: carboxylic acid} + \text{alcohol} \rightleftharpoons \text{ester} + \text{water}$$

$$C_3H_7OH(aq) + CH_3COOH(aq) \rightleftharpoons CH_3COOC_3H_7(aq) + H_2O(l)$$

Chapter 21 — Polymers

$$n\ CH_2{=}CH_2 \rightarrow -(CH_2{-}CH_2)_n-$$

$3O_2(g) \rightleftharpoons 2O_3(g)$ (natural formation/breakdown cycle, absorbing UV energy)

Rate of removal of atmospheric CO_2 = Rate of return of CO_2 to atmosphere
(to keep CO_2 level constant)

Incomplete combustion of carbon-containing fuels, e.g. $C_8H_{18}(g) + 17/2\ O_2(g) \rightarrow 8CO(g) + 9H_2O(g)$

High-temperature reaction of N_2 and O_2 in car engines/factories, or from lightning: $N_2(g)+O_2(g) \rightarrow 2NO(g)$;
$2NO(g)+O_2(g) \rightarrow 2NO_2(g)$

Combustion of sulfur impurities in fossil fuels; volcanic eruptions: $S(s)+O_2(g) \rightarrow SO_2(g)$

Photosynthesis: $6CO_2(g)+6H_2O(l) \rightarrow C_6H_{12}O_6(aq)+6O_2(g)$

Respiration: $C_6H_{12}O_6(aq)+6O_2(g) \rightarrow 6CO_2(g)+6H_2O(l)$

Combustion of fuels: e.g. $CH_4(g)+2O_2(g) \rightarrow CO_2(g)+2H_2O(g)$

$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$; $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$ (strong acid)

$4NO_2(g) + 2H_2O(l) + O_2(g) \rightarrow 4HNO_3(aq)$ (strong acid)

$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$ — CO oxidised to less harmful CO_2

$2NO(g) + 2CO(g) \rightarrow N_2(g) + 2CO_2(g)$ — NO reduced to harmless N_2; CO oxidised to CO_2

$2NO_2(g) \rightarrow N_2(g) + 2O_2(g)$ — decomposition of NO_2

$2C_8H_{18}(g) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$ — unburnt hydrocarbons oxidised

$CaCO_3(s) + SO_2(g) \rightarrow CaSO_3(s) + CO_2(g)$

$2CaSO_3(s) + O_2(g) \rightarrow 2CaSO_4(s)$ (further oxidation)

$CaSO_4(s) + 2H_2O(l) \rightarrow CaSO_4.2H_2O(s)$ (hydrated to gypsum, a useful by-product)

Alternative using lime (CaO): $CaO(s) + SO_2(g) \rightarrow CaSO_3(s)$ (this variant is called "liming")

✦ ✦ ✦

You've Reached the End

Twenty-two chapters, hundreds of definitions, and countless reactions later, WE are done :)

GOOD LUCK, cheers and godspeed!

May there be an **ADDITION** *(reaction) to your* ***A1(kene)*** *grades.*

★ ★ ★ ★ ★

Notes compiled with care for your O-Level Chemistry revision, 2026.