Pearson Edexcel Chemistry

EDEXCEL IGCSE 9-1 · 4CH1

A comprehensive revision guide covering all four major topics: Principles of Chemistry, Inorganic Chemistry, Physical Chemistry, and Organic Chemistry — plus practical skills assessed in written examinations.

01

Principles of Chemistry

States of matter, atomic structure, bonding, electrolysis

02

Inorganic Chemistry

Groups 1 & 7, reactivity series, acids, salts, chemical tests

03

Physical Chemistry

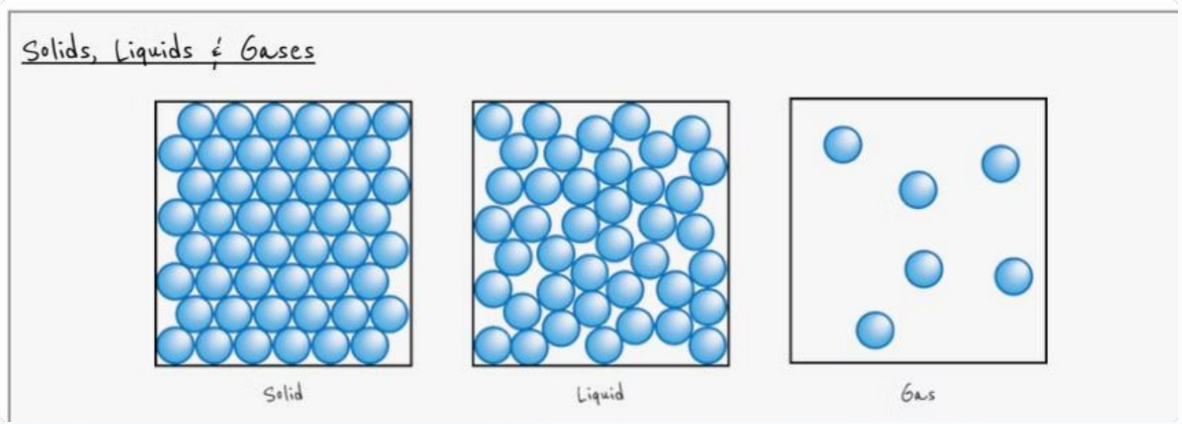
Energetics, rates of reaction, equilibria

04

Organic Chemistry

Crude oil, alkanes, alkenes, alcohols, polymers

States of Matter



Solid

- Particles arranged regularly and packed closely together
- Vibrate in fixed positions
- Strong forces between particles
- Small amount of kinetic energy

Liquid

- Particles mostly touching with some gaps
- Particles move about at random
- Medium forces between particles
- Moderate amount of kinetic energy

Gas

- Particles move at random and quickly
- Particles are far apart
- Weak forces between particles
- High kinetic energy; collide with each other and sides of container

State Changes

- Solid → liquid = **melting**
- Liquid → gas = **boiling / evaporating**
- Gas → liquid = **condensing**
- Liquid → solid = **freezing**

Particle Behaviour During Changes

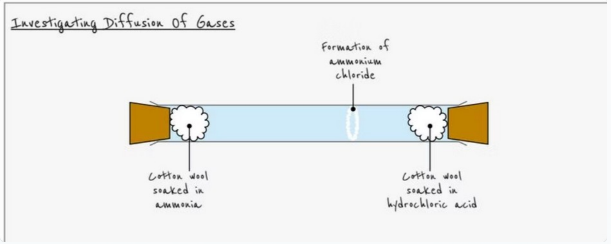
- **Cooling (condensing/freezing):** particles lose kinetic energy, move closer together, move slower and less randomly
- **Heating (boiling/melting):** particles gain kinetic energy, move further apart, move quicker and more randomly
- **Melting ice:** forces of attraction overcome; particles vibrate more and spread apart

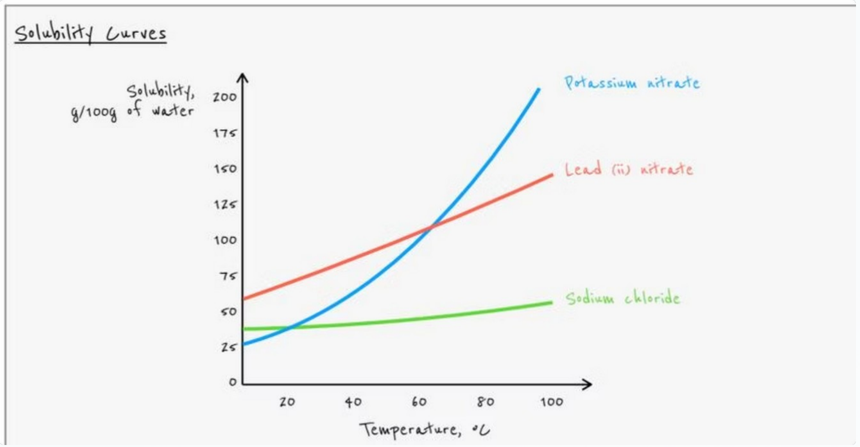
Diffusion & Solubility

Diffusion

Definition: Net movement of particles from an area of high concentration to low concentration.

Experimental demonstration: Cotton wool soaked in NH_3 and HCl placed at opposite ends of a glass tube. A white ring of ammonium chloride forms *closer to the HCl end* because NH_3 has a lower M_r and diffuses faster.





Solubility Key Terms

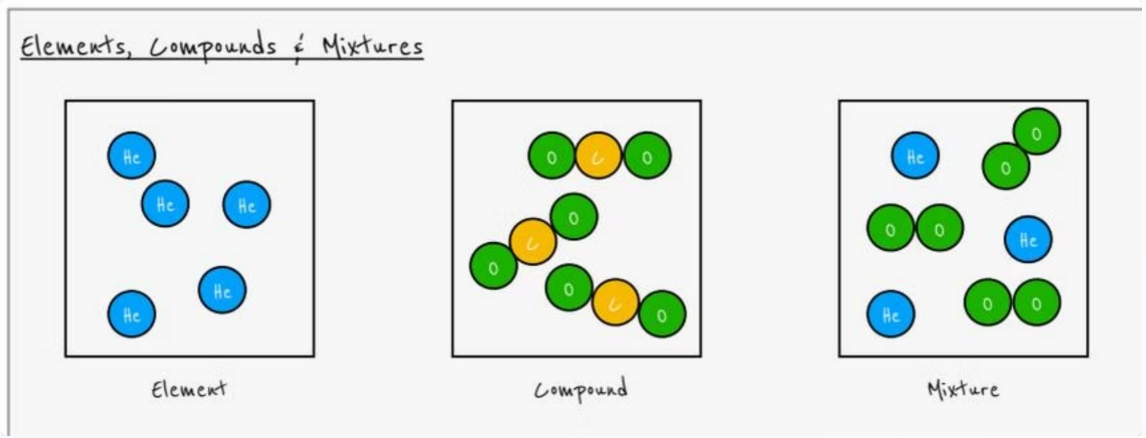
- **Solvent** – liquid in which a solute dissolves
- **Solute** – solid which dissolves in a solvent
- **Solution** – mixture of solute and solvent
- **Saturated solution** – no more solute can dissolve
- **Solubility** – mass of solute dissolving in 100 g of solvent at a given temperature to form a saturated solution

Solubility Trends

- As temperature increases → solubility of **solids increases**
- As pressure increases → solubility of **gases increases**
- As temperature increases → solubility of **gases decreases**

Solubility (g/100g) = **100 × mass of solute / mass of evaporated solvent**

Elements, Compounds & Mixtures



Element

Contains one type of atom only. Cannot be split into anything simpler by chemical means.

Examples: Iron, Lead, Sulfur, Nitrogen, Oxygen

Compound

Made up of two or more elements **chemically combined**.

Examples: Calcium carbonate, Ammonia, Carbon dioxide, Water, Iron sulfide

Mixture

Made up of two or more elements **NOT chemically bonded** together.

Examples: Honey, Air, Sea water, Blood, Soup

Pure Substances vs Mixtures

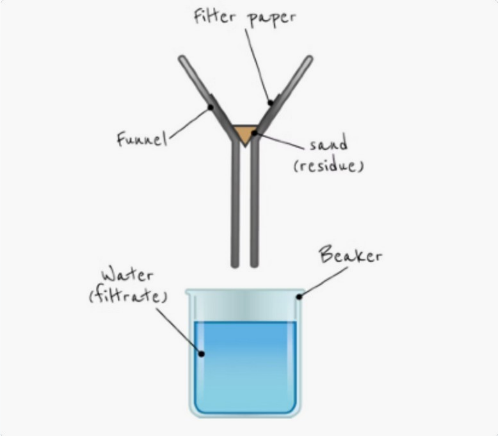
- **Pure substance:** contains one type of material only — has fixed melting and boiling points (e.g. pure water boils at exactly 100°C)
- **Mixture:** melts and boils over a range of temperatures

Separation Techniques

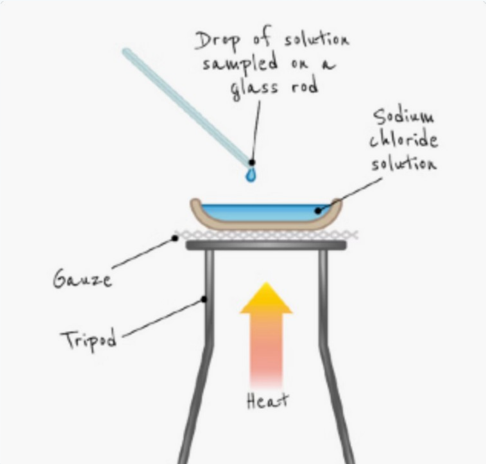
- **Filtration** – separates insoluble solute from solvent
- **Evaporation / Distillation** – separates soluble solute from solvent
- **Fractional distillation** – separates liquids of different boiling points
- **Simple distillation** – separates pure water from sea water
- **Separating funnel** – separates immiscible liquids (e.g. petrol and water)
- **Paper chromatography** – separates dyes/inks with different solubilities

Separation Techniques

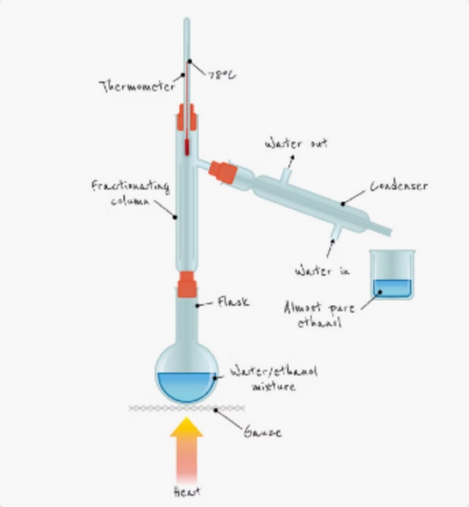
Filtration



Evaporation



Fractional Distillation



Simple Distillation

Atomic Structure

Key Definitions

Atom: smallest particle of a chemical element that can exist.

Molecule: two or more atoms chemically bonded together.

Atomic number: number of protons (= number of electrons in a neutral atom).

Mass number / Nucleon number: total number of protons + neutrons.

Number of neutrons = mass number − atomic number.

Isotopes: atoms of the same element with the same number of protons but different number of neutrons.

Subatomic Particles

Particle	Relative Charge	Relative Mass
Proton	+1	1
Neutron	0	1
Electron	−1	1/1836

Relative Atomic Mass (A_r)

The ratio of the average mass of one atom of an element compared with 1/12th of the mass of an atom of carbon-12.

$$A_r = \frac{(\%\text{ isotope } 1 \times \text{mass}_1) + (\%\text{ isotope } 2 \times \text{mass}_2) + \ldots}{100}$$

Worked Example: Chlorine: 75% Cl-35 and 25% Cl-37. A_r = [(75×35)+(25×37)]/100 = **35.5**

The Periodic Table

Layout

- Elements arranged in order of **atomic number**
- Vertical columns = **groups**; horizontal rows = **periods**
- Metals found left of stepped line; non-metals to the right
- **Period number** = number of electron shells (e.g. Ca: 2.8.8.2 → period 4)
- **Group number** = number of electrons in outer shell (e.g. F has 7 outer electrons → group 7)

Electronic Configurations

- Na = 2.8.1 · O = 2.6 · Cl = 2.8.7 · Mg = 2.8.2

Metals vs Non-Metals

Metals	Non-Metals
Good conductors of heat & electricity	Poor conductors
Shiny, malleable, ductile, sonorous	Brittle
Form positive ions; ionic bonding	Form negative ions; ionic & covalent bonding
Form basic oxides	Form acidic oxides
Solids at room temperature (except Hg)	Various states

Elements in the same group have **similar chemical properties** because they have the same number of electrons in their outer shell. Noble gases (group 0) are unreactive because they have a **full outer shell**.

Chemical Formulae, Equations & Mole Calculations

Balancing Equations

Only change the large number in **front** of each element/compound — never change the subscript numbers.

> **Worked Example:** $Cu + O_2 \rightarrow CuO$
> Tally: Cu: 1 vs 1; O: 2 vs 1 → add extra CuO → Cu: 1 vs 2 → add extra Cu
> **Answer: $2Cu(s) + O_2(g) \rightarrow 2CuO(s)$**

The Mole

- Amount of substance containing the same number of units as atoms in 12 g of carbon-12
- **Avogadro constant** = 6.02×10^{23}
- Moles = mass / M_r
- % by mass = (total mass of element / mass of compound) × 100

> **Examples:** M_r of $MgCO_3$ = 24+12+(3×16) = 84. Moles in 54 g H_2O = 54/18 = 3. Mass of 0.2 mol $CaCO_3$ = 0.2×100 = 20 g. % Mg in MgO = 24/40 × 100 = 60%

Concentration & Volume

- Units of concentration: mol/dm^3
- **Moles = Concentration × Volume** (volume in dm^3; 1 dm^3 = 1000 cm^3)

> **Example:** 25 cm^3 of 2 mol/dm^3 HCl → moles = 2 × 0.025 = **0.05 mol**

Gas Volumes

1 mole of gas occupies **24 dm^3** at RTP (24,000 cm^3)

> **Example:** 60 g ethane (C_2H_6, M_r=30) → 2 mol → 4 mol CO_2 produced → volume = 4 × 24 = **96 dm^3**

Percentage Yield

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

> **Example:** Actual yield 1.6 g, theoretical 2.0 g → % yield = 1.6/2.0 × 100 = **80%**

Reacting Masses, Empirical Formula & Water of Crystallisation

Reacting Masses (Worked Example)

4.5 g HCl reacted with $Ca(OH)_2$. Find mass of $CaCl_2$ formed.

$2HCl + Ca(OH)_2 \rightarrow CaCl_2 + H_2O$

- M_r HCl = 36.5; moles HCl = 4.5/36.5 = 0.1233
- Moles $CaCl_2$ = 0.1233/2 = 0.0616 (stoichiometry 2:1)
- M_r $CaCl_2$ = 111; mass = 111 × 0.0616 = **6.84 g**

Limiting Reagents (Worked Example)

$Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2H_2O + 2NO_2$

3.2 g Cu (M_r=63.5) → 0.05 mol Cu → requires 0.05×4 = 0.2 mol HNO_3. We have 0.4 mol HNO_3 → **HNO_3 is in excess; Cu is limiting.**

Empirical Formula (Worked Example)

Compound contains 5.85 g K, 2.10 g N, 4.8 g O.

	K	N	O
Mass	5.85	2.10	4.80
M_r	39	14	16
Moles	0.15	0.15	0.30
Ratio	1	1	2

Answer: KNO_2

Water of Crystallisation (Worked Example)

11.25 g $CuSO_4 \cdot xH_2O$ heated → 7.19 g. Mass H_2O = 4.06 g.

- Moles $CuSO_4$ (M_r=159.5) = 7.19/159.5 = 0.04508
- Moles H_2O (M_r=18) = 4.06/18 = 0.2256
- Ratio = 0.2256/0.04508 = 5 → **x = 5 ($CuSO_4 \cdot 5H_2O$)**

Ionic Bonding

Ion Formation

- Atom **loses** electrons → positive ion (cation)
- Atom **gains** electrons → negative ion (anion)
- Groups 1–3: charge = group number (e.g. Mg^{2+})
- Groups 5, 6, 7: charge = 8 − group number (e.g. N^{3-})

Key Ions to Learn

Cations: H^+, Ag^+, Cu^{2+}, Fe^{2+}, Fe^{3+}, Pb^{2+}, Zn^{2+}, NH_4^+

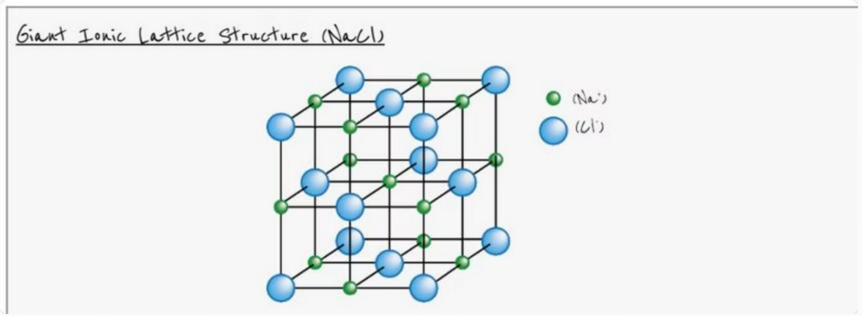
Anions: OH^-, NO_3^-, CO_3^{2-}, SO_4^{2-}

Ionic Bond

Electrostatic force of attraction between **oppositely charged ions**. Forms between a metal and a non-metal.

Drag & Drop Method: Cross the charges to get subscripts. e.g. $Fe^{3+} + SO_4^{2-} \rightarrow Fe_2(SO_4)_3$; $Ca^{2+} + Cl^- \rightarrow CaCl_2$

Properties of Giant Ionic Lattices



- **High melting/boiling points:** strong electrostatic forces between oppositely charged ions require lots of energy to break
- **Do NOT conduct when solid:** ions held tightly in fixed positions, not free to move
- **Conduct when molten/dissolved:** ions are free to move and carry electric charge
- **Brittle:** applying force causes like charges to align and repel → lattice breaks apart

Covalent Bonding

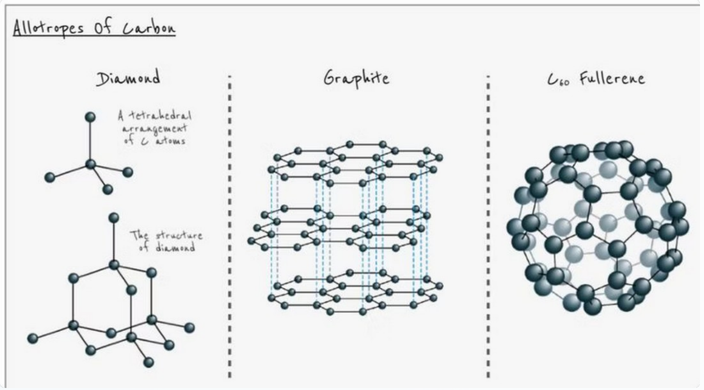
Covalent Bond

- **Basic:** a pair of electrons shared between two atoms
- **Detailed:** strong electrostatic forces of attraction between nuclei (positive) and shared pair of electrons (negative)
- Forms between **two non-metals**

Simple Molecular Substances

- Small, covalently bonded molecules (e.g. H_2O, CO_2)
- **Low melting points:** weak intermolecular forces, little energy needed to overcome
- **Boiling point increases with M_r:** greater M_r = greater intermolecular forces to overcome
- **Do not conduct electricity:** no overall electric charge, no free electrons

Allotropes of Carbon



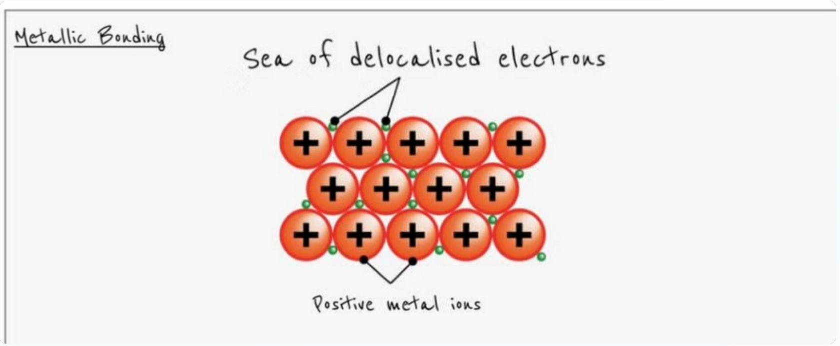
Property	Diamond	Graphite	C_{60}
Melting point	Very high	Very high	Low
Conducts electricity?	No	Yes	No
Bonds per C atom	4	3	3

- **Diamond:** each C bonded to 4 others — no free electrons, very hard, giant covalent structure
- **Graphite:** each C bonded to 3 others — 4th electron free to move (conducts); layers slide (lubricant)
- **C_{60}:** simple molecular structure — weak intermolecular forces; 4th electron stays within molecule (does not conduct)

Metallic Bonding & Bonding Summary

Metallic Bond

Electrostatic forces of attraction between **positive metal ions** and a **'sea' of delocalised electrons**. Structure: regular arrangement of positive ions in a giant metallic lattice.

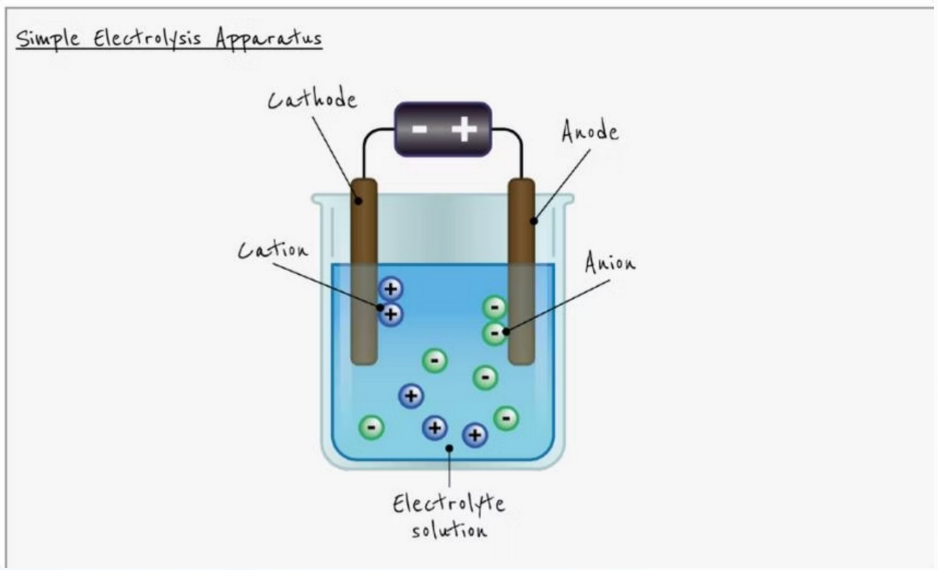


- **Good conductors of heat & electricity:** delocalised electrons free to move, transferring energy throughout structure
- **High melting/boiling points:** strong electrostatic forces require lots of energy to break
- **Malleable & ductile:** layers of ions can slide over each other

Bonding Summary

Structure	Simple Covalent	Giant Covalent	Giant Ionic	Giant Metallic
Made of	Non-metals	Non-metals	Metal + Non-metal	Metals
Examples	C_{60}, CH_4, NH_3	Diamond, Graphite, SiO_2	NaCl, MgO	Fe, Zn, Al
Melting point	Low	High	High	High
Conducts (solid)?	No	No (except graphite)	No	Yes
Conducts (liquid)?	No	No	Yes	Yes

Electrolysis



Key Definitions

- **Electrolysis:** breaking down of a substance using electricity
- **Anode:** positive electrode; **Cathode:** negative electrode (**PANC**)
- **Anion:** negative ion (gained electrons); **Cation:** positive ion (lost electrons)
- **OIL RIG:** Oxidation Is Loss; Reduction Is Gain (of electrons)
- Electrodes usually made of inert graphite

Ionic Solutions — Rules

- **Cathode (−):** H_2 produced unless metal cation is *less reactive* than hydrogen → metal deposited instead
- **Anode (+):** O_2 produced unless halide ions (Cl^-, Br^-, I^-) present → halogen produced instead

Worked Examples

Aqueous NaCl: Cathode: $2H^+(aq) + 2e^- \rightarrow H_2(g)$ (H less reactive than Na). Anode: $2Cl^-(aq) - 2e^- \rightarrow Cl_2(g)$ (halide present). Remaining: NaOH solution. Uses: H_2 as fuel; Cl_2 for bleach/pools; NaOH for paper/bleach.

Aqueous $CuSO_4$ (impure Cu anode): Cathode: $Cu^{2+} + 2e^- \rightarrow Cu$ (gains pure Cu, increases in mass). Anode: $Cu - 2e^- \rightarrow Cu^{2+}$ (loses Cu & impurities, decreases in mass).

Molten $PbBr_2$: Cathode: $Pb^{2+} + 2e^- \rightarrow Pb$. Anode: $2Br^- - 2e^- \rightarrow Br_2$.

Molten Ionic Compounds — Rules

- Cathode: solid metal forms from cations
- Anode: negative ion forms its element

Group 1 — Alkali Metals

Physical Properties

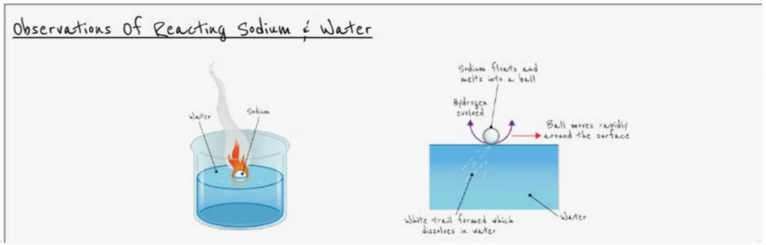
- Soft (can be cut with a knife)
- Low melting and boiling points; low density (float on water)
- Shiny (tarnish when exposed to air)
- Stored in oil — very reactive

Chemical Properties

- React with water → metal hydroxide + hydrogen: **$2Na + 2H_2O \rightarrow 2NaOH + H_2$**
- React with oxygen → metal oxides: **$4Li + O_2 \rightarrow 2Li_2O$**
- Form ionic compounds; react with halogens
- Universal indicator turns **purple** (hydroxide produced is alkaline)
- Similar properties because all have **1 electron in outer shell**
- Reactivity **increases down the group**: outer electron further from nucleus, more shielded, more easily lost

Reactions with Water

Metal	Observations
Lithium	Fizzes gently; moves across surface; disappears; red flame with O_2
Sodium	Melts into ball; fizzes; moves across surface; white trail; yellow flame with O_2
Potassium	Melts; fizzes vigorously/explodes; sparks; burns with lilac flame; white trail



Francium (predicted): Soft, low mp/bp, low density, shiny. Reactions with air and water will be more violent than any other group 1 metal.

Group 7 — Halogens

Physical Properties at Room Temperature

Halogen	State	Colour
Fluorine	Gas	Yellow
Chlorine	Gas	Green
Bromine	Liquid	Red-brown
Iodine	Solid	Grey (purple vapour when heated)
Astatine (predicted)	Solid	Dark grey/black

Trends Down Group 7

- Colour gets darker; melting/boiling point increases
- **Reactivity decreases** — outer shell further from nucleus, more shielded, harder to gain an electron
- All poor conductors of heat and electricity

Reaction with Hydrogen

Hydrogen halides formed (acidic, highly poisonous, very soluble in water). e.g. $H_2(g) + Br_2(g) \rightarrow 2HBr(g)$; $HCl(g) \rightarrow HCl(aq)$ (hydrochloric acid)

Displacement Reactions

A more reactive halogen displaces a less reactive halide from its compound.

	KCl	KBr	KI
Cl_2	—	$Cl_2+2KBr\rightarrow2KCl+Br_2$ (orange)	$Cl_2+2KI\rightarrow2KCl+I_2$
Br_2	No reaction	—	$Br_2+2KI\rightarrow2KBr+I_2$ (brown)
I_2	No reaction	No reaction	—

Ionic equation example: $Cl_2(aq) + 2Br^-(aq) \rightarrow Br_2(aq) + 2Cl^-(aq)$ (K^+ are spectator ions, removed)

A halogen cannot displace itself or a more reactive halogen — hence the "no reaction" and "—" entries.

Gases in the Atmosphere

Composition of Air

78%

Nitrogen

21%

Oxygen

0.96%

Argon

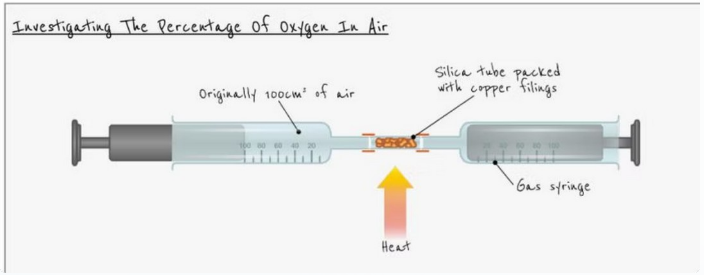
0.04%

Carbon Dioxide

Demonstrating % Oxygen in Air

- **Copper method:** 100 cm^3 air passed over heated copper → copper reacts with O_2 forming black CuO → volume decreases to ~80 cm^3. Cool before measuring (gas expands when heated).
- **Iron filings:** placed in burette full of air; iron reacts with O_2; water moves into burette. % O_2 = (initial − final height)/initial × 100
- **Phosphorus:** lit phosphorus reacts with O_2 ($4P + 5O_2 \rightarrow P_4O_{10}$); volume decreases; water moves into tube.

Combustion Reactions



Substance	Observations	Equation
Magnesium	Bright white light; white solid (MgO)	$2Mg + O_2 \rightarrow 2MgO$; $MgO + H_2O \rightarrow Mg(OH)_2$ (alkaline)
Hydrogen	Pale blue flame; water forms	$2H_2 + O_2 \rightarrow 2H_2O$
Sulfur	Blue flame; colourless poisonous SO_2	$S + O_2 \rightarrow SO_2$; $SO_2 + H_2O \rightarrow H_2SO_3$ (acidic)

Thermal Decomposition of Carbonates

- $CuCO_3 \rightarrow CuO + CO_2$ (green → black)
- $CaCO_3 \rightarrow CaO + CO_2$
- CO_2 is a greenhouse gas → global warming → ice caps melt → sea levels rise → floods, habitat loss

Reactivity Series

Displacement Reactions

A more reactive element pushes out a less reactive element from its compound.

- $Mg + CuO \rightarrow MgO + Cu$ (Mg more reactive than Cu)
- $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$ (Zn more reactive than Cu)

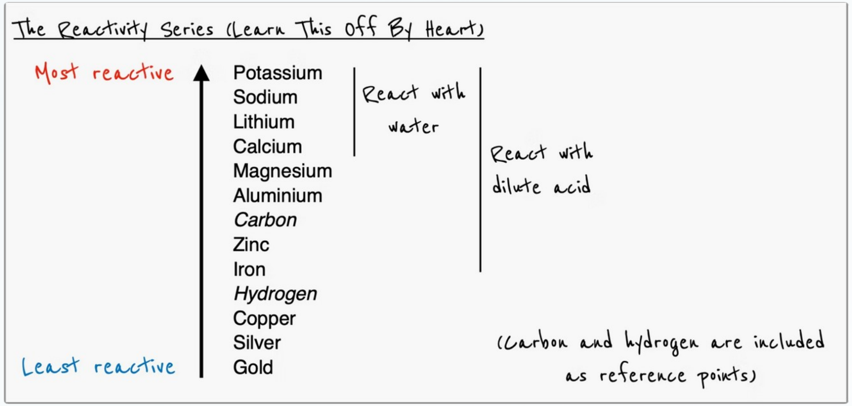
Rusting & Prevention

- **Conditions needed:** water AND oxygen
- $4Fe + 3O_2 + nH_2O \rightarrow 2Fe_2O_3{\cdot}nH_2O$
- **Barrier methods:** paint, grease, oil — prevent water/oxygen reaching iron; if damaged, iron rusts
- **Galvanising:** coat iron in zinc; zinc reacts with oxygen first (Zn^{2+} forms instead of Fe)
- **Sacrificial protection:** more reactive metal attached; undergoes oxidation in preference to iron

Redox

- **Oxidation:** gain of oxygen / loss of electrons (OIL)
- **Reduction:** loss of oxygen / gain of electrons (RIG)
- **Oxidising agent:** causes oxidation; is itself reduced
- **Reducing agent:** causes reduction; is itself oxidised

Redox example: $Mg(s) + 2Ag^+(aq) \rightarrow Mg^{2+}(aq) + 2Ag(s)$. Mg is oxidised (loses e^-) = reducing agent. Ag^+ is reduced (gains e^-) = oxidising agent.



The Reactivity Series

Most → Least reactive: Potassium · Sodium · Lithium · Calcium · Magnesium · Aluminium · *Carbon* · Zinc · Iron · *Hydrogen* · Copper · Silver · Gold

Metals above hydrogen react with dilute acid. K, Na, Li, Ca are too reactive to use with acids safely.

Extraction & Uses of Metals

Iron — Blast Furnace

Iron is less reactive than carbon → extracted by **reduction with coke (carbon)**. Strong, malleable, cheap to extract. Uses: bridges, car bodies.

Aluminium — Electrolysis

More reactive than carbon → cannot use Blast Furnace. Electrolysis used instead. **Expensive** due to high electricity cost and carbon anode burning away ($C + O_2 \rightarrow CO_2$). Low density; good conductor. Uses: aeroplanes, cables, saucepans.

Cathode: $Al^{3+} + 3e^- \rightarrow Al$. Anode: $2O^{2-} - 4e^- \rightarrow O_2$

Copper

Malleable, ductile, good conductor of heat and electricity. Uses: water pipes, wires.

Alloys

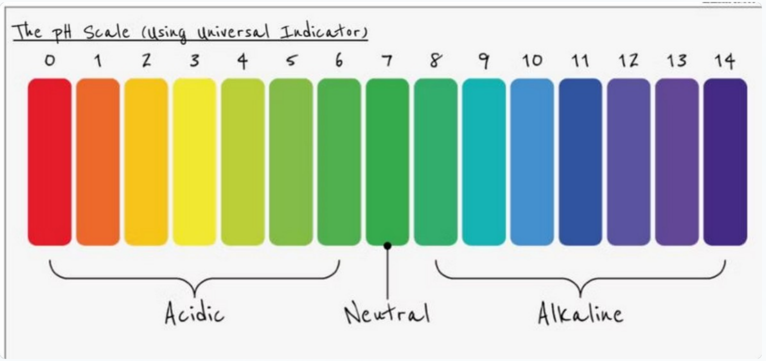
Mixture of metal with other metals or carbon. Harder than pure metals because ions of different sizes make it harder for layers to slide.

- **Low-carbon steel** (0.25% C): hard, strong, malleable — nails, car bodies. Rusts easily.
- **High-carbon steel** (0.6–1.2% C): harder, more wear-resistant, brittle — cutting tools.
- **Stainless steel** (Fe + Cr + Ni): oxide layer prevents corrosion — sinks, cutlery.

Acids, Alkalis & Titrations

Indicators & pH Scale

Indicator	In Acid	In Alkali
Methyl orange	Red	Yellow
Phenolphthalein	Colourless	Pink
Litmus paper	Red	Blue
Universal indicator	Red	Purple



- 0–3 = strongly acidic; 4–6 = weakly acidic; 7 = neutral (green); 8–10 = weakly alkaline; 11–14 = strongly alkaline
- **H^+** ions make solutions acidic; **OH^-** ions make solutions alkaline
- **Acid** = H^+ donor; **Base** = H^+ acceptor / OH^- donor; **Alkali** = soluble base

Titration Procedure

- Use pipette to add acid to conical flask; add indicator; place on white tile
- Use burette to add alkali drop-wise; swirl to mix
- Record volume at colour change; first attempt gives approximation; repeats give exact volumes

Titration Calculation (Worked Example)

20.0 cm^3 HCl neutralises 25.0 cm^3 of 0.10 mol/dm^3 $Ca(OH)_2$. Find concentration of HCl.

$2HCl + Ca(OH)_2 \rightarrow CaCl_2 + 2H_2O$

- Moles $Ca(OH)_2$ = 0.10 × 0.025 = 0.0025 mol
- Moles HCl = 2 × 0.0025 = 0.005 mol (stoichiometry 2:1)
- Concentration HCl = 0.005 / 0.02 = **0.25 mol/dm^3**

Acids, Bases & Salt Preparations

Salt Formation Reactions

- **Metal + acid → salt + hydrogen**
  $Mg + 2HCl \rightarrow MgCl_2 + H_2$
- **Metal hydroxide + acid → salt + water**
  $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$
- **Metal oxide + acid → salt + water**
  $MgO + 2HCl \rightarrow MgCl_2 + H_2O$
- **Metal carbonate + acid → salt + water + CO_2**
  $K_2CO_3 + 2HCl \rightarrow 2KCl + H_2O + CO_2$

Solubility Rules

Salt	Soluble	Insoluble
Na, K, NH_4	All	None
Nitrates	All	None
Sulfates	Most	Pb^{2+}, Ba^{2+}, Ca^{2+}
Chlorides	Most	Pb^{2+}, Ag^{+}
Carbonates	NH_4, K, Na	Most
Hydroxides	NH_4, K, Na (Ca slightly)	Most

Methods for Making Salts

Crystallisation Method

For soluble salts (not Na, K, NH_4). React acid + insoluble metal/oxide/hydroxide/carbonate. Steps: REACT → FILTER → EVAPORATE → COOL → DRY.

Titration Method

For Na, K, NH_4 salts (all soluble — no excess to filter). Steps: REACT (acid from burette + alkali in flask) → INDICATOR → REPEAT without indicator → EVAPORATE → COOL → DRY.

Precipitation Method

For insoluble salts. Mix two soluble salts → insoluble precipitate forms. Steps: REACT → FILTER → WASH → DRY. e.g. $AgNO_3 + NaCl \rightarrow AgCl\downarrow + NaNO_3$

Brønsted-Lowry: Acid = H^+ donor; Base = H^+ acceptor. e.g. $HCl + H_2O \rightarrow H_3O^+ + Cl^-$ (HCl donates H^+; H_2O accepts H^+)

Chemical Tests

Tests for Gases

Gas	Test	Positive Result
Hydrogen	Lit splint	Squeaky pop
Oxygen	Glowing splint	Splint relights
Carbon dioxide	Bubble through limewater	Turns cloudy/milky
Ammonia	Damp red litmus paper	Turns blue
Chlorine	Damp blue litmus paper	Bleaches (turns white)

Flame Tests for Cations

Ion	Flame Colour
Li^{+}	Red
Na^{+}	Yellow
K^{+}	Lilac
Ca^{2+}	Orange-red
Cu^{2+}	Blue-green

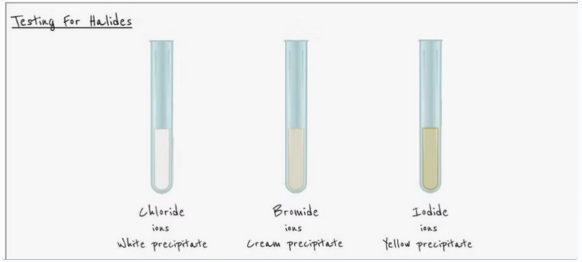
Procedure: dip nichrome wire in HCl to clean → dip in sample → hold in roaring blue Bunsen flame.

Tests for Cations (NaOH Precipitation)

Ion	Precipitate
Cu^{2+}	Light blue $Cu(OH)_2$
Fe^{2+}	Green $Fe(OH)_2$
Fe^{3+}	Brown $Fe(OH)_3$
NH_4^{+}	No precipitate; choking ammonia gas given off → turns damp red litmus blue

Tests for Anions

Ion	Test	Result
Cl^{-}	Dilute HNO_3 + $AgNO_3$	White AgCl precipitate
Br^{-}	Dilute HNO_3 + $AgNO_3$	Cream AgBr precipitate
I^{-}	Dilute HNO_3 + $AgNO_3$	Yellow AgI precipitate
SO_4^{2-}	Dilute HCl + $BaCl_2$	White $BaSO_4$ precipitate
CO_3^{2-}	Dilute HCl	Fizzing; CO_2 turns limewater cloudy



Test for Water

- White anhydrous $CuSO_4$ turns blue (hydrated $CuSO_4$ forms)
- Pure water boils at exactly 100°C

Energetics

Key Definitions

- **ΔH** = enthalpy (energy) change of a reaction
- **Exothermic:** heat energy released; ΔH is **negative**; more energy released making bonds than breaking them
- **Endothermic:** heat energy taken in; ΔH is **positive**; more energy needed to break bonds than is released making them
- **Activation energy:** minimum energy required for a reaction to occur
- **Bond energy:** energy needed to break the bond between two atoms

Calorimetry Experiments

- **Combustion:** copper calorimeter; measure temperature change of water
- **Displacement, dissolving, neutralisation:** insulated polystyrene cup calorimeter (reduces heat loss to environment)

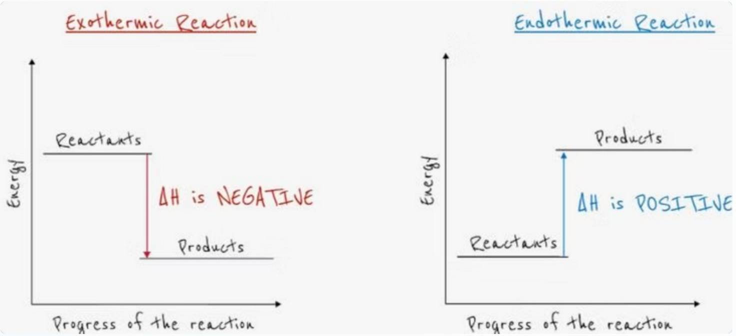
Calculations

$$Q = mc\Delta T \quad (c = 4.2\ \text{J/g°C})$$

$$\Delta H = Q/n$$

Example: 6.3 g LiI dissolved in 100 cm^3 water; ΔT = 4.9°C. Q = 100 × 4.2 × 4.9 = **2058 J**

Energy Level Diagrams



Bond Energy Calculations

$$\Delta H = \text{bonds broken} - \text{bonds formed}$$

Worked Example: $2CH_4 + O_2 \rightarrow 2CH_3OH$. Bonds broken: 2×(4×C–H) + O=O = 3480 + 497 = 3977 kJ. Bonds formed: 2×(3×C–H + C–O + O–H) = 4210 kJ. ΔH = 3977 − 4210 = **−233 kJ (exothermic)**

Bond breaking = endothermic. Bond making = exothermic.

Experimental ΔH often differs from data book values due to: heat loss to atmosphere, incomplete combustion, slow dissolving, water evaporation.

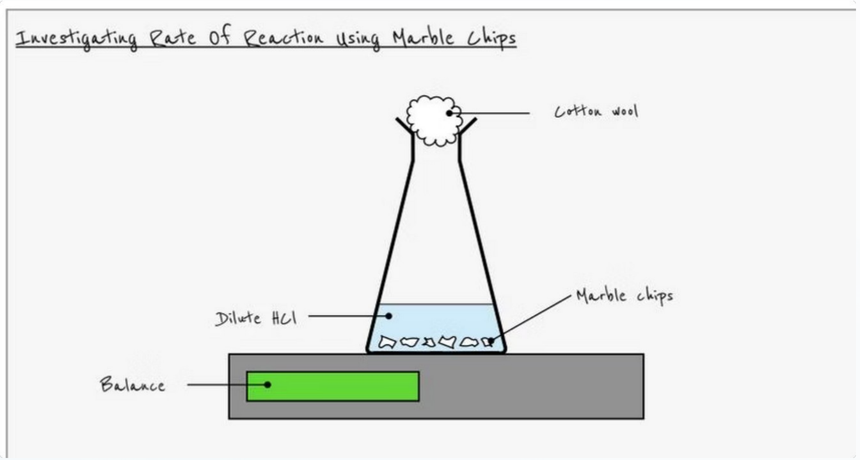
Rates of Reaction

Measuring Rate

- Rate = amount of reactant used / time OR amount of product formed / time

Experimental Methods

- **Precipitate formed (e.g. $Na_2S_2O_3$ + HCl):** time until X on paper can no longer be seen through solution
- **Gas produced (e.g. H_2O_2 decomposition):** time to collect 100 cm^3 O_2 in gas syringe
- **Change in mass (e.g. $CaCO_3$ + HCl):** place flask on balance; measure mass of CO_2 lost at intervals; cotton wool in neck to stop acid spraying



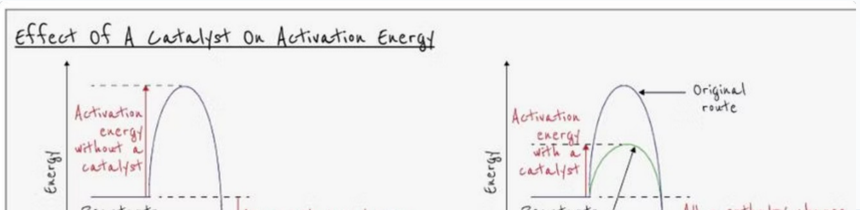
Factors Affecting Rate

↑ Temperature
Particles have greater KE → more collisions exceed activation energy → more successful collisions per second

↑ Concentration / Pressure
More particles in same volume → collisions occur more frequently → rate increases

↑ Surface Area
More particles exposed → collisions occur more frequently → rate increases

Catalyst
Provides alternative pathway with lower activation energy → rate increases without being used up



Reversible Reactions & Equilibria

Key Concepts

- ⇌ arrow indicates a **reversible reaction**
- **Dynamic equilibrium:** forward and reverse reactions occur at the same rate; concentrations of reactants and products remain constant; requires a sealed container
- A catalyst has **no effect** on the position of equilibrium — it increases the rate of both forward and reverse reactions equally; equilibrium is reached faster

Le Chatelier's Principle

When a change is made to a system in equilibrium, the system moves to **oppose that change**.

Change	Effect on Equilibrium
↑ Temperature	Favours endothermic reaction
↓ Temperature	Favours exothermic reaction
↑ Pressure	Favours side with fewer moles of gas
↓ Pressure	Favours side with more moles of gas

The Haber Process

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta H = -92$ kJ/mol

Conditions: **450°C, 200 atmospheres, iron catalyst**

- Forward reaction is exothermic → favoured by low temperature, but rate too slow → **compromise: 450°C**
- Forward reaction gives fewer moles of gas → favoured by high pressure, but dangerous/expensive → **compromise: 200 atm**
- Iron catalyst has no effect on yield; only increases rate

Worked Example: $NO_2 \rightleftharpoons N_2O_4$ ($\Delta H = -57$ kJ/mol)

- ↑ **Temperature:** favours endothermic (reverse) reaction → equilibrium shifts left → mixture turns **brown** (more NO_2)
- ↑ **Pressure:** 2 moles gas on left, 1 mole on right → equilibrium shifts right → mixture turns **colourless** (more N_2O_4)

Reversible Reactions in Practice

- **$CuSO_4 \cdot 5H_2O$:** heat → white anhydrous powder; add water → blue crystals reform
- **NH_4Cl:** heat → splits into NH_3 + HCl; cools higher up tube → white crystals reform

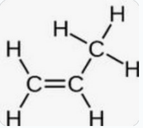
Introduction to Organic Chemistry

Key Definitions

- **Hydrocarbon:** compound containing hydrogen and carbon atoms ONLY
- **Molecular formula:** exact number of atoms of each element in a compound
- **Empirical formula:** simplest ratio of atoms of each element
- **General formula:** relationship between number of atoms within a molecule
- **Structural formula:** how atoms are joined together
- **Displayed formula:** drawing of all bonds within a compound
- **Functional group:** atom/group of atoms determining chemical properties
- **Homologous series:** group of compounds with same functional group, same chemical properties, trend in physical properties, same general formula
- **Isomerism:** compounds with same molecular formula but different structural/displayed formula

Example: Propene (C_3H_6)

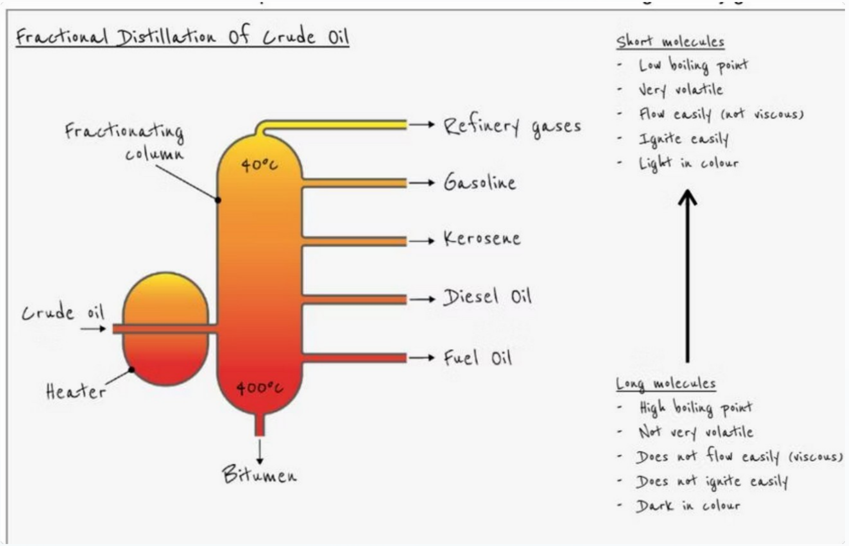
Formula Type	Propene
Molecular	C_3H_6
Empirical	CH_2
General	C_nH_{2n} (alkenes)
Structural	CH_3–CH=CH_2



Carbon Chain Mnemonic

Monkeys **E**at **P**eanut **B**utter **P**roperly = **Meth-** · **Eth-** · **Prop-** · **But-** · **Pent-**

Crude Oil & Fractional Distillation



Fractions & Their Uses

Fraction	Use
Refinery gases (CH_4, C_2H_6, C_3H_8)	Bottled gas
Gasoline	Fuel for cars
Kerosene	Fuel for planes
Diesel	Fuel for buses, lorries
Fuel oil	Fuel for ships
Bitumen	Road surfacing

Chain Length vs Properties

Short chains: low boiling point, very volatile, flow easily, ignite easily, light in colour.

Long chains: high boiling point, not volatile, viscous, do not ignite easily, dark in colour.

Combustion of Hydrocarbons

- **Complete combustion** (plentiful O_2): produces CO_2 + H_2O. e.g. $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$
- **Incomplete combustion** (insufficient O_2): produces CO + C + H_2O. CO is poisonous — combines irreversibly with haemoglobin, reducing O_2 transport.

Acid Rain

- Nitrogen oxides (from high-temperature car engines) dissolve in rainwater → nitric acid
- Sulfur impurities in crude oil → SO_2 → dissolves in rainwater → sulfuric acid
- Effects: plants die, fish die, limestone buildings corrode

Cracking

- Breaking down long alkane chains into shorter alkanes + alkenes
- Conditions: 600–700°C, alumina or silica catalyst
- Important: shorter molecules more useful as fuels; alkenes used to make alcohols/polymers

Cracking example: $C_{16}H_{34} \rightarrow C_8H_{18} + 2C_3H_6 + X$. Left: C=16, H=34. Right: C=14, H=30. X = C_2H_4 (ethene)

Alkanes

General Formula: C_nH_{2n+2}

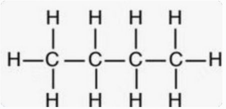
n	Name	Molecular Formula
1	Methane	CH_4
2	Ethane	C_2H_6
3	Propane	C_3H_8
4	Butane	C_4H_{10}
5	Pentane	C_5H_{12}

Properties

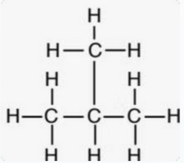
- **Saturated** hydrocarbons — all carbons form 4 single bonds; no C=C double bonds
- React with halogens in a **substitution reaction** (UV radiation required)
- e.g. $CH_4 + Br_2 \rightarrow CH_3Br + HBr$

Structural Isomers

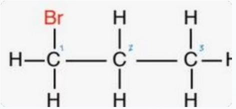
Compounds with the same molecular formula but different structural arrangements.



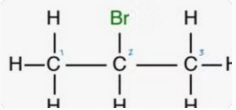
Butane (C_4H_{10}) — straight chain



Methylpropane (C_4H_{10}) — branched chain



1-bromopropane (C_3H_7Br)



2-bromopropane (C_3H_7Br)

Alkenes

General Formula: C_nH_{2n} · Functional Group: C=C

n	Name	Molecular Formula
2	Ethene	$CH_2{=}CH_2$
3	Propene	$CH_3CH{=}CH_2$
4	But-1-ene	$CH_3CH_2CH{=}CH_2$
5	Pent-1-ene	$CH_3CH_2CH_2CH_2CH{=}CH_2$

The number in the name (e.g. but-**1**-ene) indicates which carbon the double bond is attached to.

Properties

- **Unsaturated** hydrocarbons — contain C=C double bonds
- Test for unsaturation: add to **bromine water** → orange turns colourless (addition reaction; dibromoalkane formed)
- Alkanes do NOT decolourise bromine water (saturated — no double bond)
- e.g. $C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$

Addition Polymers from Alkenes

One bond in the C=C double bond breaks; monomers join to form a long chain with only single bonds.

Monomer	Polymer	Use
Ethene	Poly(ethene)	Plastic bags
Propene	Poly(propene)	Water pipes
Chloroethene	Poly(chloroethene) PVC	Window frames
Tetrafluoroethene	Poly(tetrafluoroethene) PTFE	Non-stick pan coating

⚠ **Disposal problems:** Addition polymers are inert (do not biodegrade); burning produces toxic gases.

Alcohols

Functional Group: –OH

n	Name	Molecular Formula
1	Methanol	CH_3OH
2	Ethanol	CH_3CH_2OH
3	Propan-1-ol	$CH_3CH_2CH_2OH$
4	Butan-1-ol	$CH_3CH_2CH_2CH_2OH$

Oxidation of Alcohols

- **Complete combustion:** $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$
- **Microbial oxidation:** ethanol + O_2 (in air) → ethanoic acid (vinegar)
- **Potassium dichromate(VI) in dilute H_2SO_4:** heated under reflux; solution turns orange → green

Uses

Fuels, perfumes, alcoholic drinks.

Manufacturing Ethanol

Comparison	Hydration of Ethene	Fermentation
Rate	Fast	Slow
Process type	Continuous	Batch
Renewable?	No	Yes
Temperature/Pressure	300°C, 60–70 atm, H_3PO_4 catalyst	30°C, low pressure, yeast
Purity	Pure	Impure (<15% ethanol)
Equation	$C_2H_4 + H_2O \rightarrow C_2H_5OH$	$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$

Fermentation at 30°C is the optimum temperature for yeast enzymes. Too high → enzymes denature. Too low → rate too slow. Yeast is killed when ethanol content exceeds 15%.

Carboxylic Acids, Esters & Condensation Polymers

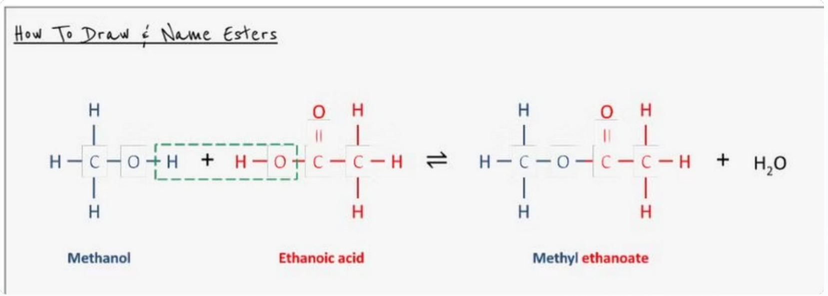
Carboxylic Acids — Functional Group: –COOH

n	Name	Formula
1	Methanoic acid	$HCOOH$
2	Ethanoic acid	CH_3COOH
3	Propanoic acid	CH_3CH_2COOH
4	Butanoic acid	$CH_3CH_2CH_2COOH$

- **With metals:** salt + hydrogen (slower than HCl). e.g. $Mg + 2CH_3COOH \rightarrow (CH_3COO)_2Mg + H_2$
- **With metal carbonates:** salt + CO_2 + water (fizzes). e.g. $Na_2CO_3 + 2CH_3COOH \rightarrow 2CH_3COONa + CO_2 + H_2O$
- Ethanoic acid is found in **vinegar**

Esters — Functional Group: –COO–

- **Esterification:** carboxylic acid + alcohol ⇌ ester + water
- Catalyst: strong acid (sulfuric acid)
- e.g. ethanoic acid + ethanol ⇌ ethyl ethanoate + water
- Used in perfumes and food flavourings — highly volatile with distinctive smells



Condensation Polymerisation

- A **dicarboxylic acid** reacts with a **diol** → polyester + water
- Differs from addition polymerisation: two different monomers required; small molecule (water) also produced
- **Biopolyesters:** condensation polymers that are biodegradable (broken down by microorganisms)

Addition & Condensation Polymer Structures

Addition Polymer Drawing Guide

Monomer	Repeating Unit	Polymer Name
Ethene (C=C with 4 H)	$–CH_2–CH_2–$	Poly(ethene)
Propene (C=C, one CH_3)	$–CH_2–CH(CH_3)–$	Poly(propene)
Chloroethene (C=C, one Cl)	$–CH_2–CHCl–$	Poly(chloroethene) PVC
Tetrafluoroethene (C=C, 4 F)	$–CF_2–CF_2–$	Poly(tetrafluoroethene) PTFE

In addition polymerisation: the C=C double bond opens; monomers join to form a long chain with only single bonds. The repeating unit is shown in square brackets with subscript n.

Condensation Polymer Drawing Guide

A dicarboxylic acid (–COOH at both ends) reacts with a diol (–OH at both ends). Water is eliminated at each bond formed.

e.g. Ethanoic acid + Ethanediol → Polyester + $2nH_2O$

The ester linkage (–COO–) is the functional group in polyesters. The –OH from the acid and –H from the alcohol's hydroxyl group combine to form water.

Biodegradability

- **Addition polymers:** inert, do not biodegrade, toxic gases when burned — major disposal problem
- **Biopolyesters:** biodegradable — broken down by microorganisms — more environmentally friendly

PRACTICAL SKILLS

Practical Skills — Experimental Design

1

Independent Variable

The variable that is being changed in the experiment. Must be clearly identified and stated.

2

Dependent Variable

The variable that is being measured. Include how it would be measured and a sensible time frame for taking measurements.

3

Control Variables

Variables kept constant throughout. State a minimum of 5. Ensures validity — only the independent variable is responsible for changes in the dependent variable.

4

Control Experiment

Allows comparison; shows what would normally happen so a comparison can be made when the independent variable is changed.

Improving Reliability

- Repeat experiment at least **3 times** and calculate the mean
- Identify and exclude anomalous results
- Ensure equal sizes/volumes of samples are used

Improving Accuracy & Validity

- Carry out more tests within existing range
- Test a wider range of values for the independent variable
- Introduce method to ensure no double counting
- Ensure all control variables are the same for each repeat

PRACTICAL SKILLS

Answering Exam Questions & Drawing Diagrams/Graphs

Question Types

Describe

Write what the data shows — trends, changes in rate, increases/decreases. Break graphs into sections (e.g. constant from A to B, increasing from B to C). Use data points from the question.

Explain

Say why the results have come about. Use scientific knowledge to explain patterns and trends. Make sure explanation is specific to the question.

Compare

Each statement must include both pieces of data. e.g. "both A and B remained constant for 2 hours" or "the rate of increase of A was greater than that of B from 2 to 4 hours".

Drawing Diagrams

- Include a title; use a sharp pencil
- Labels should be outside the diagram; use ruled label lines (do not cross them)
- Include a scale bar and state magnification
- Do not use shading; use at least 50% of available space

Drawing Graphs

- Label both axes with units
- Use a sensible scale — use at least 50% of available space
- Plot points accurately using an 'X'
- Draw a line of best fit if required

REVISION SUMMARY

Key Formulae & Equations Reference

Mole Calculations

$$\text{Moles} = \frac{\text{mass}}{M_r}$$

$$\text{Moles} = \text{Concentration} \times \text{Volume (dm}^3\text{)}$$

$$\text{Volume of gas} = \text{moles} \times 24\ \text{dm}^3$$

$$\%\ \text{by mass} = \frac{\text{total mass of element}}{\text{mass of compound}} \times 100$$

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

Energetics

$$Q = mc\Delta T$$

$$\Delta H = Q/n$$

$$\Delta H = \text{bonds broken} - \text{bonds formed}$$

Rates & Chromatography

$$\text{Rate} = \frac{\text{amount of product formed}}{\text{time}}$$

$$R_f = \frac{\text{distance travelled by component}}{\text{distance travelled by solvent}}$$

Solubility

$$\text{Solubility (g/100g)} = \frac{100 \times \text{mass of solute}}{\text{mass of evaporated solvent}}$$

Relative Atomic Mass

$$A_r = \frac{(\%\ \text{isotope 1} \times \text{mass}_1) + (\%\ \text{isotope 2} \times \text{mass}_2)}{100}$$

Neutralisation

$$H^+ + OH^- \rightarrow H_2O$$

REVISION SUMMARY

Organic Chemistry Homologous Series Summary

Alkanes

General formula: C_nH_{2n+2}

Functional group: none (C–C single bonds only)

Saturated — substitution reactions with halogens (UV light)

Alkenes

General formula: C_nH_{2n}

Functional group: C=C

Unsaturated — addition reactions; decolourise bromine water

Alcohols

General formula: $C_nH_{2n+1}OH$

Functional group: –OH

Oxidised to carboxylic acids; used as fuels and solvents

Carboxylic Acids

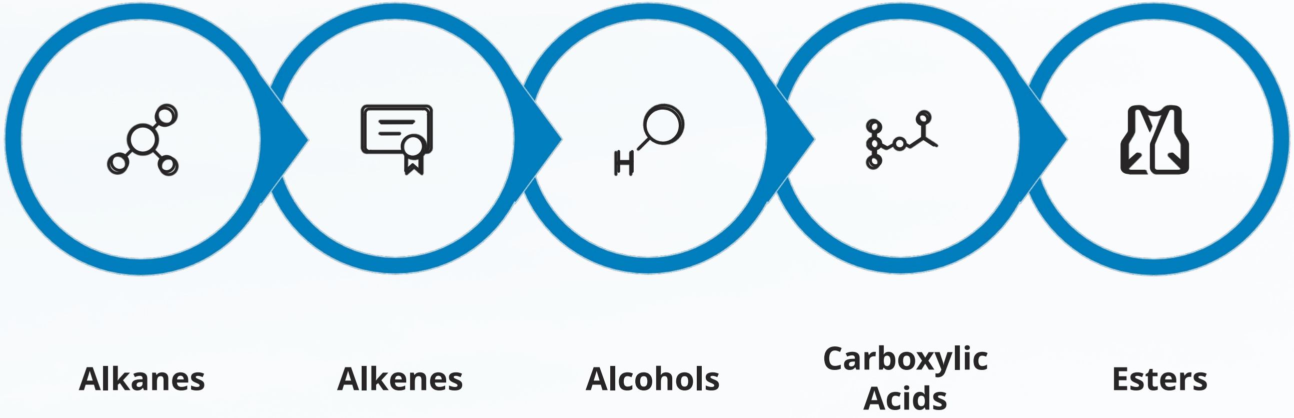
Functional group: –COOH

React with metals → salt + H_2; with carbonates → salt + CO_2 + H_2O

Esters

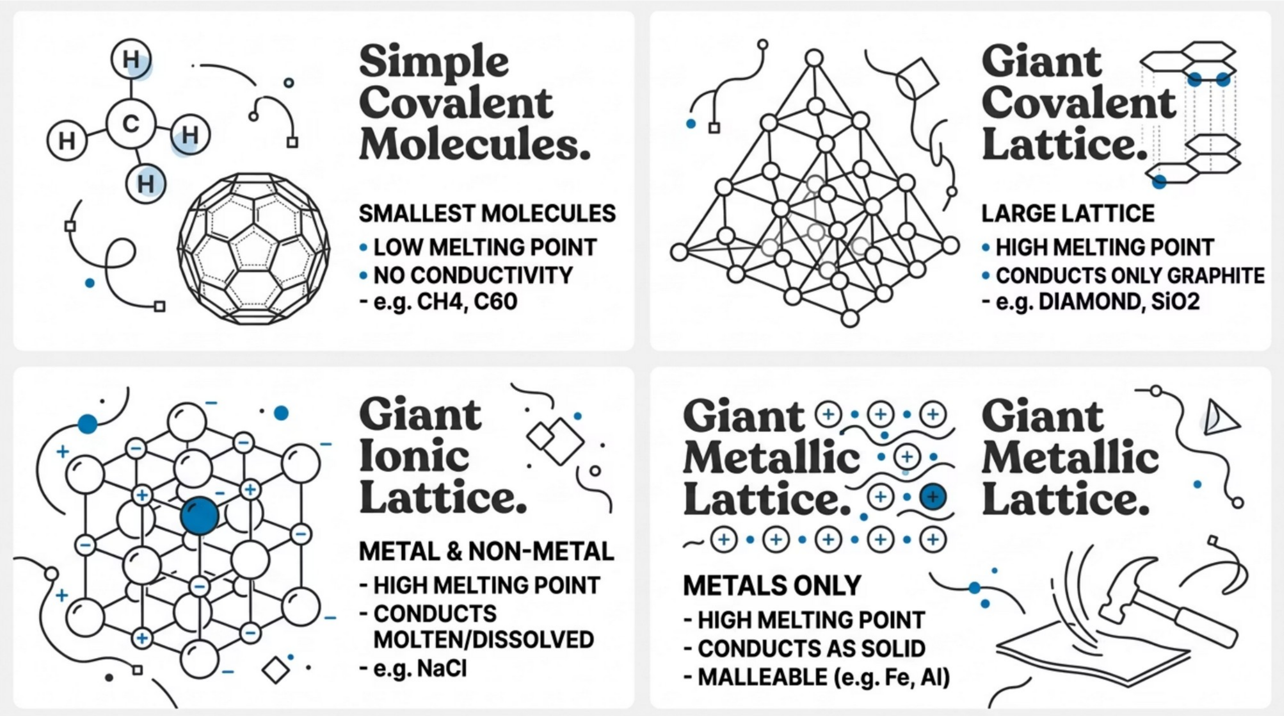
Functional group: –COO–

Formed from carboxylic acid + alcohol (H_2SO_4 catalyst); used in perfumes and flavourings



REVISION SUMMARY

Bonding & Structure — Quick Reference



Conductivity Rules

- **Simple covalent:** never conducts (no free electrons, no charge)
- **Giant covalent:** never conducts except graphite (delocalised 4th electron)
- **Giant ionic:** conducts only when molten or dissolved (ions free to move)
- **Giant metallic:** always conducts (delocalised electrons free to move)

Melting Point Rules

- **Simple covalent:** low — weak intermolecular forces only
- **Giant covalent, ionic, metallic:** all high — strong bonds/forces require lots of energy to break
- C_{60} fullerene has lower mp than diamond/graphite — simple molecular structure, weak intermolecular forces

REVISION SUMMARY

Electrolysis — Complete Summary

Ionic Solutions — Decision Rules

→ Cathode (−)

H_2 produced by default. Exception: if metal cation is less reactive than hydrogen → metal deposited instead (e.g. $Cu^{2+} \rightarrow Cu$).

→ Anode (+)

O_2 produced by default. Exception: if halide ions (Cl^-, Br^-, I^-) present → halogen produced instead (e.g. Cl_2).

Half Equations Summary

Electrolyte	Cathode	Anode
Aq. NaCl	$2H^+ + 2e^- \rightarrow H_2$	$2Cl^- - 2e^- \rightarrow Cl_2$
Aq. $CuSO_4$	$Cu^{2+} + 2e^- \rightarrow Cu$	$Cu - 2e^- \rightarrow Cu^{2+}$ (or O_2 if inert)
Dil. H_2SO_4	$2H^+ + 2e^- \rightarrow H_2$	$4OH^- - 4e^- \rightarrow 2H_2O + O_2$
Molten $PbBr_2$	$Pb^{2+} + 2e^- \rightarrow Pb$	$2Br^- - 2e^- \rightarrow Br_2$

Commercial Uses of Electrolysis Products

Hydrogen

Used as a fuel (clean energy source)

Chlorine

Used to make bleach; kills pathogens in swimming pools

Sodium Hydroxide

Used in making paper and bleach

Copper Purification

Impure Cu anode loses mass; pure Cu cathode gains mass — used to purify copper for electrical wiring

REVISION SUMMARY

Reactivity Series & Salt Preparation — Decision Flowchart

Reactivity Series Recap

K · Na · Li · Ca · Mg · Al · (C) · Zn · Fe · (H) · Cu · Ag · Au

- Metals above H react with dilute acid → salt + H_2
- Metals above C extracted by electrolysis; below C by reduction with carbon
- Displacement: more reactive metal displaces less reactive from its compound

Rusting Prevention Summary

Barrier

Paint, grease, oil — physical barrier; if damaged, iron rusts

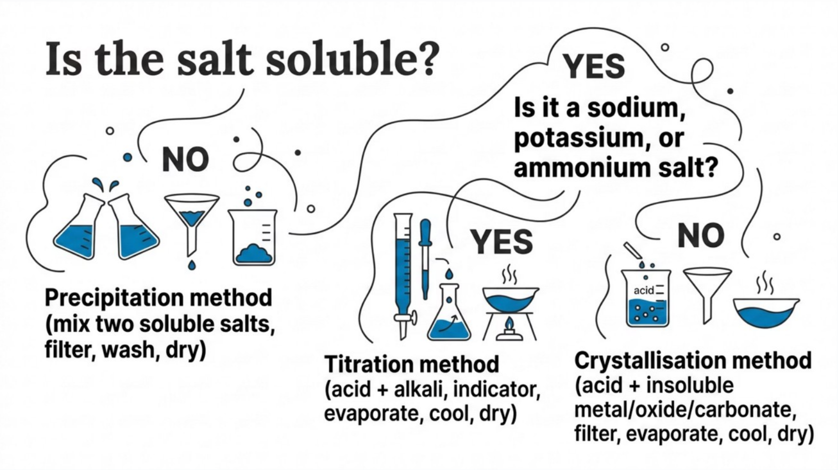
Galvanising

Zinc coating; Zn oxidises preferentially even if scratched

Sacrificial

More reactive metal block attached; oxidises instead of iron

Salt Preparation Decision Guide



REVISION SUMMARY

Physical Chemistry — Key Concepts at a Glance

Energetics Summary

Term	Exothermic	Endothermic
ΔH	Negative	Positive
Energy diagram	Products lower than reactants	Products higher than reactants
Bond energy	More energy released making bonds than breaking	More energy needed to break bonds than released making
Examples	Combustion, neutralisation, displacement	Thermal decomposition, photosynthesis

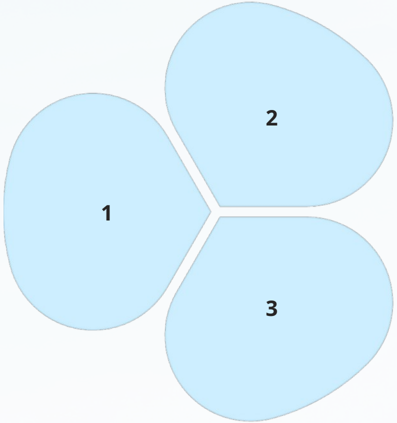
Equilibrium Summary

Change	Equilibrium Shift
↑ Temperature	→ endothermic direction
↓ Temperature	→ exothermic direction
↑ Pressure	→ fewer moles of gas
↓ Pressure	→ more moles of gas
Add catalyst	No shift; equilibrium reached faster

Haber Process Conditions

$N_2 + 3H_2 \rightleftharpoons 2NH_3$ (ΔH = −92 kJ/mol). Conditions: **450°C** (compromise — low T favours yield but too slow), **200 atm** (compromise — high P favours yield but dangerous/expensive), **iron catalyst** (no effect on yield; increases rate only).

Rates of Reaction Summary



Temperature
More KE → more successful collisions

Concentration
More particles → more frequent collisions

Surface Area
More exposed particles → more collisions

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Revision Complete — Key Takeaways

Principles of Chemistry

- States of matter, diffusion, solubility, separation techniques
- Atomic structure, periodic table, electronic configurations
- Ionic, covalent and metallic bonding; giant structures
- Mole calculations, reacting masses, empirical formulae, electrolysis

Inorganic Chemistry

- Group 1 (alkali metals) and Group 7 (halogens) trends and reactions
- Reactivity series, rusting, redox, extraction of metals
- Acids, alkalis, titrations, salt preparations
- Chemical tests for gases, cations, anions and water

Physical Chemistry

- Energetics: ΔH, $Q=mc\Delta T$, bond energies, energy level diagrams
- Rates of reaction: temperature, concentration, surface area, catalysts
- Reversible reactions, dynamic equilibrium, Le Chatelier's principle, Haber process

Organic Chemistry

- Crude oil, fractional distillation, cracking, combustion
- Alkanes, alkenes, alcohols, carboxylic acids, esters
- Addition and condensation polymerisation
- Manufacturing ethanol: hydration vs fermentation

Remember to practise worked examples, learn key equations off by heart, and apply the practical skills guidance to all experimental design questions. Good luck! 🎓