

PREMIUM STUDY NOTES · CLASS 10

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Chemistry

4CH1

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19 chapters · ~5,218 words of student-ready notes

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Chapter 1 – States of Matter

Diffusion, solute, solvent, solubility – the language of matter and its mixing.

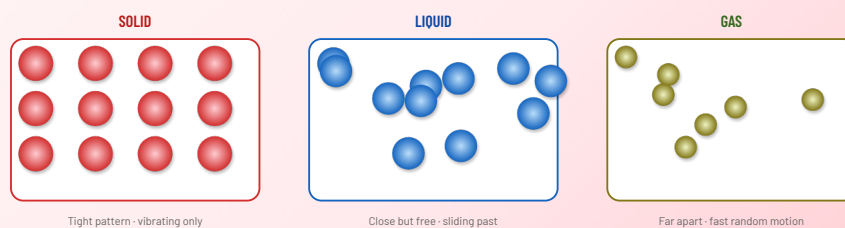
Everything around you is made of **tiny particles**. Whether something is a solid, a liquid, or a gas depends on how those particles are arranged and how much energy they have to move. The particle picture is the foundation of every chapter that follows.

DIFFUSION

DIFFUSION

Diffusion is the spreading out of particles from a region of **higher concentration** to a region of **lower concentration**.

Diffusion happens because particles in liquids and gases are constantly in random motion. Over time, they mix evenly throughout the available space. No external force is needed – the particles move themselves. Diffusion is faster in gases than liquids (gas particles are further apart and move faster), and does not occur in solids.

FIGURE 1.1

Particles in solid, liquid, gas. Same particles – arrangement loosens and motion increases as energy rises.

HOW TO READ FIGURE 1.1

Solid (red): tightly packed grid; particles only vibrate. **Liquid (blue):** particles still touching but slide past each other. **Gas (yellow):** particles spread far apart with empty space between – volume expands roughly 1,000× from solid to gas.

SOLUTIONS AND SOLUBILITY

FOUR DEFINITIONS TO MEMORISE

Solute – the substance that dissolves (in a liquid).

Solvent – the liquid that dissolves the solute.

Solution – a uniform mixture of solute and solvent.

Saturated solution – a solution that contains as much dissolved solid as possible at a particular temperature.



SOLUBILITY

Solubility is the mass of solute that dissolves in 100 g of solvent at a given temperature to form a saturated solution.

Units: g per 100 g of solvent (or sometimes g/dm^3).

HOW TO HELP A SOLUTE DISSOLVE

(1) **Stir** – brings undissolved solute into contact with fresh solvent. (2) **Raise the temperature** – particles move faster and break apart more easily. (3) **Crush** the solute into smaller pieces – more surface area exposed.

Chapter 2 – Elements, Compounds, Mixtures & Separation Techniques

Filtration, crystallization, distillation, chromatography – and how to read a chromatogram.

THREE DEFINITIONS

Element: a pure substance that cannot be split into anything simpler by any chemical means – contains only one type of atom.

Compound: formed when two or more elements are chemically combined in a fixed ratio.

Mixture: various substances mixed together without any chemical reaction (so they can be separated physically).

FILTRATION

Filtration is used to separate an undissolved (insoluble) solid from a mixture of the solid and a liquid – e.g. sand from a mixture of sand and water. **Centrifugation** can also be used for the same task.

1. Filter paper is placed in a filter funnel above another beaker.
2. The mixture of insoluble solid and liquid is poured into the filter funnel.
3. Small liquid particles pass through the filter paper as the **filtrate**.
4. Solid particles are too large to pass through – they stay behind on the paper as the **residue**.

CRYSTALLIZATION

Crystallization is used to separate a **solute** from a **solution**.

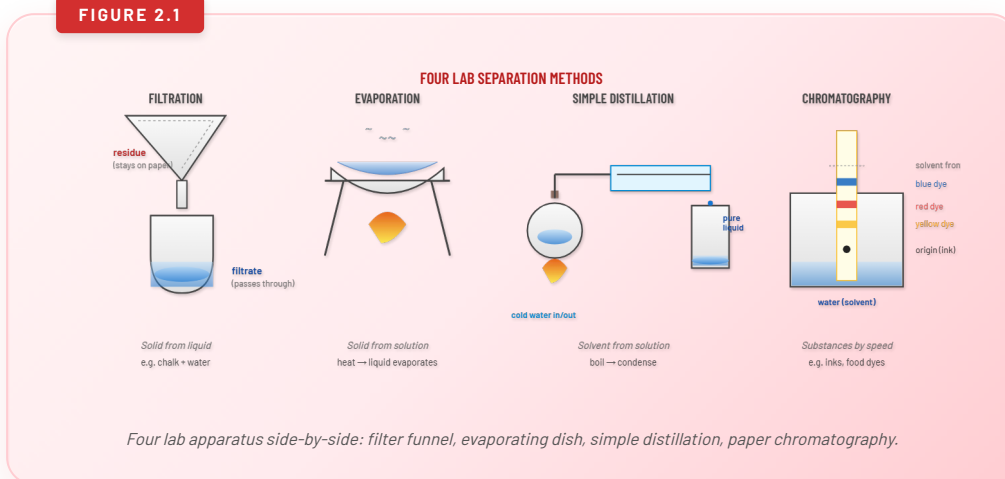
1. The solution is heated in an evaporating basin to boil off some of the water.
2. Test for the point of crystallization by dipping a glass rod – crystals form on the rod.
3. Leave the solution to cool. On cooling, more crystals form.
4. Crystals are removed from the mixture by filtration.
5. Leave the crystals in a warm place to dry.

SIMPLE DISTILLATION

Simple distillation is used to separate a liquid from a soluble solid in a solution (e.g. pure water from saltwater) or a pure liquid from a mixture of liquids.

1. The solution is heated and pure water evaporates, producing vapour that rises through the neck of the round-bottomed flask.
2. The vapour passes through the **condenser**, where it cools and condenses, turning into pure water.
3. The pure water is collected in a beaker.
4. After all the water has evaporated, only the solid solute remains in the flask.

FIGURE 2.1



HOW TO READ FIGURE 2.1

Each panel is a labelled cross-section. **Filtration**: residue stays on the paper, filtrate drips through. **Evaporation**: heat drives the liquid into vapour, solid stays behind. **Simple distillation**: liquid is boiled, vapour passes through a cooled condenser, drips out as pure liquid. **Chromatography**: ink is spotted on paper, the rising solvent carries different dyes at different speeds.

FRACTIONAL DISTILLATION

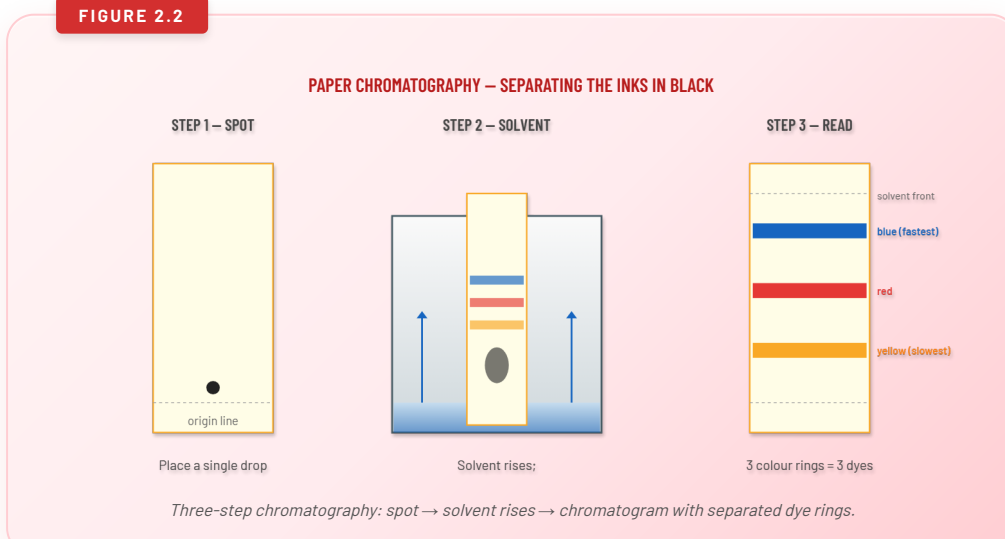
Fractional distillation separates mixtures of two or more liquids that are completely miscible, by taking advantage of their different boiling points (e.g. ethanol and water). The most volatile (lowest boiling point) liquid evaporates first, condenses at the top of the column, and is collected first.

PAPER CHROMATOGRAPHY

Chromatography separates substances that have different solubilities in a given solvent – e.g. different coloured inks that have been mixed to make black ink.

1. A **pencil line** is drawn on chromatography paper and spots of the sample are placed on it. (Pencil is used because ink would run into the chromatogram.)
2. The paper is lowered into the solvent container, making sure the pencil line sits **above** the level of the solvent so the samples don't wash off.
3. The solvent travels up the paper by **capillary action**, taking some of the coloured substances with it.
4. Different substances have different solubilities, so they travel at different rates – substances with higher solubility travel further than the others.
5. This shows the different components of the ink or dye as separate spots on the paper.

FIGURE 2.2



 HOW TO READ FIGURE 2.2

Step 1: a single ink dot on the origin line is the unknown mixture. Step 2: solvent climbs the paper, pulling each dye component along at its own pace. Step 3: the finished **chromatogram** shows three coloured bands – meaning the original ink contained three different dyes (blue fastest, then red, then yellow slowest).

INTERPRET SIMPLE CHROMATOGRAMS

- If two or more substances are the **same**, they produce identical chromatograms.
- It is common practice to include a **known compound as a reference spot** – this can help match an unknown spot to identify it.
- If the substance is a **mixture**, it separates on the paper to show all its different components as separate spots.
- An **impure substance** shows up with more than one spot. A **pure substance** shows up with only one spot.

RETARDATION FACTOR (R_f)

 R_f VALUE

The **retardation factor (R_f)** is the ratio of the distance moved by a spot from the baseline (pencil line) to the distance moved by the solvent front from the baseline.

 FORMULA

$R_f = \frac{\text{distance moved by spot}}{\text{distance moved by solvent front}}$
 R_f is a ratio – it has **no units** and is always between 0 and 1.

Chapter 3 – Atomic Structure

Atomic number, mass number, isotopes – the numbers that identify every atom.

ATOMIC NUMBER / CHARGE NUMBER / PROTON NUMBER / Z

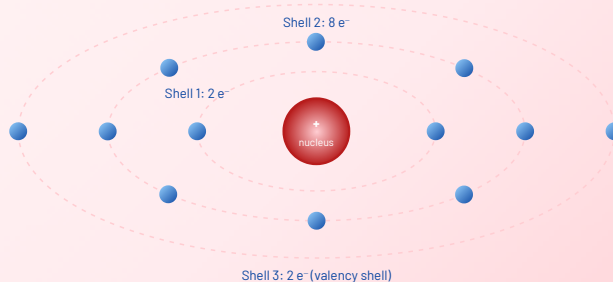
It is the **number of protons** in an atom. Denoted by 'Z'. It identifies the element – only sodium has 11 protons, only carbon has 6, etc.

MASS NUMBER / NUCLEON NUMBER / A

It is the **number of protons + neutrons** in the nucleus of an atom. Denoted by 'A'. The number of neutrons can be calculated as $A - Z$.

FIGURE 3.1

MAGNESIUM (2,8,2)



Magnesium atom ($Z = 12$). Nucleus + three shells filled in the order 2, 8, 2.

HOW TO READ FIGURE 3.1

The red dot at the centre is the **nucleus** – it carries 12 protons (atomic number) plus 12 neutrons. The three orbits are the electron **shells**. Count electrons on each shell: $2 + 8 + 2 = 12$ total – matching the 12 protons in a neutral atom. We write this as the *electronic configuration* 2,8,2.

ISOTOPES

ISOTOPES (IN TERMS OF SUB-ATOMIC PARTICLES)

Isotopes are atoms of the **same element** that have the **same number of protons** but a **different number of neutrons**.

WHY DO ISOTOPES HAVE SIMILAR CHEMICAL PROPERTIES?

Because they have the **same electronic configuration** – the same number of outer-shell electrons. Chemistry is determined by the electrons (not the neutrons), so isotopes react identically. They differ only in mass.

Chapter 4 – The Periodic Table

Periods, groups, electronic configuration – and why metals conduct.

PERIODS & GROUPS

Periods: horizontal rows in the periodic table.

Groups: vertical columns in the periodic table.

Electronic configuration: the arrangement of electrons in an atom (e.g. 2,8,1 for sodium).

WHY ARE GROUP 8 (GROUP 0) ELEMENTS NOBLE (UNREACTIVE)?

- They have **full outer shells**
- They don't lose, gain, or share electrons
- They don't form stable ions

WHY ARE METALS GOOD CONDUCTORS OF ELECTRICITY?

Because they have **delocalized electrons** that can move when a voltage (potential difference) is applied. The electrons travel through the lattice, carrying charge.

WHY ARE METALS MALLEABLE AND DUCTILE?

In metals, atoms are arranged in **layers**, and these layers can **slide** over each other when a force is applied to them. Because the metallic bond is non-directional, the bond is preserved as the layers slip.

MALLEABLE & DUCTILE

Malleable: can be hammered into shapes (e.g. sheet metal). **Ductile:** can be drawn into wires (e.g. copper wires).

Chapter 5 – Chemical Formulae, Equations and Calculations (Part I)

Relative atomic mass, the mole, empirical and molecular formulas, hydrated salts.

RELATIVE ATOMIC MASS (RAM OR A_r)

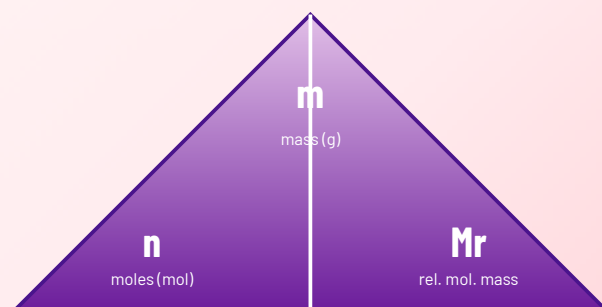
The **relative atomic mass** of an element is the weighted average mass of the isotopes of the element. It is measured on a scale where a carbon-12 (^{12}C) atom has a mass of exactly 12.

MOLE

A **mole** is a unit of the amount of substance – the quantity of a substance containing **Avogadro's number** of particles (6.02×10^{23}).

FIGURE 5.1

THE MOLE TRIANGLE – MASS · MOLES · M_r



$$n = m / M_r \cdot m = n \times M_r \cdot M_r = m / n$$

The mole triangle. Cover what you want; what's left is the formula. $n = m / M_r$

HOW TO READ FIGURE 5.1

Mass (m) sits at the top; moles (n) and molar mass (M_r) at the bottom. Cover the variable you want and the remaining two give the formula. Want moles? Cover $n \rightarrow n = m \div M_r$. Want mass? Cover $m \rightarrow m = n \times M_r$.

EMPIRICAL FORMULA VS MOLECULAR FORMULA

EMPIRICAL FORMULA

Shows the **simplest whole-number ratio** of the atoms present in a compound. Example: the empirical formula of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) is **CH_2O** .

MOLECULAR FORMULA

Shows the **actual number** of atoms of each element present in a molecule or compound. Example: the molecular formula of glucose is **$\text{C}_6\text{H}_{12}\text{O}_6$** .

HYDRATED AND ANHYDROUS COMPOUNDS

CHEMISTRY · 4CH1

EXAM FREQUENCY: HIGH

Syllabus ref: Unit 1 / Ch 5

CHAPTER 05 / Chemical Formulae I · continued

HYDRATED VS ANHYDROUS

Hydrated means containing water of crystallization, e.g. $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.

Anhydrous means without water of crystallization, e.g. CuSO_4 .

WATER OF CRYSTALLIZATION

When some substances crystallize from solution, water becomes **chemically bound up** with the salt. This water is called the **water of crystallization**, and the salt is said to be a **hydrated salt**.

Chapter 6 – Chemical Formulae, Equations and Calculations (Part II)

Avogadro's law, molar volume, neutralization, titration.

AVOGADRO'S LAW

Equal volumes of gases at standard temperature and pressure (STP) contain equal numbers of molecules.

MOLAR VOLUME (V_m)

The volume occupied by one mole of a gas. At standard temperature and pressure: $V_m = 24 \text{ dm}^3$ (or $24,000 \text{ cm}^3$).

MOLAR MASS (M)

The mass of one mole of a substance. Unit: g mol^{-1} (or g/mol).

USEFUL EQUATIONS

Solids: $n = m \div M_r$

Solutions: $n = c \times V$ (concentration in $\text{mol/dm}^3 \times$ volume in dm^3)

Gases at STP: $n = V \div 24$ (volume in dm^3)

NEUTRALIZATION

NEUTRALIZATION REACTION

The reaction of an acid with a base or alkali to produce salt and water.

General: $\text{Acid} + \text{Base} \rightarrow \text{Salt} + \text{Water}$

TITRATION

TITRATION

A technique used to find out how much of one solution reacts with a certain volume of another solution of known concentration. Used to find unknown concentration, or to prepare a pure salt.

Chapter 7 – Ionic Bonding

Cation, anion, the ionic lattice – and why ionic compounds behave the way they do.

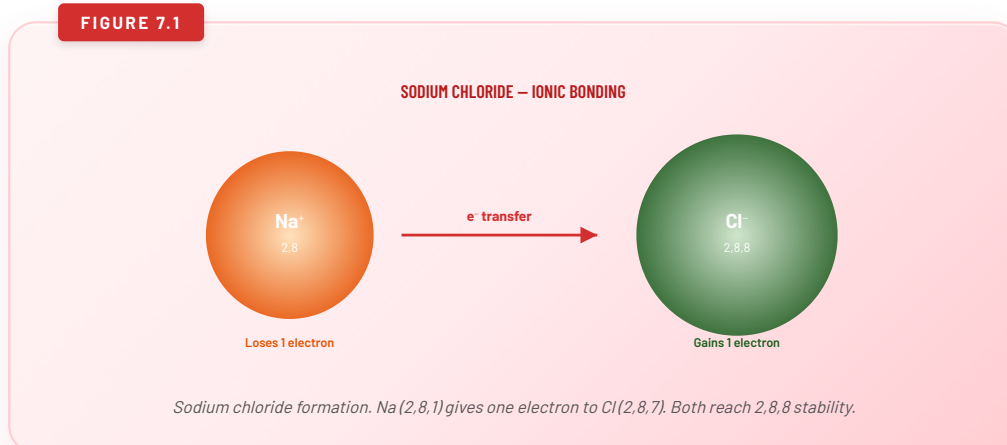
IONIC BONDING

Ionic bonding is the **electrostatic attraction** between oppositely charged ions.

CATION & ANION

Cation: a positively charged ion (an atom that has lost electrons).

Anion: a negatively charged ion (an atom that has gained electrons).

FIGURE 7.1

HOW TO READ FIGURE 7.1

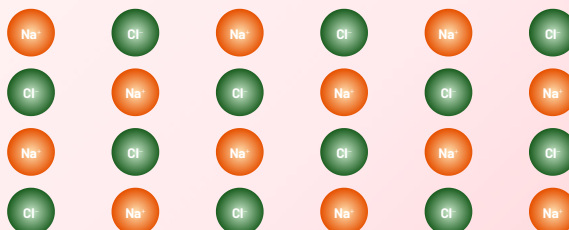
The orange Na on the left has one extra outer electron. The green Cl on the right is one short of a full shell. The red arrow shows the electron transferring. After transfer: Na has 10 e⁻ vs 11 p⁺ (charge +1, Na⁺); Cl has 18 e⁻ vs 17 p⁺ (charge -1, Cl⁻). The opposite charges attract – that's the ionic bond.

WHY DOES SODIUM CHLORIDE (NaCl) HAVE A HIGH MELTING POINT?

It has a **giant ionic structure**. There are strong forces of attraction between oppositely charged ions throughout the lattice. To overcome these strong attractions, a **large amount of heat energy** is required – so the melting point is high.

FIGURE 7.2

IONIC LATTICE – NaCl GIANT STRUCTURE



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

NaCl giant ionic lattice – alternating Na^+ and Cl^- ions in 3D. Each ion is surrounded by 6 ions of opposite charge.

HOW TO READ FIGURE 7.2

Each ion isn't paired with one partner – it sits in a 3D pattern surrounded by ions of the opposite charge. Strong electrostatic attractions in every direction explain the high m.p. and the inability to conduct as a solid.

WHY DON'T IONIC COMPOUNDS CONDUCT ELECTRICITY IN THE SOLID STATE?

Because in the solid state, the **ions are not free to move** – they are locked in the lattice. (Don't write 'electrons' here; that is a marking trap. Ionic conduction is by ion movement.) When melted or dissolved, the ions become free to move and the compound conducts.

WHY ARE IONIC COMPOUNDS BRITTLE?

Any small distortion of the crystal brings ions with the **same charge** alongside each other. **Like charges repel**, so the crystal splits itself apart along the new mismatched plane. This is why hitting an ionic crystal with a hammer shatters it.

Chapter 8 – Covalent Bonding

Sharing electrons; simple molecules; diamond, graphite, fullerene.

COVALENT BONDING

Covalent bonding is the force of attraction between a **shared pair of electrons** and the nuclei of two atoms.

WHY CO₂ / NH₃ / CCL₄ HAVE LOW BOILING POINTS

These have **simple molecular structures** with **weak intermolecular forces**. Only a little energy is needed to overcome these weak forces between molecules – so they boil at low temperatures. (The covalent bonds *inside* each molecule are strong; it is the forces *between* molecules that are weak.)

FIGURE 8.1

GIANT COVALENT STRUCTURES – ALLOTROPES OF CARBON + SILICA



- Hardest known solid
 - High m.p.
 - Insulator (no free e⁻)
- Each C bonded to

4 others. Use: drill bits, jewellery
GRAPHITE layers of hexagonal rings free e⁻ flows along layer → · Soft & slippery · **Conducts electricity** · High m.p. Each C bonded to **3** others. Use: pencil leads, electrodes
SILICA (SiO₂) tetrahedral Si-O-Si Si O O O O · Hard, scratches things · High m.p. · Lets light through Each Si bonded to 4 O; each O to 2 Si. Use: glass, sand, lenses

Three giant covalent structures: Diamond (tetrahedral), Graphite (layered), Silica (Si-O network).

HOW TO READ FIGURE 8.1

Look at bond count per atom. **Diamond**: each carbon bonded to 4 others – rigid 3D web (hardest substance, no free electrons). **Graphite**: each carbon bonded to only 3 in flat layers – the 4th electron is FREE (conducts electricity; layers slide so it's soft). **Silica**: tetrahedral with Si and O alternating.

WHY IS DIAMOND HARD, AND GRAPHITE SOFT?

Diamond: each carbon atom has four electrons in its outermost shell – it forms **four covalent bonds** to other carbons in a tetrahedral arrangement. A lot of energy is required to break these strong bonds – diamond is the hardest naturally occurring substance.

Graphite: each carbon forms only **three covalent bonds** within a layer. The forces holding atoms together within each layer are very strong, but the attractions *between* the layers are **much weaker** – not much energy is needed to overcome them. Layers slide over each other and can easily be flaked off – that's why graphite is soft and feels slippery.

WHY DOES C₆₀ FULLERENE HAVE A LOWER MELTING POINT THAN DIAMOND AND GRAPHITE?

C_{60} fullerene has a **simple molecular structure** with **weak intermolecular forces** between separate C_{60} molecules. To overcome these forces, only a relatively small amount of heat energy is needed – so the melting and boiling points are much lower than those of the giant covalent diamond and graphite.

WHY DOESN'T C_{60} FULLERENE CONDUCT ELECTRICITY?

Although all the carbon atoms in C_{60} form only three bonds, the 4th electron on each atom can **only move around within the same C_{60} molecule** – electrons cannot jump from one molecule to another. Without electrons free to flow between molecules, no current can pass.

Chapter 9 – Metallic Bonding

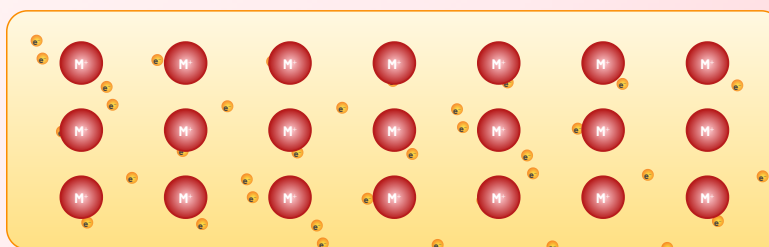
The sea of delocalized electrons that explains every metal property.

METALLIC BONDING

Metallic bonding is the **electrostatic force of attraction** between the cations (positive metal ions) and the **delocalized electrons**.

FIGURE 9.1

METALLIC BONDING – LATTICE OF +VE IONS IN A SEA OF ELECTRONS



Free electrons (yellow) carry charge and heat · +ve ions (red) hold the lattice together

Lattice of M^+ ions in a sea of free electrons – one diagram explains every metal property.

HOW TO READ FIGURE 9.1 (ONE PICTURE, FIVE PROPERTIES)

(1) **High m.p.** – the +ions are pulled in every direction by the electron sea. (2) **Conducts electricity** – free electrons carry charge through the lattice. (3) **Conducts heat** – free electrons carry kinetic energy. (4) **Malleable & ductile** – layers of +ions slide past each other without breaking the bond, because the electron sea adjusts. (5) **Shiny when polished** – the electron sea reflects light.

Chapter 10 – Electrolysis

Breaking ionic compounds with electricity; PANIC and OIL/RIG mnemonics.

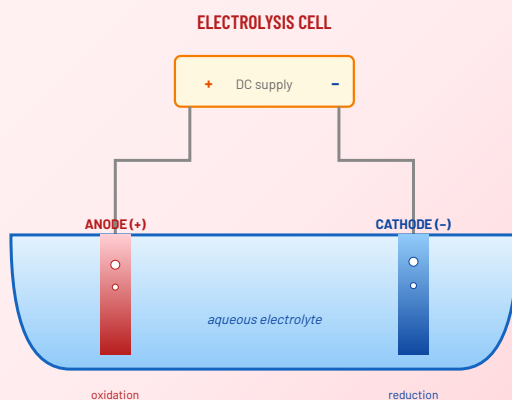
ELECTROLYSIS

Electrolysis is a chemical change caused by passing an electric current through a compound that is either **molten** or in **solution**. In simpler words: the breakdown of a compound by using electricity.

ELECTROLYTE

An **electrolyte** is a liquid or solution that undergoes electrolysis. Electrolytes all contain **ions**. The movement of these ions is responsible for both the **conduction of electricity** and the **chemical changes** that take place.

FIGURE 10.1



Electrolysis cell. Cathode (-) attracts cations → reduction. Anode (+) attracts anions → oxidation.

HOW TO READ FIGURE 10.1

DC supply provides electrons. The **cathode** (negative electrode, on the right) attracts positive ions – metals deposit here. The **anode** (positive electrode, on the left) attracts negative ions – non-metals form here. The electrolyte must be molten or in aqueous solution so the ions can move.

MEMORISE THESE NAMES

TERM	MEANING
Anode	Positive (+) electrode – attracts anions
Cathode	Negative (-) electrode – attracts cations

MNEMONICS – PANIC & OIL RIG

PANIC = Positive Anode, Negative Is Cathode.

OIL RIG = Oxidation Is Loss (of electrons), Reduction Is Gain (of electrons).

DISCHARGING, OXIDATION, REDUCTION

THREE DEFINITIONS

Discharging: an ion **loses its charge** – happens at the electrodes during electrolysis.

Oxidation: when a substance **loses electrons** or **gains oxygen**.

Reduction: when a substance **gains electrons** or **loses oxygen**.

BRINE

BRINE

Brine is a concentrated aqueous solution of sodium chloride (NaCl). Industrially, electrolysis of brine produces three valuable products at once: chlorine gas (anode), hydrogen gas (cathode), and sodium hydroxide solution (left in the cell).

Chapter 11 – Alkali Metals

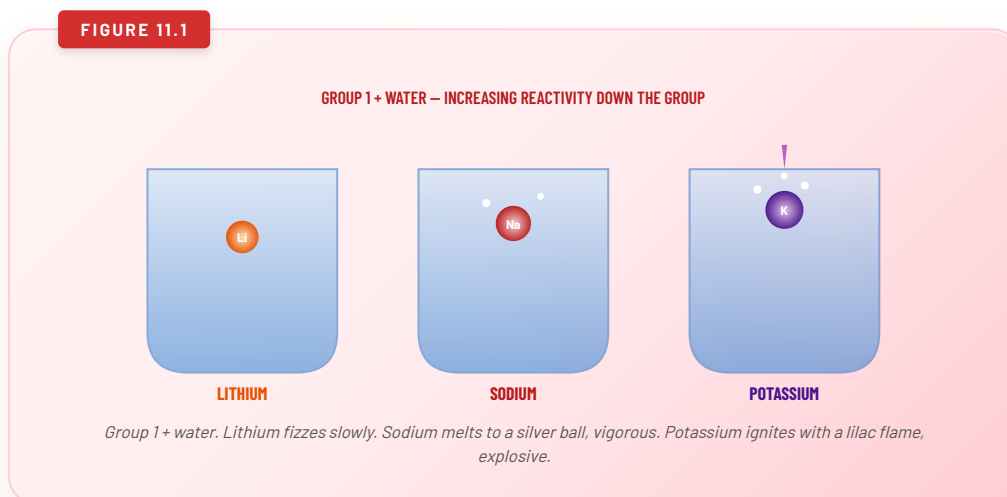
Group 1 – why they're similar, why they're called 'alkali metals', and why they get more reactive down the group.

WHY DO THE CHEMICAL PROPERTIES OF GROUP 1 ELEMENTS RESEMBLE ONE ANOTHER?

Because they all have the **same number of electrons in their outermost shell** – exactly one. Atoms with the same number of valency electrons react in the same way. That single outer electron is the engine of all Group 1 chemistry.

WHY ARE GROUP 1 ELEMENTS REFERRED TO AS ALKALI METALS?

Because when they react with water, they produce **soluble bases called alkalis**. Example: $2\text{Na} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2$ – the NaOH solution is alkaline (pH around 14).



HOW TO READ FIGURE 11.1

Three beakers of water, each with a different alkali metal. As you go $\text{Li} \rightarrow \text{Na} \rightarrow \text{K}$ (down the group), the reaction gets **more violent** – from gentle fizzing through vigorous melting to flame ignition.

WHY DOES THE REACTIVITY OF GROUP 1 ELEMENTS INCREASE AS YOU GO DOWN THE GROUP?

Three reasons working together:

1. Shells are gradually added down the group – the atom gets **larger**.
2. The **average distance** between the nucleus and the outer shell increases.
3. **Shielding** (inner electrons partly blocking the nucleus) also increases.

Together, these three effects mean the attraction between the nucleus and the outer-shell electron **decreases** down the group. The outer electron becomes easier to lose – so the atom reacts more readily.

Chapter 12 – The Halogens

Group 7 – salt-formers, displacement, and the reactivity trend that goes the other way.

Halogens are salt-producing (or salt-forming) substances. They form a variety of salts, including potassium iodide (KI) and calcium chloride (CaCl₂), when they react with metals.

DISPLACEMENT REACTIONS INVOLVING HALOGENS

DISPLACEMENT RULE

A halogen that is **more reactive** can **displace** a less reactive halogen from its salt solution.

Example: Cl₂ (more reactive) displaces Br⁻ from KBr solution: Cl₂ + 2KBr → 2KCl + Br₂ (the colourless solution turns orange-brown as Br₂ is released).

FIGURE 12.1

GROUP 7 (HALOGENS) – PHYSICAL TRENDS DOWN THE GROUP



Down Group 7: F₂ (yellow gas) → Cl₂ (green gas) → Br₂ (red-brown liquid) → I₂ (purple-black solid). Reactivity DECREASES; m.p./b.p. and density INCREASE; colour darkens.

HOW TO READ FIGURE 12.1

Colour bar runs from yellow at top (F₂, gas) to dark purple at bottom (I₂, solid). Red arrow: reactivity drops down the group. Blue arrow: m.p./b.p. and density rise. State change too: F₂ and Cl₂ are gases, Br₂ is a liquid, I₂ is a solid at room temperature.

WHY DOES REACTIVITY DECREASE DOWN GROUP 7?

Going down the group, the **attraction between the nucleus and the outer electron decreases** as the number of shells gradually increases. Halogens react by **gaining** one electron to fill their outer shell – but a larger atom with more shielding has a weaker pull on a new incoming electron. Going down the group makes it **more difficult to attract one electron** – so the element becomes less reactive.



CHEMISTRY · 4CH1

EXAM FREQUENCY: HIGH

Syllabus ref: Unit 2 / Ch 12

CHAPTER 12 / The Halogens · continued

GROUP 1 VS GROUP 7 – MIRROR TRENDS

Group 1 metals get **more** reactive down the group (easier to LOSE the outer electron when it's further from the nucleus). Halogens get **less** reactive down the group (harder to GAIN an electron when the outer shell is further from the nucleus). Same physics, opposite directions, because the two groups are doing opposite things.

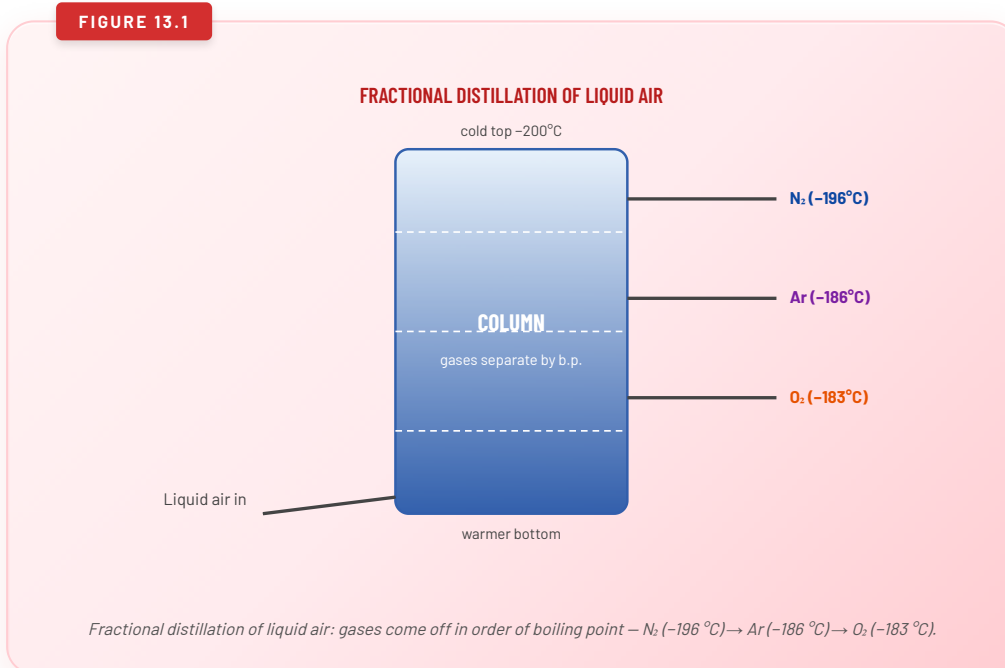
Chapter 13 – Gases in the Atmosphere

Composition of air; combustion observations; thermal decomposition; the greenhouse effect.

COMPOSITION OF UNPOLLUTED DRY AIR

GAS	% IN AIR
Nitrogen (N ₂)	78%
Oxygen (O ₂)	21%
Argon (Ar)	0.9%
Carbon dioxide (CO ₂)	0.04%

FIGURE 13.1



HOW TO READ FIGURE 13.1

Air is first liquefied (cooled and compressed). The liquid feeds the bottom of a tall column. The column is COLD at the top and warmer at the bottom. Each gas boils off at its own boiling point and is drawn off through a tap at the matching height: N₂ first (lowest b.p.), then Ar, then O₂.

BURNING MAGNESIUM IN AIR

MAGNESIUM + OXYGEN



OBSERVATIONS

(1) **Bright white flame** is produced. (2) **White powdery ash** (magnesium oxide, MgO) is formed.

CHAPTER 13 / Gases in the Atmosphere · continued

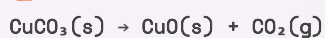
- Sulphur burns with a **blue** flame
- Hydrogen burns with a **pale-blue** flame

THERMAL DECOMPOSITION

THERMAL DECOMPOSITION

Thermal decomposition is breaking up a compound by **heating** it.

WORKED EXAMPLE – COPPER(II) CARBONATE



Green solid → black solid (the colour change is the proof of reaction).

THE GREENHOUSE EFFECT

GREENHOUSE EFFECT

Occurs when high-energy **UV and visible light** from the Sun passes through the atmosphere and warms up the surface of the Earth. Greenhouse gases re-absorb the heat radiated back, trapping it. Main greenhouse gases: **CO₂, CH₄, H₂O, CFCs**. (CFCs = Chlorofluorocarbons.)

Chapter 14 – Reactivity Series

Reducing/oxidizing agents, metals + water/steam, rust and how to prevent it.

REDUCING AGENT (REDUCTANT)

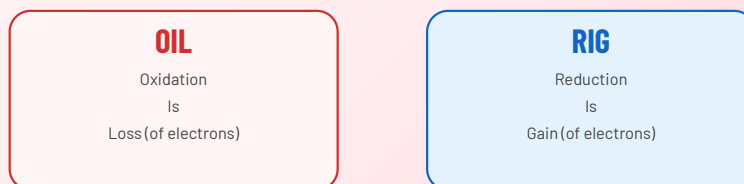
A substance that **reduces** something else but is **itself oxidized**.

OXIDIZING AGENT (OXIDANT)

A substance that **oxidizes** something else but is **itself reduced**.

FIGURE 14.1

OIL RIG – REMEMBER REDOX



OIL RIG – Oxidation Is Loss of electrons; Reduction Is Gain. Memorise it.

REACTIONS WITH COLD WATER (ABOVE HYDROGEN ONLY)

GENERAL EQUATION

Metal + Water → Metal hydroxide + Hydrogen

- $2\text{Na(s)} + 2\text{H}_2\text{O(l)} \rightarrow 2\text{NaOH(aq)} + \text{H}_2\text{(g)}$ (Group 1)
- $\text{Ca(s)} + 2\text{H}_2\text{O(l)} \rightarrow \text{Ca(OH)}_2\text{(aq)} + \text{H}_2\text{(g)}$ (Group 2)

REACTIONS WITH STEAM

GENERAL EQUATION

Metal + Steam → Metal oxide + Hydrogen

- $\text{Mg(s)} + \text{H}_2\text{O(g)} \rightarrow \text{MgO(s)} + \text{H}_2\text{(g)}$
- $\text{Zn(s)} + \text{H}_2\text{O(g)} \rightarrow \text{ZnO(s)} + \text{H}_2\text{(g)}$
- $3\text{Fe(s)} + 4\text{H}_2\text{O(g)} \rightarrow \text{Fe}_3\text{O}_4\text{(s)} + 4\text{H}_2\text{(g)}$

RUST

RUST

Rust is **hydrated iron(III) oxide**, formula $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$. ($\text{Fe}^{3+} + \text{O}^{2-}$ gives Fe_2O_3 = iron(III) oxide.)

WORD EQUATION FOR RUST FORMATION

Iron + Water + Oxygen → Rust (hydrated iron(III) oxide)

TWO ESSENTIAL CONDITIONS FOR RUSTING

Both **water AND oxygen** must be present. Remove either one and rusting stops. The classic three-tube experiment proves this.

GALVANIZED IRON

GALVANIZED IRON

Iron coated with a **layer of zinc**. Zinc is more reactive than iron, so it rusts in preference and protects the iron underneath.

METHODS TO PROTECT IRON FROM RUSTING

- Painting
- Oiling
- Greasing
- Covering in plastic
- **Galvanizing** (zinc coat)
- **Sacrificial anodes** (cathodic protection) – bolt blocks of more reactive metal to the iron object

WHAT IS THE IRON COMPOUND IN RUST?

Fe_2O_3 (**haematite**), hydrated – same compound found in iron ore.

Chapter 15 – Extraction and Uses of Metals

Ore, alloy, and three steels – mild, high-carbon, and stainless.

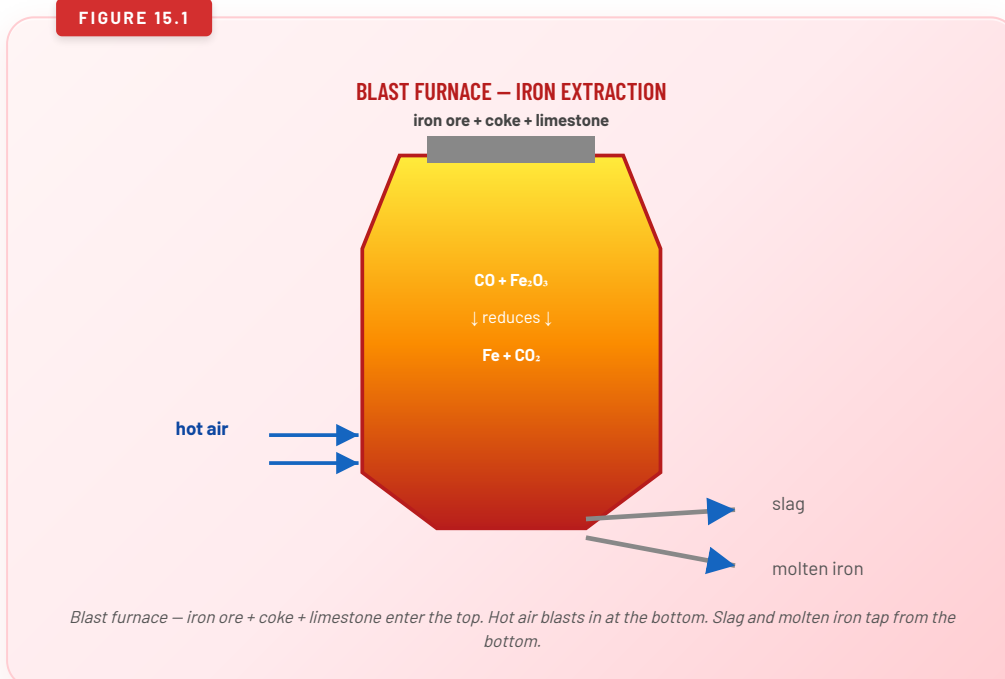
ORE

An **ore** is a sample of rock that contains **enough of a mineral** for it to be worthwhile to extract the metal from it.

ALLOY

An **alloy** is a mixture of a metal with, usually, other metals or carbon. Examples: **brass** = copper + zinc · **steel** = iron + carbon.

FIGURE 15.1



HOW TO READ FIGURE 15.1

Iron is extracted by reduction with carbon. Stage 1 (bottom): coke burns $\rightarrow \text{CO}_2 + \text{heat}$. Stage 2: CO_2 reacts with more coke $\rightarrow \text{CO}$. Stage 3: CO reduces iron(III) oxide: $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$. Limestone removes silica impurities as slag. Molten iron is tapped at the bottom.

WHY ARE ALLOYS HARDER THAN THE PURE METALS FROM WHICH THEY ARE MADE?

In an alloy, the different metals or elements have **slightly different sized atoms**. This breaks up the regular lattice arrangement and makes it **more difficult for the layers of ions to slide** over each other. Result: the alloy is harder and stronger but less malleable.

COMMON ALLOYS TO MEMORISE

Bronze = copper + tin. **Stainless steel** = iron + chromium + nickel. **Cupronickel** = copper + nickel (used to make 'silver' coins).

THREE MAIN TYPES OF STEEL

CHEMISTRY · 4CH1

EXAM FREQUENCY: HIGH

Syllabus ref: Unit 2 / Ch 15

CHAPTER 15 / Extraction and Uses of Metals · continued

TYPE	CARBON %	PROPERTIES	USES
Mild steel	Up to 0.25%	Strong, malleable, ductile (rusts in air)	Nails, car bodies, ship building, girders, bridges
High-carbon steel	0.6–1.2% (often + Mn)	Harder, more wear-resistant; brittle (less malleable)	Cutting tools, hammers
Stainless steel	+ chromium and often nickel	Very corrosion-resistant (chromium oxide layer protects)	Kitchen sinks, saucepans, knives & forks, gardening tools, brewing/dairy/chemical industry vessels

⚠ DISADVANTAGES OF MILD STEEL

Mild steel **rusts** when exposed to oxygen and water. It is also about **three times denser** than aluminium – so some car bodies are made from aluminium instead, which doesn't rust and saves fuel because the car is lighter. Stainless steel resists corrosion but is significantly more expensive than mild steel.

Chapter 16 – Acids, Alkalis and Titrations

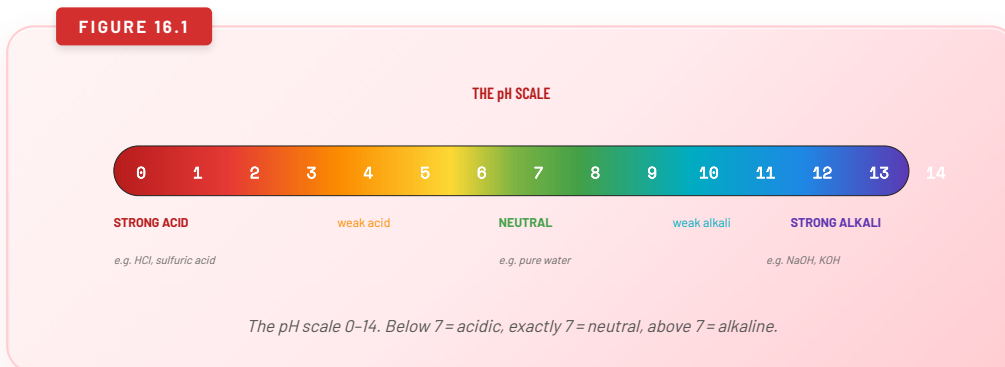
The pH scale and the four acid-base indicators you must memorise.

PH SCALE

The **pH scale** ranges from about 0 to 14 and tells how acidic or alkaline a solution is.

PH RANGE	DESCRIPTION
0-3	Strongly acidic
4-6	Weakly acidic
7	Neutral
8-10	Weakly alkaline (basic)
11-14	Strongly alkaline

FIGURE 16.1



HOW TO READ FIGURE 16.1

Colour bar runs from red (very strong acid, pH 0) through green (neutral, pH 7) to purple (very strong alkali, pH 14). Each step represents a 10^x change in H⁺ concentration. Pure water sits at exactly 7.

ACID-BASE INDICATORS

ACID-BASE INDICATOR

A substance that has different **colours depending on pH** – can be used to detect acidity or alkalinity.

CHEMISTRY · 4CH1

EXAM FREQUENCY: HIGH

Syllabus ref: Unit 2 / Ch 16

CHAPTER 16 / Acids, Alkalis & Titrations · continued

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

Chapter 17 – Acids, Bases and Salts Preparation

Common acids, bases, alkalis – and how to perform a titration step-by-step.

ACID

An **acid** is a **proton donor** – that is, an H^+ donor.

COMMON EXAMPLES OF ACIDS

ACID	CATION	ANION	NAME
HCl	H^+	Cl^-	Hydrochloric acid
H_2SO_4	$2H^+$	SO_4^{2-}	Sulphuric acid
HNO_3	H^+	NO_3^-	Nitric acid
CH_3COOH	H^+	CH_3COO^-	Ethanoic / acetic acid
H_3PO_4	H^+	PO_4^{3-}	Phosphoric acid

BASE

Bases are substances that **neutralize acids** by combining with the H^+ ions in them. **Alkalis** are *soluble* bases – e.g. NaOH, KOH.

POLYATOMIC IONS TO MEMORISE

Hydronium ion = H_3O^+

Ammonium ion = NH_4^+

PROCEDURE FOR PERFORMING A TITRATION

1. Measure a known volume of alkali into the titration flask using a **pipette**. (First wash the pipette with a little of the alkali.)
2. Add a few drops of indicator solution to the alkali in the flask.
3. Fill a clean **burette** with acid. (First wash the burette with a little of the acid.)
4. Record the burette reading.
5. Open the burette tap and let the acid flow into the flask. Keep **swirling the flask** gently to ensure mixing.
6. Keep adding acid slowly until the indicator changes its colour. This is the **end point**.
7. Record the new burette reading. (Final reading – initial reading) = the **titre**.
8. **Repeat at least three times** to get concordant results (results within 0.1 cm^3 of each other).

PROCEDURE FOR MAKING AN INSOLUBLE SALT BY PRECIPITATION

CHEMISTRY · 4CH1

EXAM FREQUENCY: HIGH

Syllabus ref: Unit 2 / Ch 17

CHAPTER 17 / Acids, Bases & Salts Preparation · continued

1. Mix the reactant solutions and stir.
2. **Filter** to separate the insoluble salt (residue) from the soluble products (filtrate).
3. **Wash** the residue with distilled water to remove soluble impurities.
4. **Collect** the salt and leave it in a warm place to dry.

Chapter 18 – Chemical Tests

Tests for gases, flame tests for cations, anion tests, and chemical tests for water.

CHEMICAL TESTS FOR GASES

GAS	FORMULA	TEST	RESULT
Hydrogen	H ₂	Lighted splint	Burns with a squeaky pop
Oxygen	O ₂	Glowing splint	Relights
Carbon dioxide	CO ₂	Bubble through limewater	Limewater turns milky
Chlorine	Cl ₂	Damp blue litmus paper	Turns red, then bleaches
Chlorine	Cl ₂	Damp red litmus paper	Bleaches
Ammonia	NH ₃	Damp red litmus paper	Turns blue

FLAME TEST

1. Take a clean piece of **platinum wire** (or nichrome wire).
2. Dip it in concentrated HCl, then into the solid sample of salt.
3. Place the wire into the **non-luminous Bunsen burner flame**.
4. Observe the flame colour and use it to identify the metal ion present.

💡 WHY PLATINUM WIRE?

(1) It has a **high melting point**. (2) It doesn't give any colour to the flame. (3) It doesn't react (it is inert / unreactive).

💡 WHY A NON-LUMINOUS FLAME?

So that the colour of the flame from the metal ion can be seen **easily** – a luminous (yellow) flame would mask the colour.

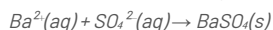
TEST FOR POSITIVE IONS (ADD NaOH)

ION	OBSERVATION ON ADDING NaOH	EQUATION
Cu^{2+}	Blue precipitate	$\text{Cu}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s})$
Fe^{2+}	Green precipitate	$\text{Fe}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_2(\text{s})$
Fe^{3+}	Brown-orange precipitate	$\text{Fe}^{3+}(\text{aq}) + 3\text{OH}^{-}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_3(\text{s})$
NH_4^{+}	No ppt; on warming, NH_3 gas evolved (turns damp red litmus blue)	$\text{NH}_4^{+}(\text{aq}) + \text{OH}^{-}(\text{aq}) \rightarrow \text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l})$

TEST FOR ANIONS

TEST FOR SULPHATE (SO_4^{2-})

Add dilute HCl followed by BaCl_2 solution. A **white precipitate** forms. (Or add dilute HNO_3 followed by $\text{Ba}(\text{NO}_3)_2$.)



TEST FOR CARBONATE (CO_3^{2-})

Add dilute $\text{HNO}_3 \rightarrow$ **bubbles or fizzing**. Pass the gas through limewater \rightarrow **turns milky**.



TEST FOR HALIDES (X^{-} – Cl^{-} , Br^{-} , I^{-})

Add dilute HNO_3 followed by silver nitrate (AgNO_3) solution. The colour of the precipitate identifies the halide:

- **White** precipitate \rightarrow **chloride**: $\text{Ag}^{+} + \text{Cl}^{-} \rightarrow \text{AgCl}(\text{s})$
- **Cream** precipitate \rightarrow **bromide**: $\text{Ag}^{+} + \text{Br}^{-} \rightarrow \text{AgBr}(\text{s})$
- **Yellow** precipitate \rightarrow **iodide**: $\text{Ag}^{+} + \text{I}^{-} \rightarrow \text{AgI}(\text{s})$

CHEMICAL TESTS FOR WATER

FIGURE 18.1

CHEMICAL TESTS FOR WATER

Anhydrous copper(II) sulfate

Anhydrous cobalt(II) chloride

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 turns white \rightarrow blue. Anhydrous CoCl_2 turns blue \rightarrow pink.

HOW TO READ FIGURE 18.1

Both reagents change colour when water adds to them. Anhydrous copper(II) sulphate (white) becomes hydrated $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ (bright blue). Anhydrous cobalt(II) chloride paper (blue) becomes $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$ (pink). Either test proves the PRESENCE of water; for purity see Chapter 19.



- **Cobalt(II) chloride paper:** anhydrous (blue) + H₂O ⇌ hydrated (pink)
- **Copper(II) sulphate:** anhydrous (white) + H₂O ⇌ hydrated (blue)

PHYSICAL TEST FOR WATER PURITY

- Measure the boiling point – pure water boils at **100 °C**
- Measure the melting point – pure water melts at **0 °C**
- Measure the freezing point – pure water freezes at **0 °C**

Chapter 19 – Energetics (Unit 3)

Exothermic vs endothermic, enthalpy change ΔH , specific heat capacity, $Q = mc\Delta T$.

EXOTHERMIC REACTION

A reaction in which heat energy is transferred **from the system to the surroundings** (heat is given out). The temperature of the surroundings rises.

Examples of exothermic reactions:

- Combustion reactions (burning fuels)
- Reaction of metals with acids
- Neutralisation reactions
- Displacement reactions

ENDOTHERMIC REACTION

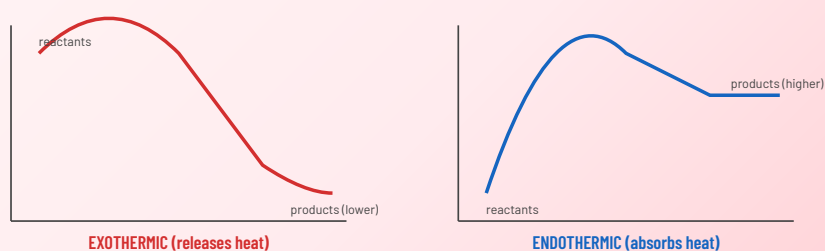
A reaction in which heat energy is transferred **from the surroundings to the system** (heat is absorbed or taken in). The temperature of the surroundings falls.

Examples of endothermic reactions:

- Thermal decomposition of calcium carbonate ($\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$)
- Photochemical reactions like photosynthesis

FIGURE 19.1

ENERGY PROFILES – EXOTHERMIC vs ENDOTHERMIC



Energy profiles. Exothermic: products lower than reactants (heat released). Endothermic: products higher (heat absorbed).

HOW TO READ FIGURE 19.1

X-axis = course of reaction (left = reactants, right = products). Y-axis = energy. **Exothermic (left, red):** the curve goes DOWN to lower-energy products – energy is released to surroundings. **Endothermic (right, blue):** the curve goes UP – energy must be absorbed from surroundings. The hump in the middle is the activation energy.

ENTHALPY CHANGE (ΔH)

ENTHALPY CHANGE ΔH

The amount of heat energy **taken in or given out** during a chemical reaction. Equivalently, the difference between the energy of the products and the energy of the reactants.

Unit of ΔH : **kJ/mol** (kilojoules per mole).

SIGN OF ΔH – ALWAYS MEMORISE

If ΔH is **positive** → the reaction is **endothermic** (heat absorbed).

If ΔH is **negative** → the reaction is **exothermic** (heat released).

SPECIFIC HEAT CAPACITY

SPECIFIC HEAT CAPACITY

The amount of heat required to raise the temperature of **1 g of a substance by 1 °C**. Unit: $\text{J g}^{-1}\text{°C}^{-1}$.

HEAT CHANGE FORMULA

$$Q = m \times c \times \Delta T$$

where Q = heat change (J), m = mass (g), c = specific heat capacity ($\text{J g}^{-1}\text{°C}^{-1}$), ΔT = temperature change ($^{\circ}\text{C}$).

WHY IS THE HEAT ENERGY RELEASED DURING COMBUSTION LOWER THAN THE REPORTED DATA BOOK VALUE?

- **Incomplete combustion** – not all the fuel is burnt cleanly to CO_2 and H_2O .
- **Heat lost to the atmosphere** – some of the heat warms the surrounding air rather than the calorimeter.
- **Conditions are not standard** – pressure, temperature or initial state may not match the data-book conditions.

Synopsis, Glossary & Examiner's Tips

A final pass over the 19 chapters of the source notebook.

SYNOPSIS

Pearson Edexcel International GCSE Chemistry – specification 4CH1 – is built around the same core idea you met in Chapter 1 – everything is made of **tiny particles** – and grows into the rich behaviour of atoms, ions, and reactions. Unit 1 covers states of matter, separation techniques, atomic structure, the periodic table, formulae & calculations, ionic / covalent / metallic bonding, and electrolysis (Chapters 1-10). Unit 2 walks through alkali metals, halogens, gases in the atmosphere, the reactivity series, metal extraction, acids & alkalis, salt preparation, and chemical tests (Chapters 11-18). Unit 3 closes with energetics – exothermic, endothermic, ΔH , and $Q = mc\Delta T$. Master the language and the exam reads itself.

EXAMINER'S TIPS (THE MARKS-WINNING SHORTLIST)

- **State symbols are free marks** – (s), (l), (g), (aq) on every full equation.
- **Balance every equation.** $Mg + O_2 \rightarrow MgO$ is wrong – it must be $2Mg + O_2 \rightarrow 2MgO$.
- **Use mark-scheme exact language.** 'An acid is a proton donor' gets the mark; 'an acid is sour' doesn't.
- **Show working in calculations.** A line of substitution earns a method mark even if the final number is wrong.
- **Distinguish observations from conclusions.** 'Bubbles formed' is observation; 'hydrogen gas was produced' is the conclusion.
- **OIL RIG:** Oxidation Is Loss (of electrons), Reduction Is Gain. **PANIC:** Positive Anode, Negative Is Cathode.
- **Sign of ΔH :** positive = endothermic, negative = exothermic.

COMPACT GLOSSARY

Atom

The smallest particle of an element that can take part in a chemical reaction.

Element

A pure substance made of one type of atom.

Compound

Two or more elements chemically combined in fixed proportions.

Mixture

Two or more substances together but not chemically combined.

Mole

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

Ion

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

Cation

A positively charged ion.

Anion

A negatively charged ion.

Electrolysis

Breakdown of an ionic compound (molten or aqueous) by electricity.

Anode / Cathode

Positive (+) electrode / negative (-) electrode (PANIC).

Oxidation

Loss of electrons (or gain of oxygen). OIL.

Reduction

Gain of electrons (or loss of oxygen). RIG.

Reducing agent

Reduces something else; itself is oxidized.

Oxidizing agent

Oxidizes something else; itself is reduced.

Acid

A proton (H^+) donor.

Base

A substance that neutralizes an acid by combining with H^+ .

Alkali

A soluble base (e.g. NaOH, KOH).

Salt

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

Ore

A rock containing enough of a mineral to make extraction worthwhile.

Alloy

A mixture of a metal with other metals or carbon.

Rust

Hydrated iron(III) oxide, $Fe_2O_3 \cdot xH_2O$.

R_f

Distance moved by spot \div distance moved by solvent front.

ΔH

Enthalpy change – heat energy taken in (+) or given out (-) per mole.



STUDENT NAME

STUDENT ID

SUBJECT

Chemistry

SUBJECT CODE

4CH1

CURRICULUM

Pearson Edexcel 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.

STUDENT SIGNATURE

INSTRUCTOR SIGNATURE

DATE