Neuratech Academy

STUDENT STUDY NOTES

PREMIUM STUDY NOTES · CLASS 10

EDEXCEL PEARSON IGCSE / O LEVEL

Chemistry (4CH1)

4CH1

STUDENT

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CHEMISTRY (4CH1) Class 10 · Edexcel Pearson IGCSE / O Level · Full syllabus compendium

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23 chapters · ~10,319 words of student-ready notes

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CHEMISTRY (4CH1) | EXAM FREQUENCY: HIGH | Syllabus ref: 1.1 / 1.2

CHAPTER 01 / Particle Theory

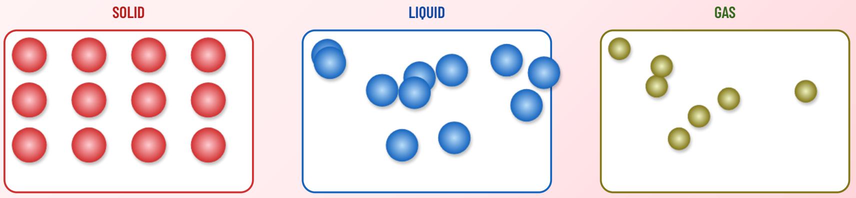
States of Matter

Particles, motion, and the kinetic theory that explains every state change.

Everything around you is made of **tiny particles**. Whether something is a solid, a liquid, or a gas depends on how those particles are arranged and how much energy they have. The same particles can be in any of three states — only their arrangement and motion change.

> 📖 **THE THREE STATES**
>
> **Solid**: particles packed in a regular pattern, vibrating in fixed positions. **Liquid**: particles still close together but able to slide past one another. **Gas**: particles far apart, moving in fast, random directions and filling any container.

FIGURE 1.1



Tight pattern · vibrating only | Close but free · sliding past | Far apart · fast random motion

Particles in solid, liquid, gas — arrangement and motion both change as energy increases.

CHANGES OF STATE (KINETIC THEORY IN ACTION)

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 1.1 / 1.2

CHAPTER 01 / Particle Theory · continued

Heat a solid and its particles vibrate harder until they break free of their fixed positions — the solid **melts** into a liquid. Heat the liquid further and the most energetic particles escape into the gas phase — the liquid **boils**. Cooling reverses each step.

- **Melting**: solid → liquid (at the melting point)
- **Boiling**: liquid → gas at the boiling point — bubbles form throughout the liquid
- **Evaporation**: liquid → gas, only at the surface, at any temperature below the boiling point
- **Condensation**: gas → liquid
- **Freezing**: liquid → solid
- **Sublimation**: solid → gas without passing through the liquid state — e.g. dry ice (solid CO_2) and iodine

BOILING VS. EVAPORATION — WHAT EXAMINERS WANT YOU TO DISTINGUISH

FEATURE	BOILING	EVAPORATION
Where it happens	Throughout the bulk of the liquid (bubbles)	At the surface of the liquid only
Temperature	Only at the boiling point	Any temperature below the boiling point
Speed	Fast	Slower; speeds up with heat or surface area
What you see	Vigorous bubbles + steam	Liquid level drops slowly; no bubbles

💡 **EXAMINER'S TIP**

Pure substances melt and boil at exact, fixed temperatures. Mixtures melt over a range. If a sample melts 'between 78 °C and 84 °C' you can immediately conclude the sample is impure or a mixture.

DIFFUSION

CHEMISTRY (4CH1) **EXAM FREQUENCY: HIGH** Syllabus ref: 1.1 / 1.2

CHAPTER 01 / Particle Theory · continued

Diffusion is the natural spreading-out and mixing of particles from a region of high concentration to a region of low concentration. It happens in liquids and gases (where particles are free to move) but not in solids.

THE COTTON-WOOL DEMONSTRATION

Soak one cotton-wool plug in concentrated ammonia solution and place it in one end of a long glass tube. At the same time, soak another plug in concentrated hydrochloric acid and place it in the other end. The two gases (NH_3 and HCl) diffuse along the tube. Where they meet, a **white smoke** of ammonium chloride forms. The smoke ring forms closer to the HCl end — meaning ammonia particles travelled faster in the same time.

WHY AMMONIA WINS THE RACE

Lighter particles diffuse faster. The relative molecular mass tells the story:
$Mr(NH_3) = 14 + (3 \times 1) = 17$
$Mr(HCl) = 1 + 35.5 = 36.5$
Ammonia is roughly half the mass — so it moves about twice as fast.

DIFFUSION IN LIQUIDS

Drop a small purple crystal of potassium manganate(VII) into a beaker of still water. Over minutes the colour spreads through the whole beaker without stirring. The $KMnO_4$ particles and water molecules slip between each other until the mixture is uniform. Diffusion is slower in liquids than in gases because the particles are closer together and move less freely.

⚠ COMMON MISTAKE

Don't describe boiling as 'evaporation getting faster.' They are different processes. Boiling needs the boiling point; evaporation does not. Boiling produces bubbles in the bulk liquid; evaporation does not.

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 2.1 / 2.2

CHAPTER 02 / Atomic Theory

Atomic Structure

Inside the atom: nucleus, electron shells, the periodic table — and the polyatomic ions you must memorise.

An **atom** is the smallest particle of an element that can take part in a chemical reaction without being changed. Every atom has the same basic plan: a tiny dense **nucleus** at the centre with one or more **electrons** arranged in shells around it. The nucleus contains two kinds of particle — protons (positive) and neutrons (no charge).

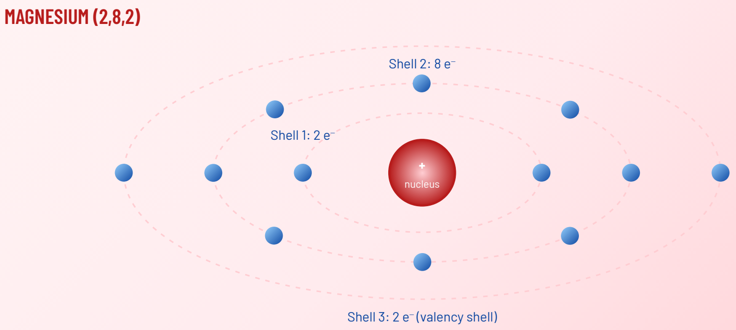
PARTICLE	CHARGE	MASS (RELATIVE)	WHERE IT LIVES
Proton	+1	1	in the nucleus
Neutron	0	1	in the nucleus
Electron	−1	1/1840 (negligible)	in shells around the nucleus

📖 **TWO NUMBERS PER ATOM**

Atomic number (Z) = number of protons. This identifies the element. **Mass number (A)** = total protons + neutrons. Both numbers are written next to the element symbol, with mass on top and atomic number below: e.g. ${}^{12}_{6}C$.

CHEMISTRY (4CH1) **EXAM FREQUENCY: HIGH** Syllabus ref: 2.1 / 2.2

FIGURE 2.1



Magnesium atom — nucleus + three electron shells filled in the order 2, 8, 2.

HOW ELECTRONS ARE ARRANGED

Electrons orbit the nucleus in fixed energy levels called **shells**. The shell nearest the nucleus has the lowest energy; shells further out have higher energy. Electrons fill from the inside out and obey strict capacity limits.

- **Shell 1** holds up to **2** electrons
- **Shell 2** holds up to **8** electrons
- **Shell 3** holds up to 18 in theory, but is treated as **stable at 8** for IGCSE / O-Level

SOME WORKED CONFIGURATIONS

- Hydrogen (Z = 1): **1**
- Carbon (Z = 6): **2,4**
- Oxygen (Z = 8): **2,6**
- Sodium (Z = 11): **2,8,1**
- Magnesium (Z = 12): **2,8,2**
- Chlorine (Z = 17): **2,8,7**
- Argon (Z = 18): **2,8,8** — full outer shell, unreactive noble gas

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 2.1 / 2.2

THE PERIODIC TABLE AT A GLANCE

Elements are arranged in order of increasing atomic number. The horizontal rows are **periods** — the period number tells you how many electron shells the atom has. The vertical columns are **groups** — the group number (for groups 1, 2, 3, 5, 6, 7) equals the number of electrons in the outer shell. All elements in a group share similar chemistry because they share the same outer-shell count.

- **Group 1** — the alkali metals (Li, Na, K, ...)
- **Group 2** — the alkaline earth metals (Mg, Ca, Sr, ...)
- **Group 7** — the halogens (F, Cl, Br, I)
- **Group 0** — the noble gases (He, Ne, Ar, Kr, ...) — full outer shells, unreactive
- **The transition metals** — the central block (Fe, Cu, Zn, etc.)

VALENCY — THE BONDING CURRENCY

> 📖 **VALENCY**
>
> **Valency** is the number of electrons an atom must gain, lose, or share to reach a full outer shell (the configuration of the nearest noble gas). Magnesium (2,8,2) loses its 2 outer electrons to look like neon — so the valency of magnesium is 2.

POLYATOMIC IONS ("GROUP RADICALS") — MEMORISE THESE

A **polyatomic ion** (sometimes called a group radical) is a group of atoms that act together as one charged ion. They appear in countless formulae, so learn them by heart.

CHEMISTRY (4CH1) | EXAM FREQUENCY: HIGH | Syllabus ref: 2.1 / 2.2

CHAPTER 02 / Atomic Theory · continued

VALENCY	POSITIVE IONS	NEGATIVE IONS
1	H^+ hydrogen · Na^+ sodium · K^+ potassium · Ag^+ silver · NH_4^+ ammonium	OH^- hydroxide · NO_3^- nitrate · Cl^- chloride · Br^- bromide
2	Mg^{2+} magnesium · Ca^{2+} calcium · Cu^{2+} copper(II) · Zn^{2+} zinc · Fe^{2+} iron(II) · Pb^{2+} lead	O^{2-} oxide · S^{2-} sulfide · SO_3^{2-} sulfite · SO_4^{2-} sulfate · CO_3^{2-} carbonate
3	Al^{3+} aluminium · Fe^{3+} iron(III)	N^{3-} nitride · PO_4^{3-} phosphate

EXAMINER'S TIP

When metals form compounds with non-metals, the non-metal's name ends in *-ide* (chloride, oxide, sulfide). Some transition metals (Fe, Cu) form more than one ion — the Roman numeral tells you which: copper(I) is Cu^+, copper(II) is Cu^{2+}.

CHEMISTRY (4CH1) EXAM FREQUENCY: MEDIUM Syllabus ref: 2.3

CHAPTER 03 / Atomic Theory

Isotopes

Same element, different mass — and why some isotopes power reactors and treat cancer.

> 📖 **ISOTOPE**
>
> **Isotopes** are atoms of the *same element* with the *same number of protons* but a *different number of neutrons*. They share chemistry (same electrons) but differ in mass.

For example, almost every carbon atom is ^{12}C (6 protons + 6 neutrons). About one in a hundred is ^{13}C (6p + 7n). A trace are ^{14}C (6p + 8n) — rare and radioactive. Chlorine has two stable isotopes: ^{35}Cl and ^{37}Cl, in a 75% : 25% ratio.

WHY ISOTOPES SHARE CHEMISTRY

Chemical reactions are decided by **electrons**, not neutrons. Both ^{35}Cl and ^{37}Cl have the same configuration (2,8,7) so they react identically — both make sodium chloride with sodium, both make hydrogen chloride with hydrogen. The mass difference shows up in physical properties (density, melting point) but not in the chemistry.

RELATIVE ATOMIC MASS CALCULATION

The relative atomic mass (Ar) shown on the periodic table is a weighted average across all naturally occurring isotopes. For chlorine, with 75% ^{35}Cl and 25% ^{37}Cl:

WORKED EXAMPLE

Ar(Cl) = (35 × 75/100) + (37 × 25/100) = 26.25 + 9.25 = **35.5**. That is exactly the value you see on the periodic table.

RADIOACTIVE ISOTOPES

Some isotopes have unstable nuclei. Sooner or later they break apart, releasing radiation in the form of particles, rays, and a large amount of energy. These are **radioisotopes**. Three real-world uses span medicine, industry, and energy.

USE CASE	EXAMPLES
Medical	Cobalt-60 to treat cancer; iodine-131 to study and treat thyroid disorders; gamma rays to sterilise medical equipment
Industrial	Detecting the thickness of paper or aluminium foil during manufacture; finding cracks and leaks in pipelines
Energy	Uranium-235 in nuclear reactors — bombarded by neutrons, the U-235 nucleus splits (nuclear fission), releasing huge amounts of heat that drive turbines to generate electricity

EXAMINER'S TIP

Don't confuse *nuclear fission* (a splitting of one atom) with a chemical reaction (a rearrangement of electrons). Fission changes the identity of the element; chemistry never does.

⚠ COMMON MISTAKE

Different elements can share a mass number — argon-40 and calcium-40 both have A = 40. They are **not** isotopes of each other; isotopes must share the proton number.

CHEMISTRY (4CH1)   EXAM FREQUENCY: HIGH   Syllabus ref: 3.1 / 3.2 / 3.3

CHAPTER 04 / Bonding

Chemical Bonding

Three ways atoms join: ionic, covalent, and metallic — and how the bond determines properties.

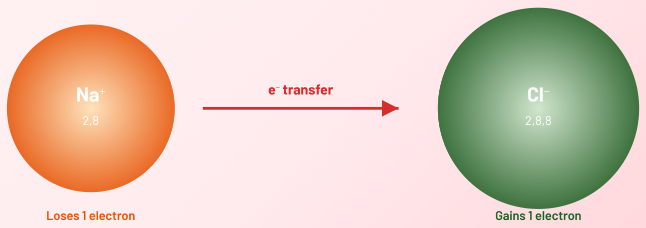
Most atoms are unstable on their own. They achieve a full outer shell — the electron configuration of the nearest noble gas — by **bonding**: gaining, losing, or sharing electrons with other atoms. There are three big bond types to know.

1. IONIC BONDING (METAL + NON-METAL)

A metal atom **loses** outer-shell electrons; a non-metal atom **gains** them. The metal becomes a positive ion (cation), the non-metal becomes a negative ion (anion). The opposite charges then attract by **electrostatic force** — that pull is the ionic bond.

FIGURE 4.1

SODIUM CHLORIDE — IONIC BONDING



Sodium gives one electron to chlorine. Sodium becomes Na^+; chlorine becomes Cl^-; the two attract.

THREE MORE WORKED EXAMPLES

- Calcium chloride: Ca(2,8,8,2) loses 2 e^- to become Ca^{2+}; two Cl(2,8,7) atoms each gain 1 e^- to become Cl^-. Formula: **$CaCl_2$**.
- Magnesium oxide: Mg(2,8,2) loses 2 e^-; O(2,6) gains 2 e^-. Formula: **MgO**.
- Aluminium oxide: 2 Al atoms lose 3 e^- each (total 6); 3 O atoms gain 2 e^- each (total 6). Formula: **Al_2O_3**.

2. COVALENT BONDING (NON-METAL + NON-METAL)

Two non-metal atoms **share** a pair of electrons so that each atom ends up with a full outer shell. A shared pair is one covalent bond. Atoms can share more than one pair if they need to.

- **Single bond** (one shared pair): hydrogen H—H, hydrogen chloride H—Cl, methane (each H—C)
- **Double bond** (two shared pairs): oxygen O=O, carbon dioxide O=C=O
- **Triple bond** (three shared pairs): nitrogen N≡N (very strong; that is why N_2 is so unreactive)

> 📖 **PROPERTIES OF SIMPLE COVALENT COMPOUNDS**
>
> Low melting and boiling points (only weak forces between molecules); usually liquids or gases at room temperature; do not conduct electricity (no charged particles to move); often soluble in organic solvents (alcohol, benzene) but most are not soluble in water.

READING CHEMICAL SHORTHAND

> ⚠ **CL_2 VS 2CL VS $2CL_2$ — THEY MEAN DIFFERENT THINGS**
>
> **Cl_2** = one molecule of chlorine (two atoms bonded). **2Cl** = two separate (uncombined) atoms of chlorine. **$2Cl_2$** = two molecules of chlorine (four atoms total). Marking schemes care — the subscript and the prefix tell different stories.

3. METALLIC BONDING (METAL + METAL)

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 3.1 / 3.2 / 3.3

CHAPTER 04 / Bonding · continued

In a metal, the outer-shell electrons leave their parent atoms and form a **sea of delocalised electrons** that can move freely throughout the structure. Positive metal ions sit in a regular lattice surrounded by this electron sea. The strong attraction between the positive ions and the negative electron cloud is the **metallic bond**.

PROPERTY ROUND-UP

PROPERTY	IONIC COMPOUND	SIMPLE COVALENT	METAL
Melting / boiling point	high	low	high (mostly)
Conducts when solid?	no	no	yes (delocalised electrons)
Conducts when molten/aqueous?	yes (free ions)	no	yes
Soluble in water?	often	rarely	no (reacts instead, if at all)
State at 25 °C	solid	liquid or gas	solid (mercury is the only liquid)

💡 **EXAMINER'S TIP**

When asked 'Why does X conduct?' the mark scheme rewards the phrase **charged particles that are free to move**. For ionic: ions, only when molten or dissolved. For metals: delocalised electrons, always.

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 3.4

CHAPTER 05 / Bonding

Chemical Formulae & Equations

Reading, writing, and balancing — the language of chemistry on paper.

A **formula** is shorthand for a compound. Na means a sodium atom; NaCl means one sodium combined with one chlorine; H_2O means two hydrogens to one oxygen. Building a formula correctly means matching charges so the compound is electrically neutral.

HOW TO BUILD AN IONIC FORMULA

1. Write the cation (positive ion) first, then the anion (negative ion).
2. Identify the charges. Cross-swap them to give the subscripts.
3. Simplify if both numbers share a common factor.
4. Check: the total positive charge equals the total negative charge.

WORKED EXAMPLES

- Sodium sulfide — Na^+ and S^{2-} → need 2 Na^+ per S^{2-} → **Na_2S**
- Aluminium oxide — Al^{3+} and O^{2-} → cross-swap gives 2 and 3 → **Al_2O_3**
- Magnesium chloride — Mg^{2+} and Cl^- → need 2 Cl^- per Mg^{2+} → **$MgCl_2$**
- Sodium sulfate — Na^+ and SO_4^{2-} → need 2 Na per SO_4 → **Na_2SO_4**
- Ammonium carbonate — NH_4^+ and CO_3^{2-} → **$(NH_4)_2CO_3$** (brackets are needed because the unit is two NH_4 not two N)
- Copper(II) sulfate — Cu^{2+} and SO_4^{2-} → **$CuSO_4$**

BALANCING EQUATIONS

CHEMISTRY (4CH1) **EXAM FREQUENCY: HIGH** Syllabus ref: 3.4

CHAPTER 05 / Bonding · continued

In a chemical reaction, atoms are not created or destroyed (the law of conservation of mass). Both sides of an equation must have the same number of each type of atom. Balance by placing whole-number coefficients *in front of* formulae — never change the subscripts inside a formula.

1. Write the unbalanced equation: $Mg + O_2 \rightarrow MgO$
2. Count: Mg 1 = 1 ✓, O 2 vs 1 ✗
3. Add coefficients: **$2Mg + O_2 \rightarrow 2MgO$**
4. Recount: Mg 2 = 2 ✓, O 2 = 2 ✓ — balanced.

EXAMINER'S TIP

After balancing, add the state symbols: **(s)** solid, **(l)** pure liquid, **(g)** gas, **(aq)** dissolved in water. Many marking schemes award a separate mark for correct state symbols.

COMMON MISTAKE

Don't change a subscript to balance. H_2O cannot become H_2O_2 just to give you another oxygen — H_2O_2 is a different chemical (hydrogen peroxide).

CHEMISTRY (4CH1) | EXAM FREQUENCY: HIGH | Syllabus ref: 10.1 / 10.4 / 9.2

CHAPTER 06 / The Periodic Table

Metals, Alloys, and the Periodic Trends

Group 1, Group 2, the transition metals, and Group 7 — patterns that let you predict reactions.

About three quarters of all elements are metals. They share the same metallic bond — positive ions held in place by a sea of delocalised electrons — which gives them a recognisable bundle of properties.

TYPICAL METAL PROPERTIES

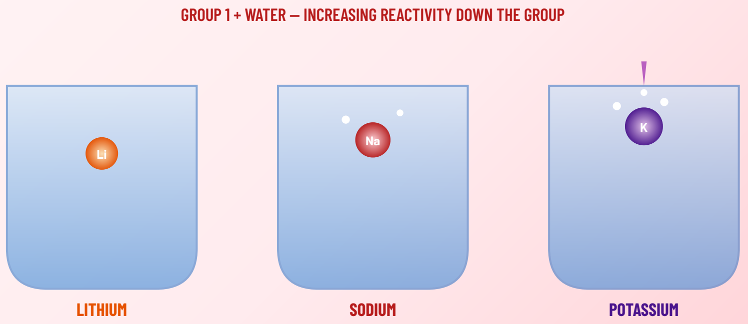
- **Shiny** when freshly cut (lustrous)
- **Malleable** — can be hammered into shape without breaking
- **Ductile** — can be drawn into wires
- **Good conductors** of heat and electricity
- **High melting and boiling points** in most cases
- **Dense** compared with non-metals; usually grey (copper and gold are exceptions)

GROUP 1 — THE ALKALI METALS

Lithium, sodium, potassium, rubidium, caesium. Soft enough to cut with a knife. Low melting points. Form white compounds that dissolve in water to give colourless solutions. Have one electron in the outer shell — valency = 1 — and react vigorously to lose it. **Reactivity increases down the group** ($\text{Li} < \text{Na} < \text{K}$) because the outer electron sits further from the nucleus and is more easily lost.

CHEMISTRY (4CH1) | EXAM FREQUENCY: HIGH | Syllabus ref: 10.1 / 10.4 / 9.2

FIGURE 6.1



Group 1 + cold water — lithium fizzes gently; sodium melts to a silver ball; potassium ignites with a lilac flame.

METAL	OBSERVATION IN COLD WATER
Lithium	Floats and fizzes; bubbles of hydrogen released; gradually disappears.
Sodium	Melts into a shiny silver ball; whizzes around the surface; hisses; bubbles of hydrogen.
Potassium	All of the above plus a **lilac flame**; the hydrogen catches fire; sometimes a tiny explosion at the end.

GENERAL EQUATION

Group 1 metal + water → metal hydroxide + hydrogen
Example: $2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$

GROUP 2 — THE ALKALINE EARTH METALS

Beryllium, magnesium, calcium, strontium, barium, plus the radioactive radium. Harder than Group 1, with higher melting points. Tarnish in air because they form an oxide layer. Have two outer electrons (valency 2) and form 2+ ions. **Reactivity again increases down the group**. Calcium reacts steadily with cold water to make calcium hydroxide; magnesium needs steam.

TRANSITION METALS (THE CENTRAL BLOCK)

CHEMISTRY (4CH1) | EXAM FREQUENCY: HIGH | Syllabus ref: 10.1 / 10.4 / 9.2

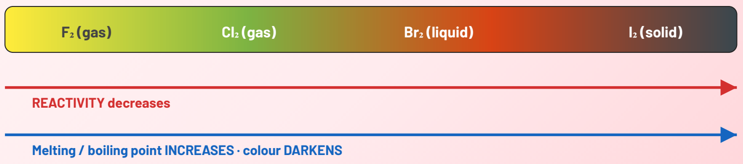
CHAPTER 06 / The Periodic Table · continued

Iron, copper, zinc, nickel, chromium, manganese and friends. They sit between Groups 2 and 3 on the periodic table and behave very differently from Group 1.

PROPERTY	GROUP 1 METALS	TRANSITION METALS
Hardness	Soft (cut with a knife)	Hard, often tough
Density / m.p.	Low	High
Compound colours	White (colourless in solution)	Often coloured (Cu^{2+} blue, Fe^{2+} pale green, Fe^{3+} yellow-brown, Ni^{2+} green)
Catalysis?	No	Yes — many transition metals or their oxides catalyse reactions (Fe in Haber, V_2O_5 in Contact, MnO_2 with H_2O_2)
Oxidation states	Just one (+1)	Often more than one (Fe^{2+} / Fe^{3+}, Cu^+ / Cu^{2+})
Reaction with water	Vigorous	Slow or only with steam

GROUP 7 — THE HALOGENS

FIGURE 6.2



GROUP 7 (HALOGENS) — PHYSICAL TRENDS DOWN THE GROUP

Down Group 7 — reactivity drops, melting/boiling points rise, colour darkens. Fluorine is the most reactive non-metal.

HALOGEN	STATE AT 25 °C	COLOUR
Fluorine (F_2)	gas	very pale yellow
Chlorine (Cl_2)	gas	pale green
Bromine (Br_2)	liquid	red-brown (gives off brown vapour)
Iodine (I_2)	solid	dark grey-black; sublimes to a violet vapour
Astatine (At)	solid	very rare and radioactive

All halogens exist as diatomic molecules. They are poisonous. Reactivity **decreases** down the group: as the outer shell sits further from the nucleus, attracting that 8th electron becomes harder. A more reactive halogen displaces a less reactive one from its compounds — a common exam test.

> **HALOGEN DISPLACEMENT EXAMPLES**
>
> $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$ (orange-brown bromine forms; chlorine wins). $Cl_2 + 2KI \rightarrow 2KCl + I_2$ (the colourless iodide solution darkens to brown). The reverse reactions do not happen.

ALLOYS

An **alloy** is a mixture of a metal with one or more other elements (often another metal). Alloys are usually *harder* than the pure metal because the different-sized atoms disrupt the regular layers, so the layers cannot slide as easily under stress.

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 10.1 / 10.4 / 9.2

CHAPTER 06 / The Periodic Table · continued

ALLOY	COMPOSITION	WHY IT IS USED
Brass	Copper + zinc	Decorative items, taps, instruments — harder and shinier than copper
Bronze	Copper + tin	Statues, ship propellers, coins — corrosion-resistant
Steel (mild)	Iron + ~0.2% carbon	Construction, vehicle bodies — strong yet workable
Steel (high-carbon)	Iron + ~1% carbon	Tools, knives, springs — very hard
Stainless steel	Iron + chromium + nickel	Cutlery, surgical tools, kitchen sinks — does not rust
Solder	Tin + lead	Joining electronic components — melts at low temperature

💡 **EXAMINER'S TIP**

Why is an alloy harder than its pure metal? The mark-scheme phrase: **different-sized atoms prevent the layers from sliding past each other**.

CHEMISTRY (4CH1) **EXAM FREQUENCY: HIGH** Syllabus ref: 10.2

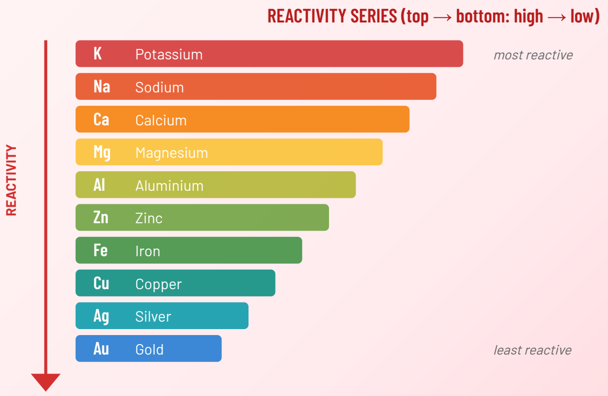
CHAPTER 07 / Metals & Reactions

The Reactivity Series

Ranking metals by their drive to lose electrons — and using the order to predict reactions.

Some metals (potassium, sodium) react fiercely with water; others (gold, platinum) sit unchanged for thousands of years in the ground. The **reactivity series** arranges metals from most reactive at the top to least reactive at the bottom. Knowing the order lets you predict what will happen in any reaction.

FIGURE 7.1



The reactivity series — learn the top group, the bottom group, and where carbon and hydrogen sit (they are non-metals but used as reactivity benchmarks).

WHAT THE SERIES LETS YOU PREDICT

1. REACTION WITH WATER OR STEAM

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 10.2

CHAPTER 07 / Metals & Reactions · continued

METAL	BEHAVIOUR
K, Na, Ca	React vigorously with cold water → metal hydroxide + H_2
Mg	Slow with cold water; fast with steam → magnesium oxide + H_2
Al, Zn, Fe	Only with steam → metal oxide + H_2
Cu, Ag, Au, Pt	No reaction with cold water or steam

2. REACTION WITH DILUTE ACID

Metals **above hydrogen** displace it from a dilute acid — you see fizzing and feel mild warmth. Metals **below hydrogen** (Cu, Ag, Au) do not react. The rate of reaction tracks the position in the series.

> **TWO STANDARD EXAMPLES**
>
> $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
>
> $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$
>
> Note: reacting potassium or sodium with dilute acid is dangerous — it would explode.

3. DISPLACEMENT REACTIONS IN SALTS

A more reactive metal kicks a less reactive metal out of its salt solution. Drop iron filings into copper(II) sulfate solution: the blue colour fades to pale green and a brown-red copper coating appears on the iron.

> **EXAMPLE**
>
> $Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$. The reverse, $Cu + FeSO_4$, does not happen — copper sits below iron.

WHY REACTIVITY DIFFERS

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 10.2

All metal reactions involve the metal **losing electrons** to form a positive ion. The easier a metal lets go of its outer electrons, the more reactive it is. Group 1 metals (one electron to lose, sitting far from the nucleus) are extremely reactive. Transition metals hold their electrons more tightly and react slowly.

REACTIVITY DECIDES EXTRACTION METHOD

POSITION IN SERIES	HOW THE METAL IS EXTRACTED
Most reactive (K, Na, Ca, Mg, Al)	Electrolysis of the molten compound — the only practical route
Mid-reactivity (Zn, Fe, Pb)	Reduction with carbon (coke) in a furnace
Below carbon (Cu, Ag, Au)	Sometimes found native (uncombined) in rocks; otherwise easy roasting or simple displacement

💡 **EXAMINER'S TIP**

Mnemonic for the order: **P**lease **S**top **C**alling **M**e **A** **Z**ebra **I** **L**ove **H**er **C**ats **S**inging **G**lorious — K, Na, Ca, Mg, Al, Zn, Fe, Pb, (H), Cu, Ag, Au.

⚠ **COMMON MISTAKE**

Aluminium *looks* unreactive because it is coated by a tough oxide layer (Al_2O_3). Beneath the layer it is highly reactive — that is why it cannot be extracted by carbon.

CHEMISTRY (4CH1) | EXAM FREQUENCY: HIGH | Syllabus ref: 3.3

CHAPTER 08 / Bonding

Solid Structures

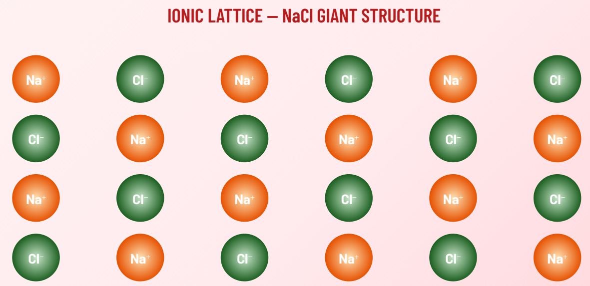
Why diamond is hard, graphite is slippery, NaCl is brittle, and metals bend.

The properties of any solid — how hard, how high-melting, whether it conducts — are decided by what kind of particles it contains and how those particles are joined. Three big families to know.

1. GIANT IONIC LATTICE

Ionic compounds (NaCl, MgO, CaO, KBr) form a regular three-dimensional pattern in which every positive ion is surrounded by negative ions and vice versa. Strong electrostatic attractions in *every direction* have to be broken to melt the solid — hence very high melting points.

FIGURE 8.1



Each ion is surrounded by ions of opposite charge — strong electrostatic attraction

NaCl giant lattice — alternating Na^+ and Cl^- ions held together by strong electrostatic forces in 3D.

- Hard but brittle (a sharp blow lines up like charges, which then repel)
- Crystalline at room temperature
- Do not conduct as solid (ions vibrate but cannot move)
- Conduct when molten or dissolved in water (ions then mobile)
- Soluble in water and other polar solvents; insoluble in alcohol or benzene

2. GIANT COVALENT STRUCTURES

Some non-metals form giant networks of covalent bonds. The two most-tested examples are diamond and graphite — both pure carbon, with very different structures and very different properties.

PROPERTY	DIAMOND	GRAPHITE
Bonding	Each carbon bonded to 4 other carbons in a tetrahedral 3-D net	Each carbon bonded to 3 others; flat layers stacked weakly
Hardness	Very hard — the hardest natural material	Soft and slippery
Electrical conduction	No (no free electrons)	Yes — one delocalised electron per carbon
Why?	Rigid 3-D network; nothing can slide	Layers slide easily because the forces between them are weak
Use	Drill bits, jewellery, abrasives	Pencil leads, lubricant, electrodes, batteries

Other giant covalent structures include silicon dioxide (sand, quartz) and silicon itself, the foundation of all electronics.

3. SIMPLE MOLECULAR STRUCTURES

Small covalent molecules — H_2O, CO_2, I_2, NH_3, CH_4 — have **strong covalent bonds inside each molecule** but only **weak intermolecular forces** between molecules. To melt or boil, only those weak forces have to be overcome — so melting and boiling points are low. The molecules are uncharged, so simple molecular substances do not conduct.

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 3.3

4. GIANT METALLIC STRUCTURE

Metal atoms sit in regular layers; their outer electrons leave the atoms and form a delocalised **sea** through the whole solid. The strong attraction between positive ions and the electron sea is the metallic bond.

WHY METALS ARE MALLEABLE AND DUCTILE

When you hammer a metal, layers of ions slide over each other. The metallic bond is non-directional — it stretches with the move — so the metal changes shape without snapping. The same sliding lets metals be drawn into wires (ductile).

WHY METALS CONDUCT

Apply a voltage and the delocalised electrons drift through the lattice carrying the current. Heat one end of a metal bar and those same electrons carry kinetic energy along the structure — so metals are also good thermal conductors.

> 💡 **EXAMINER'S TIP**
>
> Always name the bond/force you are breaking. Ionic → electrostatic attractions between ions. Giant covalent → covalent bonds (very strong). Simple molecular → weak forces between molecules (NOT covalent bonds!). Metallic → metallic bonds (positive ions and the electron sea).

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 8.1 / 8.2

CHAPTER 09 / Inorganic Chemistry

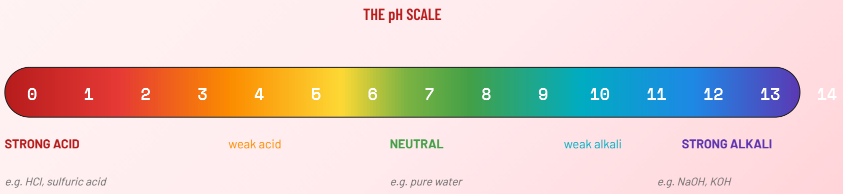
Acids, Bases, and the pH Scale

Definitions, indicators, neutralisation, and how to tell strong from weak.

📖 **ACID · BASE · ALKALI**

An **acid** releases hydrogen ions (H^+) when dissolved in water. A **base** is the chemical opposite of an acid — it neutralises one. Bases are normally metal oxides or hydroxides. A base that dissolves in water is called an **alkali**; alkalis release hydroxide ions (OH^-) in water.

FIGURE 9.1



The pH scale — 0 (very acidic) to 14 (very alkaline). 7 is neutral. Most laboratory acids are pH 1; pure water is 7.

HOW ACIDS REACT

CHEMISTRY (4CH1) **EXAM FREQUENCY: HIGH** Syllabus ref: 8.1 / 8.2

- **Acid + reactive metal** ⟶ salt + hydrogen *(only metals above hydrogen)*
- **Acid + base / alkali** ⟶ salt + water *(neutralisation)*
- **Acid + carbonate** ⟶ salt + water + carbon dioxide
- **Acid + hydrogencarbonate** ⟶ salt + water + carbon dioxide

WORKED EQUATIONS

$2HCl(aq) + Mg(s) \rightarrow MgCl_2(aq) + H_2(g)$

$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

$2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

HOW ALKALIS BEHAVE

- Turn red litmus paper **blue**
- Neutralise acids to make a salt and water
- React with ammonium salts when warmed: alkali + ammonium salt ⟶ salt + water + ammonia gas. The ammonia smell is the test.

AMMONIUM SALT + ALKALI

$NaOH(aq) + NH_4Cl(s) \rightarrow NaCl(aq) + H_2O(l) + NH_3(g)$. Test the gas: it turns damp red litmus paper blue.

STRONG VS. WEAK

A **strong** acid (HCl, H_2SO_4, HNO_3) ionises completely in water — almost every molecule splits into H^+ and the negative ion. A **weak** acid (ethanoic acid, citric acid, carbonic acid) only partially ionises — most molecules stay intact. The same idea applies to alkalis: NaOH and KOH are strong; aqueous ammonia is weak.

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 8.1 / 8.2

⚠ **STRONG ≠ CONCENTRATED**

'Strong' describes how completely the acid ionises. 'Concentrated' describes how much acid is in a given volume of water. You can have a *dilute strong acid* (e.g. 0.1 M HCl) or a *concentrated weak acid* (e.g. glacial ethanoic acid, ~17 M).

HOW TO MEASURE PH

1. **Indicator paper** (litmus, methyl orange, phenolphthalein) tells you whether a solution is acidic, neutral, or alkaline.
2. **Universal indicator** shows a different colour for each pH value — you compare with a colour chart for an approximate pH.
3. **pH meter** — a digital probe that gives a precise pH reading. Required for coloured solutions where indicators cannot be read.

INDICATOR	IN ACID	NEUTRAL	IN ALKALI
Litmus	Red	Purple	Blue
Methyl orange	Red	Orange	Yellow
Phenolphthalein	Colourless	Colourless	Pink
Universal	Red→orange	Green	Blue→violet

COMPARING STRENGTHS IN THE LAB

Two acids of the same concentration can be ranked by:

1. **Universal indicator colour**: red → strong, orange → weak.
2. **Carbonate fizz test**: add a small piece of magnesium carbonate. The acid that fizzes more vigorously is the stronger one because more H^+ ions are available to react.
3. **Conductivity**: a strong acid (more ions) conducts better than a weak one (fewer ions).
4. **Titration**: titrate equal volumes of the two acids against the same alkali. The acid needing the larger volume of alkali to neutralise is the stronger.

CHEMISTRY (4CH1) **EXAM FREQUENCY: HIGH** Syllabus ref: 8.1 / 8.2

EXAMINER'S TIP

When asked 'define' an acid, the mark-scheme answer is **proton donor** (or 'substance that releases H^+ in water'). Saying 'something sour' gets no marks.

CHEMISTRY (4CH1) EXAM FREQUENCY: MEDIUM Syllabus ref: 8.3

CHAPTER 10 / Inorganic Chemistry

Oxides

Acidic, basic, amphoteric, neutral — recognising oxide behaviour.

An **oxide** is a compound of an element with oxygen. Each oxide’s behaviour with acid or alkali tells you which family it belongs to.

TYPE	EXAMPLES	BEHAVIOUR
Basic	Na_2O, K_2O, MgO, CaO, CuO	React with acids → salt + water. Soluble basic oxides (Na_2O, K_2O) dissolve in water to give alkaline solutions.
Acidic	CO_2, SO_2, SO_3, NO_2, P_2O_5, SiO_2	Covalent. React with alkalis → salt + water. Most dissolve in water to give acidic solutions.
Amphoteric	Al_2O_3, ZnO	React with *both* acids and alkalis — insoluble in pure water.
Neutral	CO, NO, H_2O	Do not react with acids or alkalis.

💡 **THE QUICK RULE**

Metal oxide → usually basic. Non-metal oxide → usually acidic. Aluminium and zinc oxides sit on the boundary → amphoteric. A handful of non-metal oxides are simply neutral (CO, NO, H_2O).

WORKED REACTIONS

BASIC VS. AMPHOTERIC IN PRACTICE

$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$ (basic with acid)
$Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$ (amphoteric with acid)
$Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$ (amphoteric with alkali — gives sodium aluminate)
$CO_2 + 2NaOH \rightarrow Na_2CO_3 + H_2O$ (acidic with alkali)

WHY LIME WATER GOES CLOUDY WITH CO_2

Lime water is calcium hydroxide solution $Ca(OH)_2$. Bubble CO_2 through it and a white precipitate of insoluble calcium carbonate forms — that is why the lime water turns milky. Carbon dioxide is an acidic oxide reacting with the alkali. $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$.

Acidic oxides released by industry — SO_2 from burning fossil fuels, NO_2 from car engines — dissolve in atmospheric water to give **acid rain**. Acid rain corrodes limestone buildings, harms aquatic life, and strips minerals from soil.

CHEMISTRY (4CH1) | EXAM FREQUENCY: HIGH | Syllabus ref: 9.1

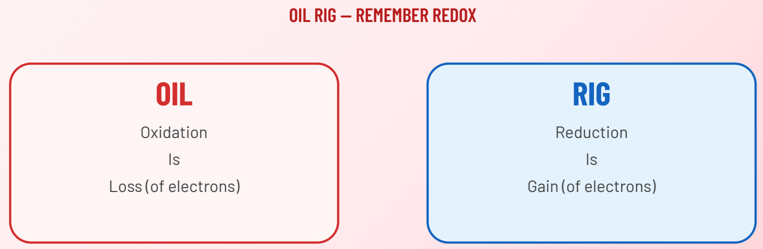
CHAPTER 11 / Reactions

Redox Reactions

Electrons changing hands — and the lab tests that prove it.

> 📖 **MODERN DEFINITION (ELECTRONS)**
>
> **Oxidation** = loss of electrons. **Reduction** = gain of electrons. Whenever one species is oxidised, another must be reduced — redox reactions always pair up.

FIGURE 11.1



__OIL RIG__: __O__xidation __I__s __L__oss, __R__eduction __I__s __G__ain — one mnemonic, full marks.

THE OLDER DEFINITIONS (STILL USEFUL)

- **Oxidation** = gain of oxygen, OR loss of hydrogen.
- **Reduction** = loss of oxygen, OR gain of hydrogen.

WORKED EXAMPLE — OXYGEN TRANSFER

$2Mg + O_2 \rightarrow 2MgO$. Magnesium gains oxygen → **oxidised**. Oxygen loses (no longer free) → **reduced**. Looking at electrons: Mg loses 2 e^- per atom (oxidised); each O atom gains 2 e^- (reduced). Magnesium is the **reducing agent**; oxygen is the **oxidising agent**.

ELECTRON-TRANSFER REACTIONS

Many redox reactions involve no oxygen at all — the electrons just move from one atom to another.

DISPLACEMENT AMONG METALS

Drop a piece of zinc into copper(II) sulfate solution. The blue colour fades; brown copper coats the zinc.

HALF-EQUATIONS

$Zn \rightarrow Zn^{2+} + 2e^-$ (zinc loses electrons — oxidised). $Cu^{2+} + 2e^- \rightarrow Cu$ (copper ion gains electrons — reduced). Zinc is the reducing agent; Cu^{2+} is the oxidising agent.

DISPLACEMENT AMONG NON-METALS

Bubble chlorine through potassium iodide solution. The colourless solution turns brown as iodine is freed:

HALF-EQUATIONS

$Cl_2 + 2e^- \rightarrow 2Cl^-$ (chlorine gained electrons → reduced; oxidising agent). $2I^- \rightarrow I_2 + 2e^-$ (iodide ions lost electrons → oxidised; reducing agent).

LAB TESTS FOR OXIDISING AND REDUCING AGENTS

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 9.1

CHAPTER 11 / Reactions · continued

TEST	WHAT YOU ADD	POSITIVE RESULT
Test for a reducing agent	Acidified potassium manganate(VII), $KMnO_4$ (purple)	Colour changes from purple to colourless
Test for a reducing agent	Acidified potassium dichromate(VI), $K_2Cr_2O_7$ (orange)	Colour changes from orange to green
Test for an oxidising agent	Potassium iodide solution + starch	Iodine freed turns starch indicator **blue-black**

💡 **EXAMINER'S TIP**

Identify the species that *changes* in oxidation state. The one that loses electrons is oxidised — it is the reducing agent. The one that gains electrons is reduced — it is the oxidising agent. Trace the electrons; the rest writes itself.

CHEMISTRY (4CH1) **EXAM FREQUENCY: HIGH** Syllabus ref: 5.1 / 5.2

CHAPTER 12 / Reactions

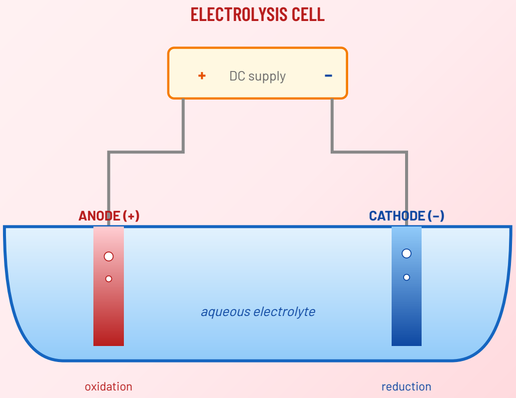
Electrolysis

Splitting compounds with electricity — at the cathode, at the anode, with inert or active electrodes.

📖 **ELECTROLYSIS**

The breakdown of an ionic compound (molten or in aqueous solution) into its elements using a direct electric current. The compound being broken down is the **electrolyte**. Electrical energy is converted to chemical energy — the process is endothermic.

FIGURE 12.1



An electrolysis cell: positive ions migrate to the cathode; negative ions migrate to the anode.

THE MECHANISM

- The electrolyte must be molten or dissolved — ions need to move. Solid ionic compounds do not work.
- Electrons flow through the external wire from anode to cathode (and through the battery).
- Positive ions (cations) drift through the electrolyte to the **cathode** (negative electrode), where they *gain* electrons — reduction.
- Negative ions (anions) drift to the **anode** (positive electrode), where they *lose* electrons — oxidation.

CHEMISTRY (4CH1) **EXAM FREQUENCY: HIGH** Syllabus ref: 5.1 / 5.2

CHAPTER 12 / Reactions · continued

> **MEMORY HOOK**
>
> **PANIC**: **P**ositive ions are **A**ttracted to the **N**egative electrode. Combine with **OIL RIG** (oxidation = loss, reduction = gain) and you have everything.

INERT VS. ACTIVE ELECTRODES

ELECTRODE TYPE	WHAT HAPPENS	EXAMPLES
Inert	Carry the current but take no part in the chemistry. Only the electrolyte's ions react.	Graphite (carbon) and platinum
Active	The electrode itself reacts. The metal anode dissolves into the solution as positive ions; the cathode plates with the metal.	Copper anodes when refining copper; silver anodes when electroplating with silver

MOLTEN COMPOUNDS — THE EASY CASE

When the electrolyte is molten, only the salt's own ions are present. The metal forms at the cathode; the non-metal at the anode.

> **MOLTEN LEAD(II) BROMIDE**
>
> Cathode (−): $Pb^{2+} + 2e^- \rightarrow Pb(l)$ (silvery liquid forms at the bottom)
>
> Anode (+): $2Br^- \rightarrow Br_2(g) + 2e^-$ (red-brown vapour is given off)

AQUEOUS SOLUTIONS — THE TRICKY CASE

Water itself ionises slightly to give H^+ and OH^- ions, so an aqueous solution always contains four kinds of ion competing at the electrodes. Two priority rules decide the winner.

1. **At the cathode**: the *less reactive* species wins. If the metal is below hydrogen in the reactivity series (Cu, Ag) the metal forms. If it is above hydrogen (K, Na, Ca, Mg, Al, Zn) hydrogen forms instead.
2. **At the anode**: *halides win over hydroxide*. Cl^-, Br^-, I^- release Cl_2, Br_2, I_2. With sulfate or nitrate solutions the OH^- wins, releasing oxygen.

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 5.1 / 5.2

CHAPTER 12 / Reactions · continued

ELECTROLYTE	CATHODE	ANODE
molten NaCl	Na	Cl_2
molten Al_2O_3 (with cryolite)	Al	O_2
concentrated NaCl(aq) (brine)	H_2	Cl_2
dilute NaCl(aq)	H_2	O_2
dilute H_2SO_4(aq)	H_2	O_2
$CuSO_4$(aq) with Cu electrodes	Cu plated on cathode	Cu dissolves from anode (purification)

WHY INDUSTRY CARES

- **Aluminium** is extracted from molten Al_2O_3 dissolved in cryolite (lowers the melting point). The only viable method.
- **Copper purification**: impure copper is the anode; pure copper is the cathode. As the anode dissolves, only Cu^{2+} ions plate out at the cathode; impurities (Au, Ag) settle as 'anode sludge'.
- **Chlor-alkali industry**: brine electrolysis gives chlorine (anode), hydrogen (cathode), and sodium hydroxide (left in solution).
- **Electroplating**: the object to be plated is the cathode; the plating metal (Ag, Cr, Ni) is the anode. Used for protection, decoration, and corrosion resistance.

⚠ **COMMON MISTAKE**

Electrolysing aqueous sodium chloride does *not* give sodium at the cathode — sodium is far too reactive in water. You always get hydrogen. Always check whether the electrolyte is molten or aqueous before predicting products.

CHEMISTRY (4CH1) | EXAM FREQUENCY: MEDIUM | Syllabus ref: 10.3

CHAPTER 13 / Industrial Chemistry

Metal Extraction

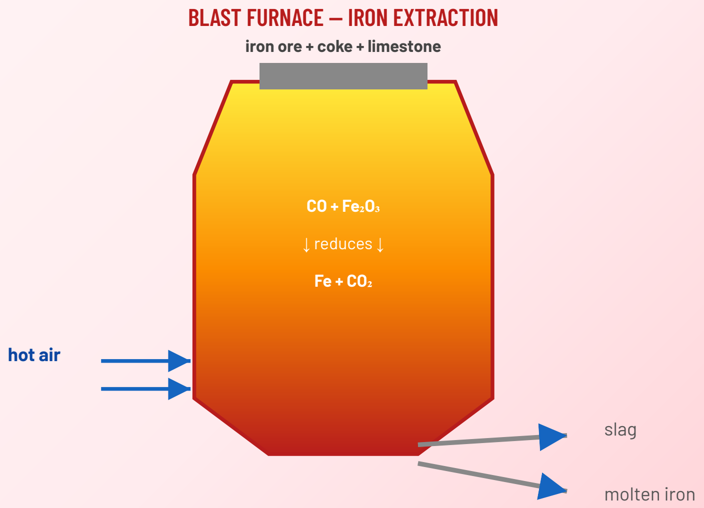
Iron from the blast furnace, aluminium from electrolysis — chosen by reactivity.

How a metal is extracted from its ore depends on where it sits in the reactivity series. The more reactive the metal, the more stable its compounds, and the harder (and more expensive) the extraction.

- **Very reactive** (K, Na, Ca, Mg, Al) — only electrolysis works.
- **Mid reactivity** (Zn, Fe, Pb) — reduce the oxide with carbon (cheap and effective).
- **Below carbon** (Cu, Ag, Au) — sometimes found native (uncombined) in rocks; otherwise simple roasting or displacement.

IRON FROM THE BLAST FURNACE

FIGURE 13.1



Blast furnace — coke + iron ore + limestone enter the top; hot air blows in near the bottom; molten iron and slag tap out separately.

Three raw materials are loaded at the top of the furnace: **iron ore** (haematite, Fe_2O_3), **coke** (almost pure carbon), and **limestone** ($CaCO_3$). Hot air is blasted in at the bottom. Several reactions stack on top of one another.

KEY EQUATIONS

$C + O_2 \rightarrow CO_2$ (coke burns in the hot air)

$CO_2 + C \rightarrow 2CO$ (carbon dioxide is converted to carbon monoxide higher up the furnace)

$\mathbf{Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2}$ (the actual reduction step)

$CaCO_3 \rightarrow CaO + CO_2$ (limestone decomposes)

$CaO + SiO_2 \rightarrow CaSiO_3$ (calcium oxide neutralises silica impurity to form slag)

The molten iron sinks to the bottom and is tapped off; the slag floats on top and is tapped off separately. The iron from the furnace is 'pig iron' and contains carbon. To make steel, oxygen is blown through the molten iron to oxidise excess carbon, and then carefully measured amounts of carbon and other elements (Cr, Ni, Mn) are added back.

ALUMINIUM BY ELECTROLYSIS

Aluminium ore is bauxite, mostly aluminium oxide (Al_2O_3). Aluminium is more reactive than carbon, so carbon cannot reduce its oxide. We electrolyse **molten aluminium oxide dissolved in molten cryolite** (Na_3AlF_6). The cryolite is added to lower the melting point from over 2000 °C to about 950 °C, saving huge amounts of energy.

- Cathode (a carbon-lined steel tank): $Al^{3+} + 3e^- \longrightarrow Al(l)$
- Anode (carbon blocks): $2O^{2-} \longrightarrow O_2(g) + 4e^-$
- The hot oxygen burns the carbon anodes ($C + O_2 \longrightarrow CO_2$), so the anodes are gradually consumed and replaced.

EXAMINER'S TIP

'Why is aluminium expensive?' The mark scheme wants: huge amounts of **electrical energy** are needed for electrolysis, and the carbon anodes are continuously consumed and must be replaced.

CHEMISTRY (4CH1) EXAM FREQUENCY: MEDIUM Syllabus ref: 10.5

CHAPTER 14 / Industrial Chemistry

Rusting and Corrosion

Why iron rusts — and the four ways we stop it.

📖 **RUST**

Rust is hydrated iron(III) oxide, $Fe_2O_3 \cdot xH_2O$. Iron only rusts when *both* water *and* oxygen are present. Salt water speeds rusting up because the dissolved ions help the redox electron transfer.

FIGURE 14.1

Three test tubes prove the point: rust appears only where iron meets both water and air.

THE THREE-TUBE DEMONSTRATION

- Tube 1 — iron nail in tap water exposed to air → rusts.
- Tube 2 — iron nail in air with anhydrous calcium chloride to absorb moisture → **no rust** (no water).
- Tube 3 — iron nail under boiled water sealed by oil → **no rust** (no oxygen).

CHEMISTRY (4CH1) EXAM FREQUENCY: MEDIUM Syllabus ref: 10.5

METHODS OF PREVENTION

METHOD	HOW IT WORKS	WHERE IT IS USED
Painting / oiling / greasing	Physical barrier — keeps water and oxygen out.	Bridges, ships, machinery
Galvanising (zinc coating)	Barrier *plus* sacrificial protection — if the coating is scratched, zinc oxidises preferentially because it is more reactive than iron.	Steel buckets, roofing, dustbins, car panels
Tin-plating	Barrier only. If scratched, the exposed iron rusts faster (tin is below iron).	Food cans (the tin is non-toxic and provides a clean surface)
Sacrificial protection	Blocks of a more reactive metal (Mg or Zn) bolted onto the iron object — the block corrodes instead.	Ship hulls, buried steel pipelines, oil rigs
Alloying (stainless steel)	Chromium and nickel form a tough oxide layer that is unreactive.	Cutlery, surgical instruments, kitchen equipment

> 💡 **EXAMINER'S TIP**
>
> Why does galvanising still protect a scratched object but tin-plating doesn't? The mark scheme wants **sacrificial protection**: zinc is more reactive than iron, so zinc oxidises first. Tin is less reactive than iron, so the iron rusts faster once the tin layer is broken.

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 6.1

CHAPTER 15 / Reactions & Energy

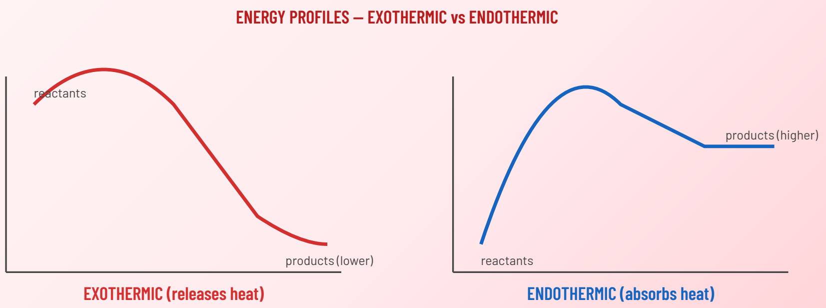
Energy Changes in Reactions

Exothermic, endothermic, and the bond-energy bookkeeping that explains the heat.

Energy is neither created nor destroyed in a chemical reaction — it simply changes form. Most reactions involve a change in heat, called the **energy change**, ΔH. The sign of ΔH tells you whether the reaction releases heat or absorbs it.

> 📖 **EXOTHERMIC VS. ENDOTHERMIC**
>
> **Exothermic** reactions *release* heat to the surroundings — the surroundings get hotter. ΔH is negative. **Endothermic** reactions *absorb* heat from the surroundings — the surroundings cool down. ΔH is positive.

FIGURE 15.1



Energy profiles — exothermic on the left (products lower); endothermic on the right (products higher). The hump in the middle is the activation energy.

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 6.1

TYPE	EXAMPLES
Exothermic	Combustion (burning fuels), neutralisation, most oxidation reactions, respiration, displacement reactions, hand warmers
Endothermic	Thermal decomposition, photosynthesis, dissolving certain salts (e.g. NH_4Cl in water), instant cold packs, melting ice

WHERE DOES ΔH COME FROM? BOND BREAKING AND BOND MAKING.

- **Breaking bonds** requires energy — an endothermic step.
- **Making bonds** releases energy — an exothermic step.
- If the energy released by new bonds is *greater* than the energy absorbed by breaking old bonds → reaction is **exothermic**.
- If the energy released by new bonds is *less* than the energy absorbed → reaction is **endothermic**.

THE MASTER EQUATION

ΔH = (energy needed to break reactant bonds) − (energy released forming product bonds). Negative ΔH → exothermic. Positive ΔH → endothermic.

BOND ENERGIES (AVERAGE VALUES TO USE IN CALCULATIONS)

CHEMISTRY (4CH1)   EXAM FREQUENCY: HIGH   Syllabus ref: 6.1

CHAPTER 15 / Reactions & Energy · continued

BOND	BOND ENERGY (KJ/MOL)
H—H	436
Cl—Cl	242
H—Cl	431
C—H	412
O=O	498
C=O	743
O—H	464
C—C	347
N≡N	945

WORKED EXAMPLE — METHANE COMBUSTION

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

1. Bonds broken: 4 × (C—H) + 2 × (O=O) = (4 × 412) + (2 × 498) = 1648 + 996 = **2644 kJ** in.
2. Bonds made: 2 × (C=O) + 4 × (O—H) = (2 × 743) + (4 × 464) = 1486 + 1856 = **3342 kJ** out.
3. ΔH = 2644 − 3342 = **−698 kJ/mol**. The negative sign tells you methane combustion is strongly exothermic — which is why methane is used as a fuel.

💡 **EXAMINER'S TIP**

Activation energy is the 'hump' on an energy diagram — the minimum energy reactants need before they can react. A catalyst lowers this hump. It does *not* change ΔH or the final yield — only how fast you reach equilibrium.

CHEMISTRY (4CH1) EXAM FREQUENCY: MEDIUM Syllabus ref: 6.2

CHAPTER 16 / Reactions

Reactions: Synthesis, Displacement, Decomposition, and Ionic Equations

How chemistry classifies reactions — and how to write the ionic core of one.

A **physical change** rearranges the form of a substance but not its identity (melting ice, dissolving sugar). A **chemical change** rearranges atoms into a new substance (burning, neutralising, rusting). Chemical changes are usually irreversible by simple physical means and often involve a noticeable energy change.

THREE BIG FAMILIES OF CHEMICAL REACTIONS

1. SYNTHESIS (COMBINATION)

Two or more substances combine to form a single product.

EXAMPLES

$Fe + S \rightarrow FeS$ (iron + sulfur, with heat, gives iron(II) sulfide)
$2Mg + O_2 \rightarrow 2MgO$ (magnesium burns brightly in air to form white powder)
$2H_2 + O_2 \rightarrow 2H_2O$ (hydrogen burns with a squeaky pop)
$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$ (photosynthesis — chlorophyll catalyses it; sunlight provides the energy; an endothermic synthesis)

2. DISPLACEMENT

A more reactive element kicks a less reactive element out of its compound. Includes both metal-metal and halogen displacements.

CHEMISTRY (4CH1) EXAM FREQUENCY: MEDIUM Syllabus ref: 6.2

CHAPTER 16 / Reactions · continued

- Metal displacement: $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$
- Halogen displacement: $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$
- Double displacement (an exchange of partners): $HCl + KOH \rightarrow KCl + H_2O$

3. THERMAL DECOMPOSITION

Heat splits a single substance into two or more simpler ones. Patterns differ by the family of compound.

COMPOUND TYPE	HEAT RESULT	EXAMPLE
Metal carbonates (most)	Metal oxide + CO_2	$CaCO_3 \rightarrow CaO + CO_2$; $CuCO_3 \rightarrow CuO + CO_2$
Sodium / potassium carbonates	Stable to heat	Na_2CO_3 and K_2CO_3 do not decompose
Metal hydroxides (most)	Metal oxide + H_2O	$Ca(OH)_2 \rightarrow CaO + H_2O$
NaOH and KOH	Stable to heat — do not decompose	no reaction
Sulfates (most)	Metal oxide + SO_3	$CaSO_4 \rightarrow CaO + SO_3$
Ammonium salts (NH_4Cl)	NH_3 + acid (re-form on cooling)	$NH_4Cl \rightleftharpoons NH_3 + HCl$
Nitrates (most)	Metal oxide + NO_2 + O_2	$2Cu(NO_3)_2 \rightarrow 2CuO + 4NO_2 + O_2$

WRITING IONIC EQUATIONS

An ionic equation strips away the ‘spectator’ ions that take no part in the reaction, leaving only the species that actually change. Four steps.

CHEMISTRY (4CH1) EXAM FREQUENCY: MEDIUM Syllabus ref: 6.2

1. Write the balanced full symbol equation with state symbols.
2. Re-write all aqueous compounds as separate ions. Solids, pure liquids, and gases stay as molecules.
3. Cross out anything that appears *identically* on both sides — those are the spectator ions.
4. Write what is left.

WORKED EXAMPLE — NEUTRALISATION

Full: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

Ionised: $H^+ + Cl^- + Na^+ + OH^- \rightarrow Na^+ + Cl^- + H_2O$

Spectators (Na^+, Cl^-) cancel out.

Ionic: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$. Every neutralisation between a strong acid and a strong alkali boils down to this single equation.

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CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 6.3

CHAPTER 17 / Reactions

Rate of Chemical Reactions

What makes reactions fast or slow — collision theory and the lab methods that prove it.

> 📖 **RATE OF REACTION**
>
> Amount of product formed (or reactant used up) per unit time. Faster reactions produce more product per second.

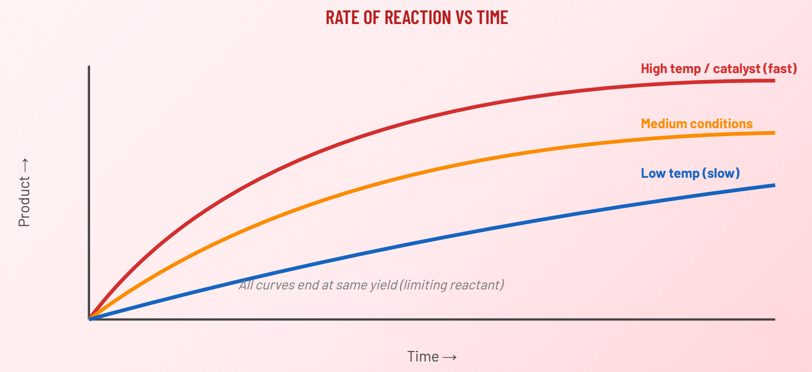
COLLISION THEORY — THE MASTER IDEA

For a reaction to happen, particles must **collide**. The collision must (a) carry at least the activation energy, and (b) be the right way round. Anything that increases the frequency of successful collisions speeds the reaction up.

THE FIVE FACTORS

1. **Concentration ↑**: more particles per unit volume → more collisions per second → faster reaction.
2. **Pressure ↑** (gas reactions only): squeezes particles closer together; same effect as concentration.
3. **Surface area ↑**: a powdered solid has far more surface particles exposed to a liquid than a single lump — faster reaction.
4. **Temperature ↑**: particles move faster *and* a bigger fraction have enough energy to clear the activation hump — rate roughly doubles for every 10 °C rise.
5. **Catalyst**: a substance that increases the rate without being used up. Provides an alternative pathway with a lower activation energy. Does not change ΔH or final yield.

FIGURE 17.1



Rate-vs-time graphs. The steeper the curve at the start, the faster the reaction. All curves end at the same yield when the reactant is the same.

LIGHT AS A SIXTH FACTOR

- **Photosynthesis**: chlorophyll uses sunlight to drive the endothermic synthesis of glucose from CO_2 and H_2O.
- **Photography**: silver halides darken in light. $2AgBr \rightarrow 2Ag + Br_2$ is the underlying photochemical reaction; AgCl is white, AgBr cream, AgI yellow before exposure.

THREE LAB METHODS TO MEASURE RATE

METHOD 1 — MASS LOSS WITH MARBLE CHIPS AND HCL

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$. Place the flask on a balance with a cotton-wool plug. As CO_2 escapes, the flask loses mass. Record the balance reading every 30 s and plot mass loss vs. time. Compare powdered chips (greater surface area) against lumps (smaller surface area) under otherwise identical conditions.

METHOD 2 — GAS VOLUME WITH MAGNESIUM AND HCL

$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$. Collect the hydrogen in a gas syringe; record the volume every 30 s. Run the experiment with two different acid concentrations (e.g. 1 M and 2 M) but the same mass of magnesium. The 2 M curve is roughly twice as steep at the start — doubling concentration roughly doubles the initial rate. Both curves end at the same total volume because the limiting reactant (magnesium) is the same.

METHOD 3 — THE DISAPPEARING CROSS

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 6.3

$Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + S(s) + SO_2(g) + H_2O(l)$. A flask sits on a piece of paper marked with a black cross. As fine yellow sulfur is produced, the solution turns cloudy. Time how long it takes for the cross to vanish. Repeat at several temperatures (warm baths). Higher temperature → shorter time → faster reaction.

METHOD 4 (BONUS) — HYDROGEN PEROXIDE + MNO_2

$2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$. Without a catalyst the reaction is slow. Add ~0.5 g of powdered manganese(IV) oxide (a catalyst) and oxygen pours off. Collect the gas in a syringe and time the volume produced. The black MnO_2 powder is unchanged at the end — that is the catalyst test.

💡 **EXAMINER'S TIP**

When two curves end at the *same* final value but reach it at different times → same amount of reactant; only rate differed. When they end at *different* values → the amount of reactant differed.

⚠ **COMMON MISTAKE**

A catalyst does not start a reaction that wouldn't otherwise happen, and it does not change the ΔH. It just lowers the activation energy so a slow reaction becomes a fast one.

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 7.1

CHAPTER 18 / Reactions

Reversible Reactions and Equilibrium

When products turn back into reactants — and how industry tunes the conditions.

Some reactions go to completion; others reach a balance point where products and reactants coexist. We mark a reversible reaction with two arrows: $A + B \rightleftharpoons C + D$.

> 📖 **DYNAMIC EQUILIBRIUM**
>
> In a **closed system**, when the rate of the forward reaction equals the rate of the reverse reaction, the amounts of reactants and products no longer change. Reactions are still happening — they just balance each other. This is dynamic equilibrium.

LE CHATELIER'S PRINCIPLE

If you change the conditions of a system at equilibrium, the equilibrium **shifts to oppose the change**.

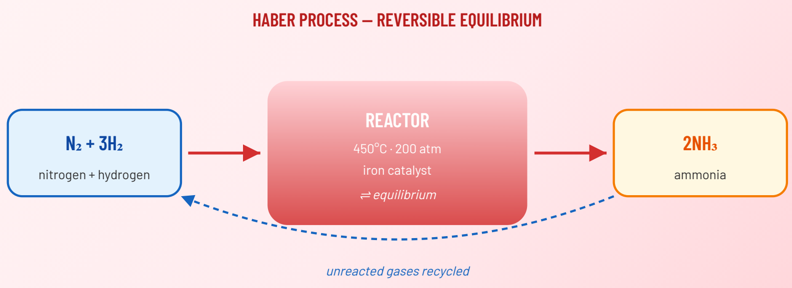
CHANGE	EQUILIBRIUM SHIFTS TO...
Increase concentration of a reactant	produce more product
Increase concentration of a product	produce more reactant
Increase pressure (gas reactions)	the side with FEWER moles of gas
Increase temperature	favour the endothermic direction
Decrease temperature	favour the exothermic direction

⚠ **CATALYSTS AND EQUILIBRIUM**

A catalyst speeds the forward and reverse reactions equally. It does NOT change the equilibrium position — only the time taken to get there.

THE HABER PROCESS — INDUSTRIAL AMMONIA

FIGURE 18.1



The Haber process — nitrogen and hydrogen react reversibly. Recycling unreacted gas pushes the overall yield much higher than the single-pass equilibrium.

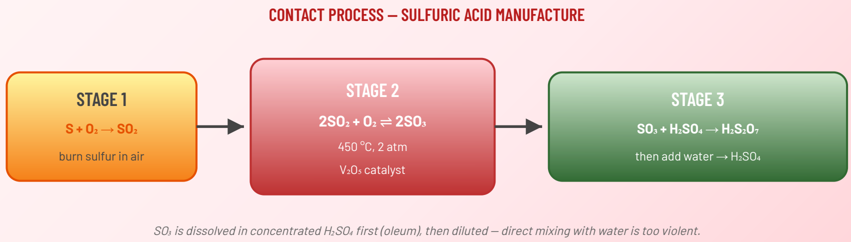
$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ — exothermic forward direction, four moles of gas going to two.

- Lower temperature would favour the product side (exothermic favoured by cold), but the rate would be too slow — uneconomic.
- Higher pressure favours the product side (smaller mole count), but extreme pressures are expensive and dangerous.
- **Compromise conditions**: 450 °C, 200 atm, with an iron catalyst. Unreacted gases are recycled.

Sources: nitrogen comes from the fractional distillation of liquid air; hydrogen comes mainly from natural gas (methane). The ammonia produced is used to make fertilisers (e.g. ammonium sulfate) and nitric acid.

THE CONTACT PROCESS — SULFURIC ACID

FIGURE 18.2



Contact process — three stages: burn sulfur to SO_2, oxidise SO_2 to SO_3 over V_2O_5, then dissolve SO_3 in conc. H_2SO_4 (oleum) and dilute.

1. $S(s) + O_2(g) \rightarrow SO_2(g)$: sulfur is burned in air.
2. $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$: the reversible heart of the process. Conditions: ~450 °C, ~2 atm, vanadium(V) oxide catalyst (V_2O_5). Compromise conditions for yield + rate.
3. $SO_3 + H_2SO_4 \rightarrow H_2S_2O_7$ (oleum), then $H_2S_2O_7 + H_2O \rightarrow 2H_2SO_4$. Direct mixing of SO_3 with water is too violent — it produces a fine acid mist.

💡 **EXAMINER'S TIP**

Always quote condition *and* reasoning. 'The Haber process uses 450 °C because lower temperatures give higher yield but the rate would be too slow — this is the compromise temperature.' Both halves earn marks.

CHEMISTRY (4CH1) **EXAM FREQUENCY: HIGH** Syllabus ref: 8.4

CHAPTER 19 / Inorganic Chemistry

Preparation of Salts

Three reliable routes — neutralisation, titration, precipitation — chosen by solubility.

> 📖 **SALT**
>
> An **ionic compound** formed when the H^+ of an acid is replaced by a metal ion or an ammonium ion. Examples: NaCl from HCl, $MgSO_4$ from H_2SO_4, KNO_3 from HNO_3, $(NH_4)_2SO_4$ from H_2SO_4.

SOLUBILITY RULES TO MEMORISE

- All **nitrates** are soluble.
- All **sodium, potassium, ammonium** salts are soluble.
- Most **chlorides** are soluble (exceptions: AgCl, $PbCl_2$).
- Most **sulfates** are soluble (exceptions: $BaSO_4$, $PbSO_4$, $CaSO_4$).
- Most **carbonates** are insoluble (exceptions: Na, K, ammonium carbonates).
- Most **hydroxides** are insoluble (exceptions: NaOH, KOH; $Ca(OH)_2$ partly soluble).

CHOOSING THE METHOD

CHEMISTRY (4CH1) | EXAM FREQUENCY: HIGH | Syllabus ref: 8.4

TARGET SALT	METHOD	REAGENTS
Soluble salt; base is INSOLUBLE	Neutralisation (excess solid)	Acid + insoluble metal oxide / hydroxide / carbonate
Soluble salt; base is SOLUBLE	Titration	Acid + alkali (both in solution)
Insoluble salt	Precipitation (mixing solutions)	Two soluble salts whose ions combine to make the insoluble target

METHOD 1 — NEUTRALISATION (SOLUBLE SALT FROM INSOLUBLE BASE)

Used for salts of copper, zinc, iron, magnesium, etc. Step-by-step preparation of magnesium sulfate (Epsom salt):

1. Warm dilute sulfuric acid in a beaker (do not boil).
2. Add magnesium oxide a spatula at a time, stirring, until no more dissolves — an excess of solid sits on the bottom.
3. Filter to remove the unreacted MgO.
4. Heat the clear filtrate gently in an evaporating dish to drive off some water; stop when crystallisation point is reached (a sample placed on a glass rod cools to crystals).
5. Allow to cool slowly — the slower the cooling, the bigger the crystals.
6. Filter the crystals; pat dry between filter paper or place in a warm oven.

EQUATION

$$MgO(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2O(l)$$

METHOD 2 — TITRATION (SOLUBLE SALT FROM SOLUBLE BASE)

Used when both reactants are solutions — you cannot 'add an excess' safely, so a titration is needed to mark the exact end point. Step-by-step preparation of sodium nitrate:

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 8.4

1. Pipette 25 cm^3 of sodium hydroxide solution into a clean conical flask.
2. Add 2–3 drops of methyl orange indicator (yellow in alkali).
3. Fill the burette with dilute nitric acid; record the starting reading.
4. Run the acid into the flask, swirling, until the indicator just changes from yellow to red — this is the end point. Record the final burette reading.
5. Repeat the titration **without any indicator**, using exactly the same volumes — otherwise the salt would be contaminated by the dye.
6. Evaporate the salt solution to crystallisation point, cool, filter, and dry.

EQUATION

$$HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)$$

METHOD 3 — PRECIPITATION (INSOLUBLE SALT)

Mix two soluble salt solutions whose ions combine to give the insoluble target. Step-by-step preparation of silver chloride:

1. Make solutions of silver nitrate and sodium chloride.
2. Mix the two solutions in a beaker. A white precipitate of silver chloride forms instantly.
3. Filter the mixture — the precipitate stays on the filter paper.
4. Wash the precipitate with distilled water (running it through the filter paper) to remove soluble impurities.
5. Place the wet precipitate in a warm oven to dry. **Keep AgCl out of bright light** — it darkens because of photodecomposition.

EQUATION (FULL AND IONIC)

Full: $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

Ionic: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 8.4

💡 **EXAMINER'S TIP**

Always identify the salt's solubility first. Soluble → method 1 or 2 (depending on base solubility); insoluble → method 3. Then pick the right acid: chloride salt → HCl; sulfate → H_2SO_4; nitrate → HNO_3.

CHEMISTRY (4CH1) EXAM FREQUENCY: MEDIUM Syllabus ref: 11.1 / 11.2 / 11.3

CHAPTER 20 / Inorganic Chemistry

Air, Water, and the Atmosphere

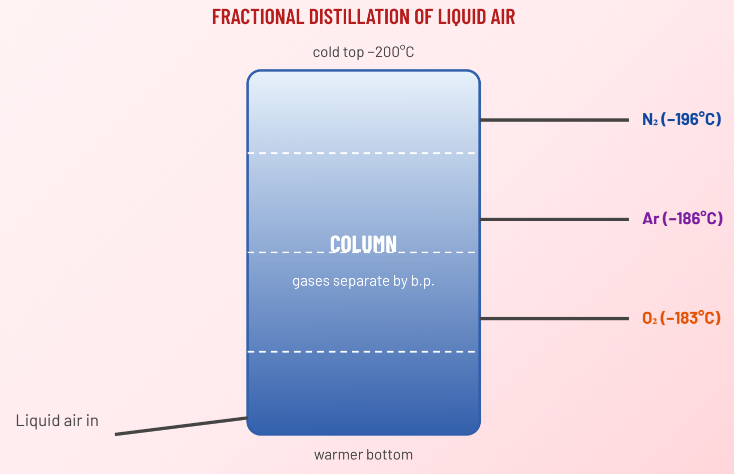
Composition of air, fractional distillation, water purification, and the chemistry of pollution.

COMPOSITION OF CLEAN DRY AIR

GAS	APPROXIMATE %	BOILING POINT (°C)
Nitrogen (N_2)	78	−196
Oxygen (O_2)	21	−183
Argon (Ar)	0.9	−186
Carbon dioxide (CO_2)	~0.04	−78 (sublimes)
Other gases	trace	—

FRACTIONAL DISTILLATION OF LIQUID AIR

FIGURE 20.1



Liquid air enters the column. As it warms, gases boil off in order of boiling point — nitrogen first, argon next, oxygen last.

1. Air is filtered, then passed through sodium hydroxide to remove acidic impurities (CO_2, SO_2, NO_2) and through concentrated H_2SO_4 to dry it.
2. The dry air is compressed and cooled to about −200 °C — everything except helium and neon liquefies.
3. The liquid air is allowed to warm in a fractionating column. Gases boil off in turn: nitrogen first (b.p. −196), then argon (−186), then oxygen (−183).
4. Each fraction is collected and stored in pressurised cylinders.

OXYGEN — PROPERTIES, LAB PREPARATION, USES

- Colourless, odourless, slightly soluble in water.
- Not flammable itself but supports combustion.
- **Lab prep**: $2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$ with MnO_2 catalyst.
- **Test**: a glowing splint relights.
- Uses: oxyacetylene welding; converting iron to steel; medical use in hospitals; high-altitude climbing; diving.

NITROGEN — PROPERTIES AND USES

- Colourless, odourless, tasteless, very unreactive (the N≡N triple bond is hard to break).
- Only with a high temperature or an electric spark does it react with oxygen — e.g. inside a car engine to produce NO and NO_2 (pollutants).
- Uses: making ammonia (Haber process), making nitric acid, producing fertilisers, and as liquid nitrogen for freezing food and biological samples.

AIR POLLUTION — SOURCES AND EFFECTS

POLLUTANT	SOURCE	HARMFUL EFFECT
Carbon monoxide (CO)	Incomplete combustion of fuels in car engines and faulty boilers.	Colourless, odourless, poisonous — binds to haemoglobin and prevents O_2 transport.
Sulfur dioxide (SO_2)	Burning fossil fuels containing sulfur (coal, petroleum).	Causes acid rain; damages lungs; aggravates breathing problems.
Oxides of nitrogen (NO, NO_2)	High-temperature reaction of N_2 + O_2 in car engines and furnaces.	Cause acid rain; brown smog; respiratory disease.
Particulates (soot, PM)	Incomplete combustion (especially diesel).	Trigger asthma and lung disease; soil walls and lungs alike.
Methane (CH_4)	Livestock digestion, rice paddies, landfill, leaks from natural-gas systems.	Powerful greenhouse gas, ~25× the warming impact of CO_2 over 100 years.

THE GREENHOUSE EFFECT

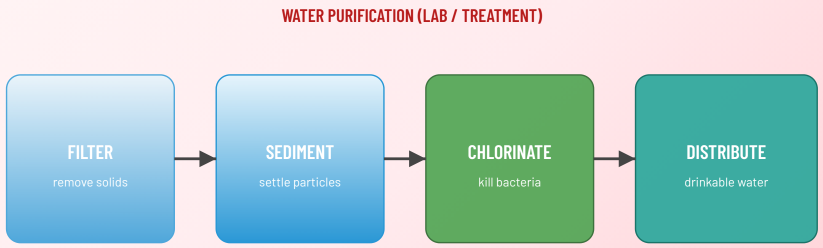
Greenhouse gases (CO_2, CH_4, water vapour) let visible sunlight through but absorb the infrared radiation re-emitted by the Earth, warming the lower atmosphere. Rising CO_2 from fossil fuel use and deforestation is intensifying the natural effect — the cause of **global warming**.

ACID RAIN

Naturally rain has a pH around 5 because dissolved CO_2 makes it slightly acidic. Industrial SO_2 and NO_2 dissolve in atmospheric water to make sulfurous, sulfuric, and nitric acid, dropping the pH below 4. Effects: damage to limestone buildings ($CaCO_3$ reacts with H^+); kills fish in lakes; strips minerals from soil; corrodes metal structures and bridges.

WATER — TREATMENT, TESTS, AND USES

FIGURE 20.2



Public water treatment: filter solids, settle particles, chlorinate to kill bacteria, distribute.

1. **Filtration** through beds of fine sand removes insoluble solids.
2. **Sedimentation** in large tanks lets fine particles settle out.
3. **Chlorination** kills bacteria and microbes.
4. **Distribution** through pressurised pipes.

FIGURE 20.3

Both tests confirm the PRESENCE of water. Pure water also boils at exactly 100°C and freezes at 0°C.

Two chemical tests for water — anhydrous $CuSO_4$ (white → blue) and anhydrous $CoCl_2$ (blue → pink).

CHEMISTRY (4CH1) EXAM FREQUENCY: MEDIUM Syllabus ref: 11.1 / 11.2 / 11.3

TEST	OBSERVATION
Anhydrous copper(II) sulfate ($CuSO_4$)	white → blue (forms $CuSO_4{\cdot}5H_2O$)
Anhydrous cobalt(II) chloride ($CoCl_2$)	blue → pink (forms $CoCl_2{\cdot}6H_2O$)
Boiling point (purity)	Pure water boils at exactly 100 °C and freezes at exactly 0 °C at 1 atm

- **Domestic uses**: drinking, cooking, washing.
- **Industrial uses**: solvent (water dissolves most ionic compounds); coolant in power stations; raw material for the manufacture of ethanol; raw material for the electrolytic production of hydrogen and oxygen.

💡 **EXAMINER'S TIP**

'Pure' in chemistry has a strict meaning: only one type of particle. Bottled mineral water is not pure — it contains dissolved salts. Pure water boils at exactly 100 °C; mineral water boils slightly above. The boiling point test confirms purity, the chemical tests just confirm presence of water.

CHEMISTRY (4CH1) EXAM FREQUENCY: MEDIUM Syllabus ref: 9.2 / 12.x

CHAPTER 21 / The Periodic Table & Industry

Non-metals: Halogens, Hydrogen, Ammonia, and Sulfur

Lab preparations, tests, and the industrial chemistry that follows.

COMMON NON-METAL PROPERTIES

- Mostly gases or low-melting-point solids; many are coloured (S yellow, Br red-brown, I dark grey).
- Poor conductors of electricity (graphite is the celebrated exception).
- Brittle when solid — not malleable.
- Form negative ions, or covalent compounds with each other.
- Their oxides are usually acidic.

THE HALOGENS (GROUP 7)

All exist as diatomic molecules; all are poisonous; reactivity decreases down the group. They undergo halogen displacement — a more reactive halogen displaces a less reactive halogen from its compounds (see Topic 6 for the colour table).

CHLORINE

- Pale green gas, denser than air, soluble in water.
- **Test**: chlorine bleaches damp blue litmus paper white.
- **Industrial production**: electrolysis of brine (concentrated NaCl). Chlorine is released at the anode.
- Uses: bleach, water treatment (kills microbes), making hydrochloric acid and PVC.

HYDROGEN

- The simplest element — one electron, one proton. Diatomic (H_2). Lowest density of any gas.
- **Lab prep**: $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$ — collect over water in a test tube.
- **Industrial production**: from natural gas (cracking) or as a by-product of brine electrolysis (cathode product).
- **Test**: a lit splint at the mouth of the tube produces a sharp **squeaky pop**.
- Uses: making ammonia (Haber); margarine production (hydrogenation of vegetable oils); a non-polluting fuel that produces only water on combustion.

AMMONIA (NH_3)

- Colourless gas with a sharp pungent smell. Less dense than air. Very soluble in water (the dissolved solution is alkaline).
- **Tests**: damp red litmus turns blue; ammonia + HCl gives **white smoke** of NH_4Cl.
- **Industrial production**: Haber process (Topic 18). Conditions 450 °C, 200 atm, iron catalyst, exothermic reversible reaction.

> **AMMONIA → FERTILISER**
>
> $2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$ (ammonium sulfate fertiliser). NPK fertilisers contain Nitrogen, Phosphorus, Potassium — the three macronutrients that limit plant growth.

SULFUR AND SULFURIC ACID

Sulfur is a yellow non-metallic solid, insoluble in water, soluble in some organic solvents. Burns in air with a blue flame to give sulfur dioxide. Manufactured into sulfuric acid by the Contact process (see Topic 18).

> **THREE STAGES**
>
> Stage 1: $S + O_2 \rightarrow SO_2$
> Stage 2: $2SO_2 + O_2 \rightleftharpoons 2SO_3$ (450 °C, V_2O_5 catalyst, ~2 atm)
> Stage 3: SO_3 absorbed in conc. H_2SO_4 (forms oleum), then diluted to give more H_2SO_4.

- Sulfuric acid is the most-produced industrial chemical worldwide.
- Uses: making fertilisers (ammonium sulfate, super-phosphates); car batteries; detergents; refining petroleum; making dyes and pigments.

NOBLE GASES (GROUP 0)

Helium, neon, argon, krypton, xenon, radon. Monoatomic. Full outer shells → almost completely unreactive.

- **Helium**: balloons, airships (much safer than hydrogen).
- **Neon**: red advertising signs, lasers.
- **Argon**: filling tungsten light bulbs to stop the filament burning; shielding gas in arc welding.
- **Krypton**: photographic flash, lasers (eye surgery), car headlamps.
- **Xenon**: high-intensity lighthouse and projector lamps.

💡 **TESTS YOU MUST MEMORISE**

Hydrogen: lit splint → squeaky pop. **Oxygen**: glowing splint relights. **Carbon dioxide**: turns lime water milky. **Chlorine**: bleaches damp litmus paper. **Ammonia**: damp red litmus turns blue. These five appear in nearly every paper.

CHEMISTRY (4CH1) | **EXAM FREQUENCY: HIGH** | Syllabus ref: 12.1 / 12.2

CHAPTER 22 / Organic Chemistry

Organic Chemistry

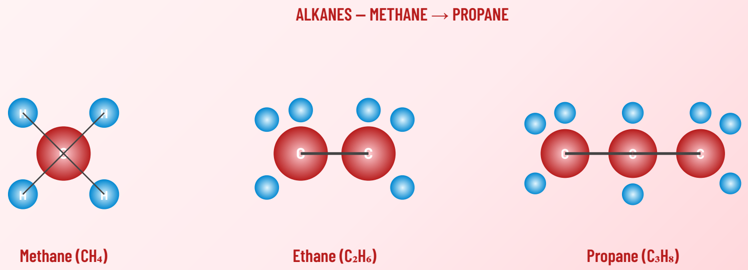
The carbon-based world: alkanes, alkenes, alcohols, carboxylic acids, esters, polymers.

Carbon atoms can form four covalent bonds, including bonds with other carbon atoms. They build chains, rings, and branches — over a million different compounds. We organise them into **homologous series**: families with the same general formula and gradually changing properties.

ALKANES (SATURATED; ONLY SINGLE BONDS)

General formula C_nH_{2n+2}. Found in crude oil; used as fuels (methane, propane, octane). Two main reactions: combustion in oxygen, and substitution with halogens in UV light.

FIGURE 22.1



First three alkanes: methane (CH_4), ethane (C_2H_6), propane (C_3H_8).

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 12.1 / 12.2

COMBUSTION

Complete: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ (clean blue flame). Incomplete: $2CH_4 + 3O_2 \rightarrow 2CO + 4H_2O$ (yellow smoky flame, gives toxic CO and soot). Always pair the equation to the air supply.

SUBSTITUTION WITH HALOGENS

$CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$ (UV light needed). One H is replaced at a time.

ALKENES (UNSATURATED; ONE C=C BOND)

General formula C_nH_{2n}. The double bond can 'open up' to add extra atoms — alkenes are far more reactive than alkanes. **Bromine water test**: the orange-brown colour disappears when shaken with an alkene (Br_2 adds across the double bond).

ADDITION REACTIONS OF ETHENE (C_2H_4)

+ Br_2 (bromine water): $C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$ — orange → colourless
+ H_2 (Ni catalyst, 200 °C): $C_2H_4 + H_2 \rightarrow C_2H_6$ — hydrogenation, used to make margarine
+ H_2O (300 °C, 60 atm, H_3PO_4 catalyst): $C_2H_4 + H_2O \rightarrow C_2H_5OH$ — industrial route to ethanol

ALCOHOLS

Functional group **—OH**. General formula $C_nH_{2n+1}OH$. Ethanol (C_2H_5OH) is the most-tested. Two production routes:

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 12.1 / 12.2

CHAPTER 22 / Organic Chemistry · continued

ROUTE	PROCESS	CONDITIONS	PROS / CONS
Fermentation	Sugars + yeast → ethanol + CO_2	30–40 °C, anaerobic	Renewable; slow; gives dilute solution that needs distilling
Hydration of ethene	$C_2H_4 + H_2O \rightarrow C_2H_5OH$	300 °C, 60 atm, H_3PO_4	Continuous; pure ethanol; ethene from finite crude oil

Ethanol burns cleanly: $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$. Used as a solvent, a fuel ("E10" petrol), and in alcoholic drinks.

CARBOXYLIC ACIDS

Functional group **—COOH**. Examples: methanoic acid (HCOOH), ethanoic acid (CH_3COOH, vinegar). Weak acids — they react with metals, bases, and carbonates just like strong acids do, but more slowly. Made by oxidising the matching alcohol (with acidified $KMnO_4$ or air over a catalyst).

ESTERS

Made by warming a carboxylic acid with an alcohol in the presence of a few drops of concentrated H_2SO_4 (catalyst). The reaction is reversible. Esters smell sweet and fruity — used in perfumes, food flavourings.

> **ESTERIFICATION**
>
> $CH_3COOH + C_2H_5OH \rightleftharpoons CH_3COOC_2H_5 + H_2O$ — ethanoic acid + ethanol gives ethyl ethanoate (smells like nail polish).

POLYMERS

Addition polymers form when many alkene monomers join up by opening their double bonds. The repeat unit has the same atoms as the monomer; only the bonds change.

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: 12.1 / 12.2

CHAPTER 22 / Organic Chemistry · continued

- **Poly(ethene)**: $n \times CH_2{=}CH_2 \rightarrow -(CH_2{-}CH_2)_n-$. Plastic bags, bottles.
- **Poly(propene)**: from propene. Crates, ropes, kitchenware.
- **PVC**: poly(chloroethene). Pipes, window frames, cable insulation.
- **PTFE**: poly(tetrafluoroethene). Non-stick coatings.

Condensation polymers form when monomers with two functional groups join, releasing a small molecule (often water) at each link. Examples: nylon (a polyamide) and Terylene (a polyester).

> ⚠ **COMMON MISTAKE**
>
> Don't confuse *addition* polymerisation (alkenes → one product, no other molecule released) with *condensation* polymerisation (releases small molecules like water).

> 💡 **EXAMINER'S TIP**
>
> Recycling vs. landfill of plastics is a frequent essay topic. Addition polymers are non-biodegradable, take centuries to break down, and burn to release toxic gases. The mark scheme rewards naming a specific polymer (e.g. PVC produces HCl) and one realistic alternative (recycling, biodegradable polymers, reduced consumption).

CHEMISTRY (4CH1) **EXAM FREQUENCY: HIGH** Syllabus ref: 4.1 / 4.2 / 4.3

CHAPTER 23 / Quantitative Chemistry

Moles and Calculations

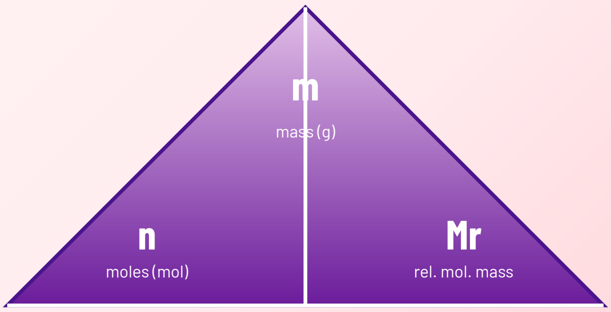
Chemistry by the numbers — moles, masses, gas volumes, concentrations, yields.

📖 **THE MOLE**

A **mole** (mol) is the amount of substance that contains **6.02×10^{23}** particles — Avogadro's number. One mole of any substance has a mass equal to its **relative formula mass (Mr)** in grams.

FIGURE 23.1

THE MOLE TRIANGLE — MASS · MOLES · Mr



n = m / Mr · m = n × Mr · Mr = m / n

The mole triangle — cover the quantity you want; what is left tells you the formula.

THE THREE FORMS

n = m / Mr · m = n × Mr · Mr = m / n (n in mol, m in g, Mr in g/mol)

WORKED EXAMPLE 1 — MASS TO MOLES

How many moles in 8.0 g of NaOH? Mr(NaOH) = 23 + 16 + 1 = 40. n = 8.0 / 40 = **0.20 mol**.

WORKED EXAMPLE 2 — REACTING MASSES

What mass of MgO is formed when 6.0 g of Mg burns completely? Equation: $2Mg + O_2 \rightarrow 2MgO$. Mr(Mg) = 24, Mr(MgO) = 40. n(Mg) = 6.0 / 24 = 0.25 mol. From the 2:2 ratio, n(MgO) = 0.25 mol. m(MgO) = 0.25 × 40 = **10.0 g**.

VOLUME OF GAS (ROOM TEMPERATURE AND PRESSURE)

MOLAR GAS VOLUME

At RTP (25 °C, 1 atm), **1 mole of any gas occupies 24 dm^3** (= 24 000 cm^3). Volume = n × 24 dm^3.

What volume of CO_2 (at RTP) is given off when 50 g of $CaCO_3$ decomposes? Mr($CaCO_3$) = 100. n = 50/100 = 0.50 mol. From $CaCO_3 \rightarrow CaO + CO_2$, n($CO_2$) = 0.50 mol. V = 0.50 × 24 = **12 dm^3**.

CONCENTRATION OF SOLUTIONS

TWO USEFUL EQUATIONS

concentration (mol/dm^3) = moles ÷ volume (in dm^3)
concentration (g/dm^3) = mass ÷ volume (in dm^3)
To convert: c (g/dm^3) = c (mol/dm^3) × Mr.

WORKED EXAMPLE 3 — TITRATION

25.0 cm^3 of 0.100 mol/dm^3 NaOH is exactly neutralised by 21.5 cm^3 of dilute H_2SO_4. What is the concentration of the acid?

CHEMISTRY (4CH1) | EXAM FREQUENCY: HIGH | Syllabus ref: 4.1 / 4.2 / 4.3

1. Equation: $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$.
2. Moles of NaOH = 0.100 × (25.0/1000) = 2.50×10^{-3} mol.
3. Mole ratio NaOH : H_2SO_4 = 2 : 1, so moles of H_2SO_4 = 2.50×10^{-3} / 2 = 1.25×10^{-3} mol.
4. Concentration of H_2SO_4 = 1.25×10^{-3} / (21.5/1000) = **0.0581 mol/dm^3**.

PERCENTAGE YIELD AND PERCENTAGE PURITY

TWO FINAL EQUATIONS

```
% yield = (actual yield ÷ theoretical yield) × 100
% purity = (mass of pure substance ÷ mass of impure sample) × 100
```

Calculate theoretical yield from the equation, then divide the mass actually obtained by it. A yield below 100% can be due to side reactions, losses during transfer, or incomplete reaction.

EXAMINER'S TIP

Mole calculations always work in three steps: (1) find moles of the known, (2) use the equation ratio to get moles of the unknown, (3) convert moles back to mass, volume, or concentration. Show every step — method marks are awarded even if the final number is wrong.

⚠ **COMMON MISTAKE**

Convert volumes to dm^3 before plugging into the concentration formula. 250 cm^3 = 0.250 dm^3, not 250.

CHEMISTRY (4CH1) EXAM FREQUENCY: HIGH Syllabus ref: all

APPENDIX / END OF NOTES

Synopsis, Glossary & Examiner's Tips

A final pass over the most exam-critical concepts from across the 23 chapters.

SYNOPSIS

Chemistry at IGCSE level builds from the simplest idea — that everything is made of **tiny particles** — into the rich behaviour of matter, atoms, ions, and reactions. The first half of the course (states of matter through bonding) trains your particle thinking. The middle (acids, oxides, redox, electrolysis, reactivity) trains your **chemical intuition**: which species attack which, who donates electrons, who gains them. The second half (rates, equilibria, salts, the Earth's atmosphere, organic chemistry, moles) gives you the tools to **predict and quantify** what happens. Master the language — *ions*, *oxidation*, *reduction*, *concentration*, *equilibrium*, *functional group* — and the exam reads itself.

EXAMINER'S TIPS (THE MARKS-WINNING SHORTLIST)

- **State symbols are free marks** — every full equation should include (s), (l), (g) or (aq). Examiners deduct for missing symbols when they are required.
- **Balance every equation**. If your answer says “$Mg + O_2 \rightarrow MgO$” it is wrong — it must be $2Mg + O_2 \rightarrow 2MgO$.
- **Define exactly what is asked**. “An acid is a proton donor” is the mark-scheme phrase. “An acid is sour” is not.
- **Show working in calculations**. Even a single line of substitution earns a method mark when the final number is wrong.
- **Distinguish observations from conclusions**. “Bubbles formed” is observation; “hydrogen gas was produced” is the conclusion.
- **Energy-level diagrams**: arrow points *down* for exothermic, *up* for endothermic. Always label reactants, products, and ΔH.
- **Electrolysis products**: at the cathode (negative), the lower-reactivity ion or hydrogen wins. At the anode (positive), halides go first, then OH^- gives oxygen.

COMPACT GLOSSARY

Atom
The smallest particle of an element that retains its chemical identity.

Element
A pure substance made of one type of atom.

Compound
A substance made of two or more elements chemically combined in fixed proportions.

Mixture
Two or more substances physically together but not chemically combined.

Mole
The amount of substance containing 6.02×10^{23} particles.

Mr
Relative molecular mass — the sum of the relative atomic masses in a molecule.

Ion
A charged particle formed when an atom or group gains or loses electrons.

Oxidation
Loss of electrons (or gain of oxygen).

Reduction
Gain of electrons (or loss of oxygen).

Catalyst
A substance that speeds up a reaction without being used up.

Equilibrium
A state in a closed system where forward and backward reactions occur at equal rates.

Electrolysis
The breakdown of an ionic compound (molten or in solution) by an electric current.

Salt
An ionic compound formed when an acid reacts with a base, metal or carbonate.

Acid
A proton (H^+) donor; turns blue litmus red; pH below 7.

Alkali
A soluble base; releases OH^- in water; pH above 7.

Hydrocarbon
A compound containing only carbon and hydrogen.

Homologous series
A family of organic compounds with the same general formula and gradually changing properties.

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STUDENT NAME	**STUDENT ID**
SUBJECT Chemistry (4CH1)	**SUBJECT CODE** 4CH1
CURRICULUM Edexcel Pearson IGCSE / O Level	**ACADEMIC SESSION**

DECLARATION OF AUTHENTICITY

I hereby declare that this study notes document is my own original work prepared under the guidance of my instructor at Neuratech Academy. All content contained herein is accurate to the best of my knowledge and has been prepared with academic integrity in accordance with the institution's code of conduct and scholarly standards.

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DATE