

PREMIUM STUDY NOTES · CLASS 10

# EDEXCEL PEARSON IGCSE / O-LEVEL Chemistry

4CH1

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11 chapters · ~7,764 words of student-ready notes

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# Unit 1 – States of Matter

*Particle theory, three states, diffusion, mixtures, separation, distillation, chromatography.*

Everything is made of **tiny particles**. In a solid, particles are not free to move. In liquids and gases, they can move – and as they move they collide with each other and bounce off in all directions. This is called **random motion**.

When two substances are mixed, their particles bounce off each other when they collide and gradually spread out. This is called **diffusion** – the movement of particles without an external force.

## ATOM, MOLECULE, ION

The **smallest particle** that cannot be broken down by chemical means is called an **atom**. In some substances, particles are single atoms (e.g. argon gas). In many, particles consist of two or more atoms joined as **molecules**. In others, particles are atoms or groups of atoms carrying a charge – **ions**.

## SOLIDS, LIQUIDS AND GASES

### SOLID – PROPERTIES & PARTICLE MODEL

- Definite shape and definite volume
- Normally hard and rigid
- Large force needed to change shape
- High density · incompressible
- *Model:* closely packed in a regular pattern, vibrate at fixed positions, not free to move

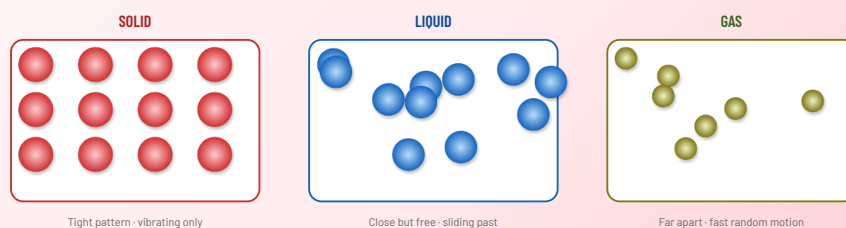
### LIQUID – PROPERTIES & PARTICLE MODEL

- Definite volume but no fixed shape
- High density · not compressible
- *Model:* occur in clusters with molecules slightly further apart than a solid; free to move within a confined vessel

### GAS – PROPERTIES & PARTICLE MODEL

- No fixed volume and no fixed shape
- Low density · compressible
- *Model:* very far apart; travel at high speed; independent random motions; negligible forces of attraction between them

FIGURE 1.1



Tight pattern · vibrating only

Close but free · sliding past

Far apart · fast random motion

Particles in solid, liquid, gas. Same particles – only arrangement and motion change.

### HOW TO READ FIGURE 1.1

**Solid (red):** tightly packed grid; vibrate but cannot move. **Liquid (blue):** still touching but slide past each other. **Gas (yellow):** spread far apart with empty space between – volume expands roughly 1,000× from solid.

## DIFFUSION IN GASES

Gases diffuse at **different rates**. Two factors decide how fast:

1. **Mass of the particles** – the lighter the molecule, the faster it diffuses (lighter particles travel faster at the same temperature).
2. **Temperature** – the higher the temperature, the faster gases diffuse (particles gain energy as they are heated).

## MIXTURES, SOLUTIONS AND SOLVENTS

### THREE DEFINITIONS

**Mixture:** contains more than one substance, just mixed together and not chemically combined (e.g. sand and water). **Solution:** when a solute and a solvent mix – the solute dissolves in the solvent (e.g. sugar dissolved in water). **Solvent:** a substance that allows solutes to dissolve in it (e.g. water, ethanol).

The solubility of every substance is different. To help a solute dissolve you can **stir** it or **raise the temperature**. If you add an excess of solute to a small amount of solvent it stops dissolving – the solution is **saturated**.

## PURE SUBSTANCES AND IMPURITIES

### PURE & IMPURE

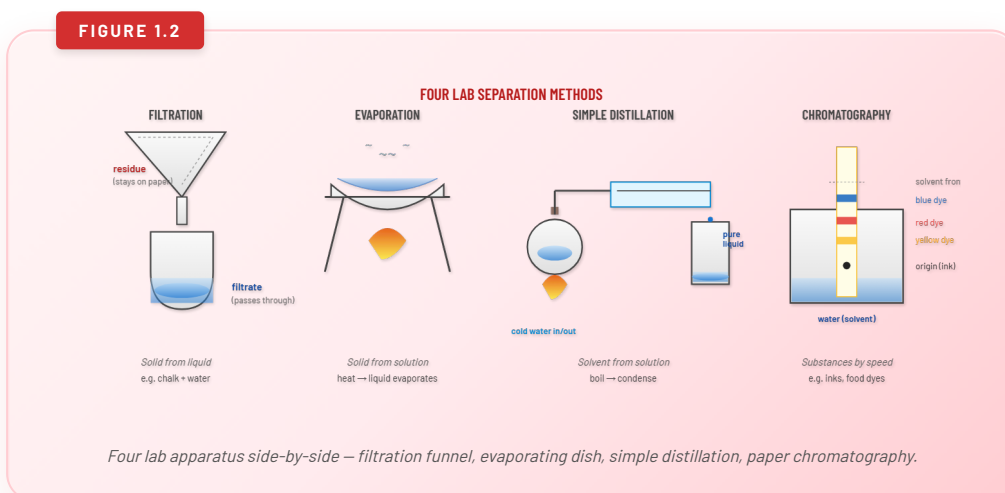
A **pure substance** contains particles of only one kind. An **impurity** is an unwanted substance mixed in with the wanted one.

### TEST FOR PURITY

Check the melting and boiling points. A pure substance has a **sharp, exact** m.p. and b.p. An impure substance: m.p. **falls** and b.p. **rises**, both spread over a range. The wider the range, the more impurity.

## SEPARATION METHODS

METHOD	SEPARATES
Filter	Solid from liquid
Centrifuge	Solid from liquid (small amounts)
Evaporation	Solid from its solution
Crystallisation	Solid (as crystals) from its solution
Simple distillation	Solvent from a solution
Fractional distillation	Liquid from each other
Chromatography	Different dissolved substances from a solution



### HOW TO READ FIGURE 1.2

Each panel is a labelled cross-section. **Filtration:** solid (residue) stays on the paper; liquid (filtrate) passes through. **Evaporation:** heat drives the liquid into vapour; solid is left behind in the dish. **Simple distillation:** the impure liquid is boiled, vapour passes through a cooled condenser, drips out as pure liquid. **Chromatography:** ink spotted on paper; rising solvent carries each component at its own speed.

### 1. FILTERING

Example: separating chalk from water. (1) Place a filter paper in a funnel; place the funnel on a flask. (2) Pour the mixture onto the filter paper. The chalk (the **residue**) stays on the paper; the water (the **filtrate**) drips through into the flask.

### 2. CENTRIFUGING

Used for separating **small amounts** of solid and liquid. Test tubes are spun very fast in a centrifuge so the solid is flung to the bottom of each tube.

### 3. EVAPORATION

Used when a solid is dissolved in a liquid. The solution is heated; the liquid evaporates; the solid remains at the bottom of the evaporating dish.

### 4. CRYSTALLISATION

Similar to evaporation, but the solid forms **crystals** as the solvent evaporates. The crystals are then left to dry.

## SEPARATING A MIXTURE OF TWO SOLIDS

(1) Dissolve one of them in an appropriate solvent. (2) Filter to remove the un-dissolved solid; (3) extract the dissolved solid by evaporating its solution.

## 5. SIMPLE DISTILLATION

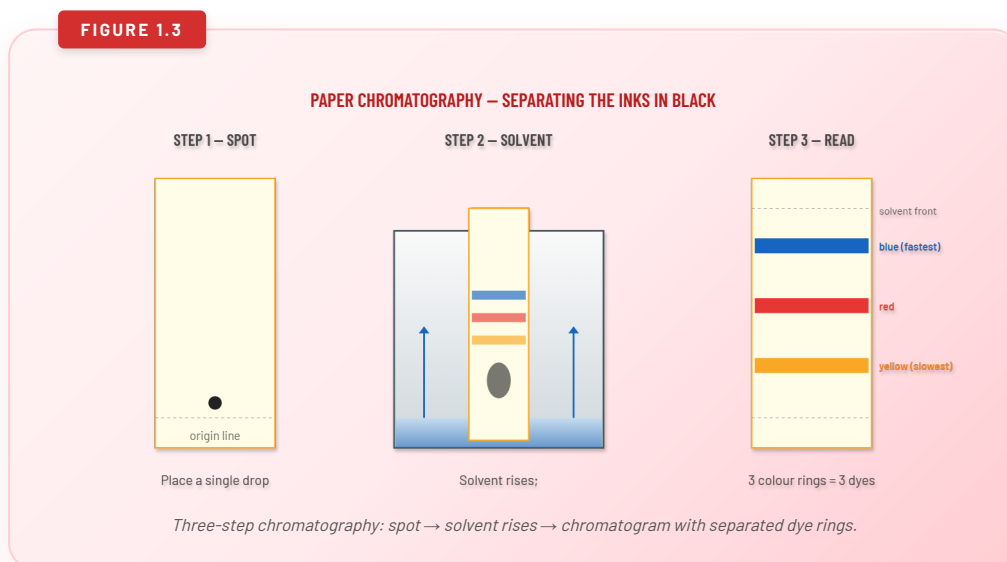
1. The impure liquid is heated.
2. It boils, and steam rises into the condenser.
3. Impurities are left behind in the flask.
4. The condenser is cold so the steam condenses to pure liquid that drops into the receiving beaker.

## 6. FRACTIONAL DISTILLATION

1. The mixture is heated.
2. The wanted substance boils and evaporates (some unwanted liquid also evaporates) and rises up the column.
3. The substance condenses on the beads in the column, heating them up.
4. Once the beads reach the boiling point of the wanted liquid, that vapour no longer condenses – it passes through the condenser and drops out as pure liquid; the unwanted liquid drips back.

## 7. CHROMATOGRAPHY

Used to separate a mixture of substances. For example, you can find how many coloured substances are in black ink. (1) Drop the black ink onto the centre of a filter paper and let it dry. (2) Drop water onto the spot one drop at a time. (3) Different dyes travel at different speeds – suppose three rings appear (yellow, red, blue) – this means the ink contains three coloured substances. The paper showing the separated rings is called a **chromatogram**.



### USES OF CHROMATOGRAPHY

- (1) Separate mixtures of substances.
- (2) Purify a substance by separating impurities from it.
- (3) Identify a substance – compare with a chromatogram of a known reference.

## Unit 2 – The Atom

*Atoms, elements, isotopes, radioactivity, electron arrangement.*

Atoms are the smallest particles. Each atom consists of a **nucleus** at the centre and a cloud of **electrons** that whizz around the nucleus.

### ELEMENT

An **element** is a substance that contains only one kind of atom.

The **periodic table** is the 'map' for elements: each element gets a symbol (e.g. K for potassium). Elements with similar properties are placed in the same numbered **group** (column). The rows are **periods**. A zig-zag line separates metals (left) from non-metals (right) – most elements are metals.

### COMPOUND

A **compound** contains atoms of different elements joined together – chemically combined. For example, carbon dioxide is a compound of carbon and oxygen (1 C and 2 O). The formula uses the element symbols: **CO<sub>2</sub>**.

## ISOTOPES AND RADIOACTIVITY

You identify an atom by its number of protons. Only sodium atoms have 11 protons, only carbon atoms have 6, etc.

### ISOTOPES

**Isotopes** are atoms of the **same element** with **different numbers of neutrons** (and therefore different masses).

Some isotopes are **radioactive** – their nucleus is unstable and eventually decays, giving out radiation as rays, tiny particles, and a lot of energy. Carbon-14 is a famous radioactive isotope that occurs naturally; the other two natural carbon isotopes (carbon-12, carbon-13) are not radioactive. The time it takes for half of a sample to decay is called the **half-life**.

## USES OF RADIATION

### 1. INDUSTRY – TRACERS IN PIPELINES

A radioisotope is added to oil or gas in a pipeline. At a leak, the radiation can be detected outside the pipe. Radioisotopes used in this way are called **tracers**.

### 2. MEDICINE – CANCER TREATMENT

Although radiation can cause cancer, it can also **cure** it. In **radiotherapy**, a radioisotope's decay rays are aimed precisely at the cancer cells, killing them while sparing healthy tissue.

### 3. ARCHAEOLOGY – CARBON-14 DATING

A tiny percentage of every living thing contains carbon-14 atoms. When the organism dies, it stops taking in new carbon – but the carbon-14 already in it slowly decays. Measuring the faint remaining radiation tells us the age of the remains.

#### ⚠ RADIATION & HEALTH

Radiation can also harm humans – causing radiation sickness. Doses must be carefully controlled.

## HOW ELECTRONS ARE ARRANGED

Electrons circle the nucleus at different distances called **energy levels** or **shells**. The further the shell from the nucleus, the higher its energy.

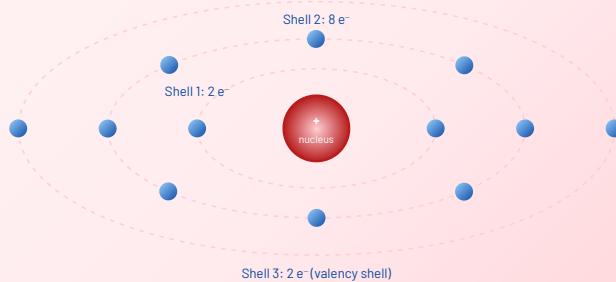
- **First shell** – up to **2** electrons
- **Second shell** – up to **8** electrons
- **Third shell** – treated as up to **8** for IGCSE / O-Level

#### 📖 ELECTRONIC CONFIGURATION

The **electronic configuration** of an atom is the arrangement of its electrons across the shells. Examples: argon = 2,8,8 · magnesium = 2,8,2 · sodium = 2,8,1.

FIGURE 2.1

#### MAGNESIUM (2,8,2)



Magnesium atom (atomic number 12). Nucleus + three shells filled in the order 2, 8, 2.

#### 💡 HOW TO READ FIGURE 2.1

The red dot at the centre is the **nucleus**. The three orbits are the **shells**. Count electrons on each shell:  $2 + 8 + 2 = 12$  electrons total – that's magnesium's atomic number, which equals its proton count in a neutral atom.

## IMPORTANT POINTS (FIVE RULES)



1. Shells fill in order from **lowest** energy to **highest**.
2. All elements in a group have the **same number** of electrons in their outer shell – called **valency electrons**.
3. The **group number** equals the number of outer-shell electrons.
4. The **period number** equals the number of shells in the atom.
5. An atom with a **full outer shell** is unreactive.

## Unit 3 – Atoms Combining (Bonding)

*Ionic, covalent, giant covalent, metallic – every bonding family in source order.*

Most elements form compounds because they want a **full outer shell**. To get one they must react with other atoms – either by giving away, gaining, or sharing electrons.

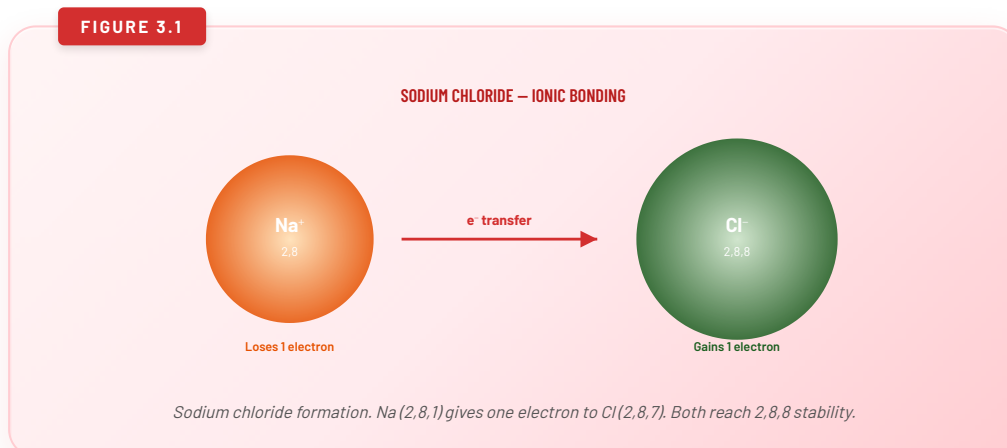
### THE IONIC BOND

Sodium has just one electron in its outer shell. It can obtain a full outer shell by **losing** that electron to another atom – making it a sodium **ion** with 10 electrons but still 11 protons, so it carries a **+1 charge**.

#### ION

An **ion** is a charged particle – an atom that has an unequal number of protons and electrons.

When sodium reacts with chlorine: sodium gives its outer electron to chlorine. Both now have full outer shells, but with charges. Sodium has **+1** (10 e<sup>-</sup> vs 11 p<sup>+</sup>); chlorine has **-1** (18 e<sup>-</sup> vs 17 p<sup>+</sup>). Opposite charges attract – the strong attraction is the **ionic bond**.

**FIGURE 3.1**

#### HOW TO READ FIGURE 3.1

Two atoms approach. 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. Result: Na<sup>+</sup> and Cl<sup>-</sup> pulled together by electrostatic attraction.

Billions of Na<sup>+</sup> and Cl<sup>-</sup> ions don't stay in pairs – they cluster so each ion is surrounded by 6 ions of opposite charge. The pattern grows into a giant **ionic lattice**. The overall charge is zero (every +1 balanced by a -1).

### ⚠ IMPORTANT NOTES ABOUT IONIC BONDING

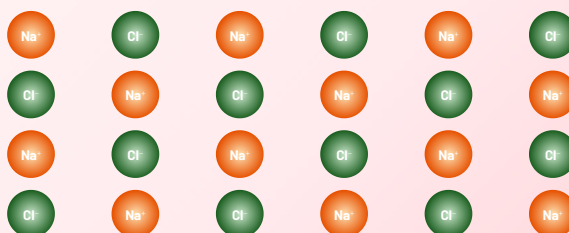
(1) Ionic bonding occurs **only between metals and non-metals**. (2) Hydrogen and metals form positive ions. (3) Non-metals form negative ions; their names end in *-ide*. (4) Group 4 and 5 rarely form ions – would have to lose/gain too many electrons. (5) Group 0 elements never form ions – already have full outer shells. (6) Some transition metals form more than one ion. (7) Some ions are formed from groups of joined atoms – **compound ions** or polyatomic ions.

## PROPERTIES OF IONIC COMPOUNDS

- **High melting and boiling points** – ionic bonds are very strong; lots of heat energy needed to break the lattice.
- **Usually soluble in water** – water molecules pull ions away from the lattice.
- **Conduct electricity when melted or dissolved** – ions become free to move and carry charge. (They do *not* conduct as solids.)

FIGURE 3.2

### IONIC LATTICE – NaCl GIANT STRUCTURE



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

*NaCl giant ionic lattice. Each Na<sup>+</sup> surrounded by 6 Cl<sup>-</sup> (and vice versa).*

### 💡 HOW TO READ FIGURE 3.2

Each ion is locked into a 3D pattern surrounded by ions of opposite charge. The drawing shows a 2D slice. Strong electrostatic attractions in every direction explain the high m.p. and the inability to conduct as a solid – ions cannot move.

## THE COVALENT BOND

Giving and losing electrons is not the only way to gain a full outer shell. Atoms can also **share** electrons. Covalent bonding is for **non-metals only** – only non-metals need to gain electrons.

### 📖 MOLECULE & BOND TYPES

A **molecule** is a group of atoms held together by covalent bonds. **Single bond** = 1 shared pair. **Double bond** = 2 shared pairs. **Triple bond** = 3 shared pairs.

## COVALENT COMPOUNDS

A **covalent compound** is when atoms of **different** elements share electrons with each other. The molecules are **not flat** – the shared pairs of electrons repel each other and the atoms spread out as far apart as possible (3D shape).

## MOLECULAR SUBSTANCES

Most molecular substances are gases or liquids at room temperature. **Molecular solids** are held in a lattice but the forces between the molecules are **weak**. The molecules are held in a regular pattern (so the solid is crystalline). Cool a molecular liquid or gas and the molecules slow down; at the freezing point they form a lattice (ice is the classic example).

## PROPERTIES OF COVALENT BONDING

- **Low melting and boiling points** – the forces *between molecules* are weak (the bonds *inside* each molecule are still strong).
- **Do not conduct electricity** – molecules are uncharged, so cannot carry charge even when melted.

## GIANT COVALENT STRUCTURES (MACROMOLECULES)

A **giant covalent structure** – or macromolecule – is made of billions of atoms bonded together in a single covalent network. The whole crystal is effectively one molecule.

### DIAMOND

Diamond is made of **carbon atoms** held in a strong lattice. Each carbon atom forms a covalent bond to **four** others. Billions of carbons bond together to form a crystal of diamond.

- **Very hard** – each atom held by 4 strong bonds
- **Very high m.p.** – the same bonds explain it
- **Cannot conduct electricity** – no free electrons to carry charge

### SILICA

**Silica (SiO<sub>2</sub>)** is similar to diamond. Each silicon atom is bonded to 4 oxygen atoms, and each oxygen is bonded to 2 silicon atoms, in a tetrahedral structure.

### GRAPHITE – A DIFFERENT GIANT STRUCTURE

Like diamond, graphite is made only of carbon – so diamond and graphite are **allotropes** of carbon (two forms of the same element). But graphite is one of the softest solids on Earth. Why? In graphite, each carbon forms a covalent bond to **three** others, giving rings of six atoms in flat layers. The fourth outer electron is **free**.

- **Soft and slippery** – layers slide over each other
- **Conducts electricity** – the free electrons can move along the layers carrying charge
- High m.p. – covalent bonds inside each layer are still strong

FIGURE 3.3

GIANT COVALENT STRUCTURES – ALLOTROPES OF CARBON + SILICA



- Hardest known solid
  - High m.p.
  - Insulator (no free e<sup>-</sup>)
- Each C bonded to

**4** others. Use: drill bits, jewellery  
**GRAPHITE** layers of hexagonal rings free e<sup>-</sup> flows along layer → · Soft & slippery · **Conducts electricity** · High m.p. Each C bonded to **3** others. Use: pencil leads, electrodes  
**SILICA (SiO<sub>2</sub>)** 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 (alternating Si-O tetrahedra).

HOW TO READ FIGURE 3.3

Look at the bond count per atom. Diamond: 4 (locked, no free electrons, hard). Graphite: 3 (one electron free per atom – conducts; layers slide – soft). Silica: tetrahedral with O between Si atoms (hard, transparent).

SUBSTANCE	PROPERTIES	USES
Diamond	Hardest known · sparkles when cut · does not conduct	Drilling and cutting tools · jewellery
Silica	Hard, can scratch · lets light through · high m.p.	Sandpaper · glass · lenses · furnace bricks
Graphite	Soft, slippery, dark · conducts electricity	Lubricant · pencil 'lead' · electrodes · brushes in generators

COMPARING BONDS

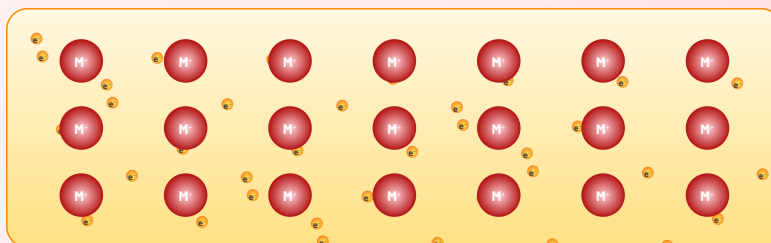
	COVALENT	IONIC	MOLECULAR IONIC
What happens	Shares electrons	Exchanges electrons	Simple molecules in a lattice
Forms between	Non-metals only	Metals and non-metals	Non-metals (with lattice)
Solubility	Dissolves in organic liquid (not water)	Dissolves in water	Varies
Conductivity	Does not conduct	Conducts when molten or aqueous	Does not conduct
m.p./b.p.	Low	High	Low

METALLIC BONDING

Metals form giant structures in which the **outer-shell electrons** of the metal atoms are **free to move**. The metallic bond is the strong force of attraction between these free electrons and the positive metal ions.

FIGURE 3.4

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. Same picture explains every metal property.

**HOW TO READ FIGURE 3.4 (ONE PICTURE, FIVE PROPERTIES)**

(1) High m.p. – the +ve 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 +ve ions slide past each other without breaking the bond. (5) Shiny when polished – the electron sea reflects light.

## PROPERTIES OF METALS

1. **High m.p.** – lots of heat energy needed to break up the lattice.
2. **Malleable** (can be bent and pressed into shapes) and **ductile** (can be drawn into wires). Layers slide without breaking the metallic bond because the electrons move with them.
3. **Good conductors of heat** – free electrons take in heat energy, move faster, transfer it through the metal.
4. **Good conductors of electricity** – free electrons move through the lattice carrying charge.

## Unit 4 – The Periodic Table

Layout, periods, groups; Group 1, Group 7, Group 0, transition metals.

The **periodic table** is a list of all the elements, in order of increasing **atomic number**. The columns are called **groups**; the rows are called **periods**.

### GROUPS

- The **group number** tells you how many electrons are in the outer shell of the atoms.
- The outer-shell electrons are also called **valency electrons** – their number shows how the elements behave.
- All elements in a group have **similar properties** (same outer-shell electron count = same chemistry).
- Group 0** elements have a *full* outer shell – this makes them unreactive.

#### SPECIAL GROUP NAMES

**Group 1** – the alkali metals. **Group 2** – the alkaline earth metals. **Group 7** – the halogens. **Group 0** – the noble gases.

### PERIODS

The **period number** gives information about the number of electron shells available in that period. **Hydrogen** sits alone in the table because it's the only element with one electron shell.

### TRENDS IN THE PERIODIC TABLE

Elements within a numbered group show **trends** in their properties. For example, as you go down Group 1 the elements become *more* reactive; as you go down Group 7 the elements become *less* reactive.

### GROUP 1 – THE ALKALI METALS

#### PHYSICAL PROPERTIES

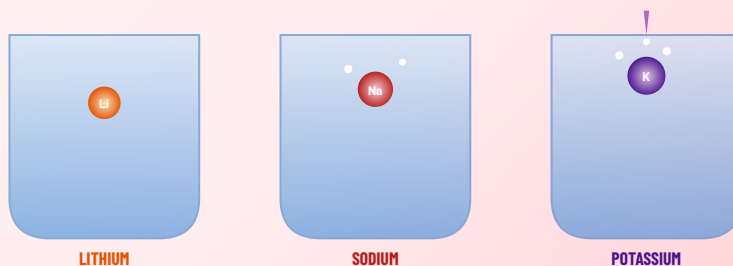
- Like all metals, they are good conductors of heat and electricity.
- They are **softer** than most other metals and have **low density**.
- They have **low** m.p. and b.p. compared to most metals.

#### CHEMICAL PROPERTIES

- All alkali metals react **vigorously with water**, releasing hydrogen gas and forming hydroxides. The hydroxides give alkaline solutions.
- They react with **non-metals**: with chlorine to make chlorides, with oxygen to make oxides.
- They form ionic compounds with metal ions of charge **+1**. The compounds are white solids that dissolve in water to colourless solutions.

FIGURE 4.1

GROUP 1 + WATER – INCREASING REACTIVITY DOWN THE GROUP



Group 1 + water. Lithium fizzes slowly. Sodium melts to a silver ball, vigorous. Potassium ignites with a lilac flame, explosive.

HOW TO READ FIGURE 4.1

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? The outer electron is further from the nucleus in larger atoms, so it's pulled less strongly – easier to give up – faster reaction.

DOWN GROUP 1	TREND
Reactivity	<b>Increases</b>
Atom size & mass	Increases
Softness	Increases (Li hardest, Cs softest)
Density	Increases
m.p. / b.p.	<b>Decreases</b>

WHY ARE THEY ALL SIMILAR?

Because they all have **one valency electron**. Atoms with the same number of outer-shell electrons react in the same way.

## GROUP 7 – THE HALOGENS

A non-metal group: F, Cl, Br, I.

- Form coloured molecules
- Are **poisonous**
- Brittle and crumbly as solids; do not conduct electricity
- **Diatomic** – exist as  $\text{F}_2$ ,  $\text{Cl}_2$ ,  $\text{Br}_2$ ,  $\text{I}_2$

FIGURE 4.2

GROUP 7 (HALOGENS) – PHYSICAL TRENDS DOWN THE GROUP



Down Group 7:  $F_2$  gas (yellow) →  $Cl_2$  gas (green) →  $Br_2$  liquid (red-brown) →  $I_2$  solid (purple-black). Reactivity DECREASES; m.p./b.p./density INCREASE; colour darkens.

HOW TO READ FIGURE 4.2

The colour bar runs yellow at top to dark purple at bottom. The red arrow shows reactivity dropping; the blue arrow shows m.p./b.p. and density rising. State change too:  $F_2$  and  $Cl_2$  are gases,  $Br_2$  is a liquid,  $I_2$  is a solid – at room temperature.

WHY REACTIVITY DECREASES DOWN GROUP 7

The smaller the atom, the easier it is for the nucleus to attract the electron needed to fill the outer shell – so the more reactive the element. Halogens are reactive because their atoms are only one electron short of a full shell.

## GROUP 0 – THE NOBLE GASES

- Colourless gases that occur naturally in air
- **Monatomic** – exist as single atoms
- **Unreactive** – full outer shell already

GAS	USE
Helium	Filling balloons and airships – lighter than air, won't catch fire
Neon	Advertising signs – glows red; mix with other gases for colour
Argon	Filling tungsten light bulbs (oxygen would burn the filament); shielding hot welded metals
Krypton	Lasers (eye surgery, car headlamps)
Xenon	Lighthouse lamps, hospital theatre lights, car headlamps

DOWN GROUP 0

Same physical trends as the halogens: atom size and density increase; boiling points rise.

## THE TRANSITION ELEMENTS

The **transition elements** are the block of 30 elements in the middle of the periodic table. They are all metals.

### PHYSICAL PROPERTIES



## CHAPTER 04 / Periodic Table · continued

- Hard, tough and strong
- **High** m.p. (mercury is an exception – liquid at room temp)
- Malleable and ductile
- Good conductors of heat and electricity
- **High density**

### CHEMICAL PROPERTIES

1. **Less reactive** than Group 1.
2. Show **no clear trend** in reactivity (unlike Group 1).
3. Most form **coloured** compounds.
4. Most form ions with different charges (**variable valency**).
5. Many form more than one compound with another element.
6. Most can form **complex ions**.

### USES OF TRANSITION METALS

- Hard, strong transition metals are used in structures – bridges, buildings, cars
- Many are used in making alloys
- Used as conductors of heat and electricity
- Many transition metals and their compounds act as **catalysts** – iron in the Haber process, nickel for hydrogenation, platinum in catalytic converters

## Unit 6 – Chemical Equations

Physical vs chemical change, reactants, products, balanced equations, state symbols.

A substance can be changed by heating it, adding water to it, mixing another substance with it, and so on. The change is either a **chemical change** or a **physical change**.

### CHEMICAL CHANGE

#### CHEMICAL CHANGE

In a **chemical change**, a new chemical substance is produced.

#### MIXTURE VS COMPOUND

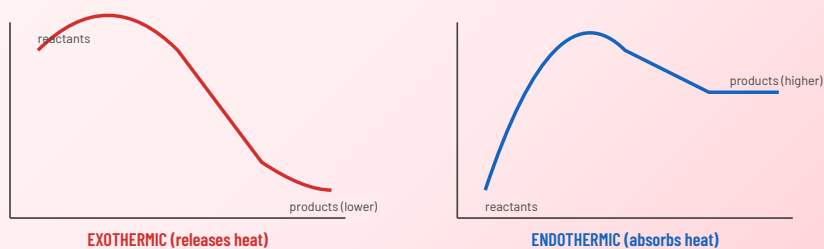
**Mixture**: 2 substances mixed together but not chemically bonded. **Compound**: 2 substances chemically bonded together.

### THREE SIGNS OF A CHEMICAL CHANGE

1. **One or more new chemical substances** are formed – usually looking different from the starting substances.
2. **Energy is taken in or given out** during the reaction. Heat-releasing change = **exothermic**; heat-absorbing change = **endothermic**.
3. **Hard to reverse** – you can't simply get the raw materials back.

FIGURE 5.1

#### ENERGY PROFILES – EXOTHERMIC vs ENDOTHERMIC



Energy profile diagrams. Exothermic reactions release heat (products lower than reactants). Endothermic reactions absorb heat (products higher).

#### HOW TO READ FIGURE 5.1

X-axis = course of reaction (left = reactants, right = products). Y-axis = energy. **Exothermic (left, red)**: products end LOWER – the energy difference is given out as heat to surroundings (the temperature rises). **Endothermic (right, blue)**: products end HIGHER – energy must be absorbed (the temperature falls). The hump in the middle is the activation energy – same in both.

### PHYSICAL CHANGE

If **no new chemical substance** is formed, the change is a **physical change** – usually reversible (e.g. ice melting, salt dissolving).

## EQUATIONS FOR CHEMICAL REACTIONS

When carbon and oxygen react, they form carbon dioxide. Carbon and oxygen are the **reactants**; carbon dioxide is the **product**.

### WORKED EQUATION 1 – CARBON + OXYGEN



1 atom of carbon + 1 molecule of oxygen → 1 molecule of carbon dioxide.

Atoms balance: 1 C, 2 O on each side.

### WORKED EQUATION 2 – HYDROGEN + OXYGEN



2 molecules of hydrogen + 1 molecule of oxygen → 2 molecules of water.

Left: 4 H atoms, 2 O atoms. Right: 4 H atoms, 2 O atoms. Balanced. ✓

## ADDING STATE SYMBOLS

You can show the state of each reactant and product by adding a state symbol after its formula.

SYMBOL	MEANING
(s)	solid
(l)	liquid
(g)	gas
(aq)	aqueous solution – dissolved in water

### STATE SYMBOLS ARE FREE MARKS

Examiners deduct for missing (s)/(l)/(g)/(aq) where they're asked for. Example:  $2\text{Na}(s) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{NaOH}(aq) + \text{H}_2(g)$ .

## Unit 8 – Acids & Alkalis

Indicators, pH, strong/weak, reactions of acids/bases, acid rain, salt preparation.

### ACIDS

You can tell if something is an **acid** by its effect on **litmus** – a purple dye that can be used as a solution or on paper. **Acids turn litmus red.**

### ALKALIS

You can tell if something is an **alkali** by its effect on litmus too. **Alkalis turn litmus blue.**

### INDICATORS

Litmus is called an **indicator** because it indicates whether something is an acid or an alkali.

### NEUTRAL SUBSTANCES

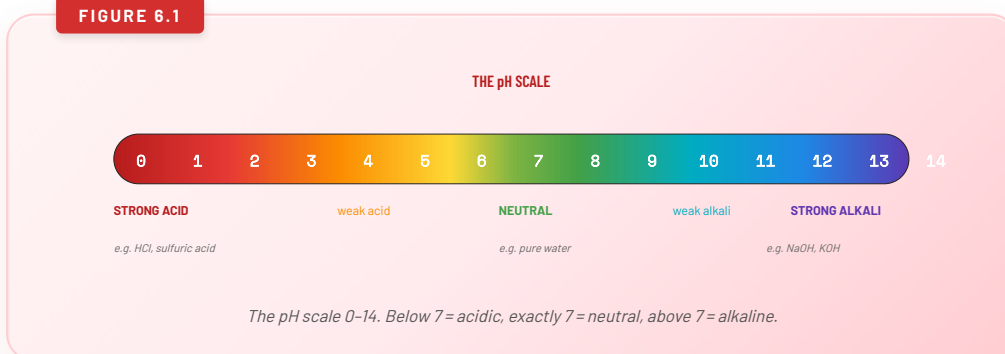
Many substances are neither acids nor alkalis – they are **neutral**. Pure water is the classic example.

### THE PH SCALE

How acidic or alkaline a solution is can be expressed on the **pH scale**, with numbers from 0 to 14.

- Acidic solution – pH less than 7
- Alkaline solution – pH greater than 7
- Neutral solution – pH exactly 7

FIGURE 6.1



**HOW TO READ FIGURE 6.1**

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× change in H<sup>+</sup> concentration. Pure water sits at exactly 7.

## ACIDS PRODUCE HYDROGEN IONS

Acidic solutions contain **hydrogen ions (H<sup>+</sup>)** – this is what makes them acidic.

## DIFFERENCE BETWEEN STRONG AND WEAK ACIDS

In a solution of a **strong acid**, ALL molecules become ions. In a solution of a **weak acid**, only some do. The higher the concentration of H<sup>+</sup>, the lower the pH, the stronger the acid.

## ALKALIS PRODUCE HYDROXIDE IONS

Alkaline solutions contain **hydroxide ions (OH<sup>-</sup>)** – this is what makes them alkaline.

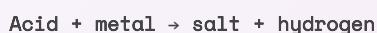
## DIFFERENCE BETWEEN STRONG AND WEAK ALKALIS

A **strong alkali** contains more OH<sup>-</sup> ions than a weak alkali. The higher the concentration of OH<sup>-</sup>, the higher the pH.

To tell strong from weak, you can also measure conductivity. A **strong acid**: high conductivity, low pH. A **weak acid**: low conductivity, higher pH. Same pattern reversed for alkalis.

## REACTION OF ACIDS WITH METALS

When an acid reacts with a metal, hydrogen is displaced, leaving a salt in solution. It is a **redox reaction**.

**GENERAL EQUATION**

## REACTION OF ACIDS WITH BASES

**Bases** are a group of compounds that react with acids and neutralise them, giving a salt and water. Bases include alkalis, insoluble metal oxides, hydroxides, and carbonates.

1. **With alkalis:** Acid + alkali → salt + water
2. **With metal oxides:** Acid + metal oxide → salt + water
3. **With carbonates:** Acid + metal carbonate → salt + water + carbon dioxide

## REACTIONS OF BASES

1. Neutralising acids, giving salt and water. With carbonates, carbon dioxide is also produced.
2. All alkalis (except ammonia) react with ammonium compounds, giving ammonia gas out.

## THE IONIC EQUATION

An **ionic equation** shows only the ions that actually take part in a reaction. It leaves out the rest. Three steps:

1. Write down all the ions present in the equation.
2. Cross out any ions that appear unchanged on both sides (these are *spectator ions*).
3. What's left is the ionic equation.

## PROTON DONORS AND ACCEPTORS

**Acids donate protons ( $H^+$ ); bases accept them.** Example: magnesium oxide is an insoluble base. The acid donates its  $H^+$  protons; the oxygen from  $MgO$  reacts with them to form water molecules.

## ACIDITY IN SOIL

Most crops grow best when the soil pH is near 7. If the soil is too acidic or too alkaline, crops grow badly or not at all. Usually **acidity** is the problem – from a lot of vegetation rotting in the soil or too much fertiliser used in the past. To reduce the acidity, the soil is treated with a base such as **limestone** ( $CaCO_3$ ), **quicklime** ( $CaO$ ), or **slaked lime** ( $Ca(OH)_2$ ).

## ACID RAIN

**Acid rain** is caused by factories, power stations, and homes burning fossil fuels. The waste gases include **sulphur dioxide** and **oxides of nitrogen**. They react with air and water to form sulphuric acid and nitric acid – both strong acids – which fall as rain.

## MAKING SALTS

You can make salts by reacting metals, insoluble bases, or soluble bases with acids.

### METHOD 1 – WITH A METAL

1. Add the zinc to the sulphuric acid in a beaker. It dissolves with hydrogen bubbles. Reaction stops when all the acid is used up.
2. Filter to remove the excess (unreacted) zinc. This leaves an aqueous solution of zinc sulphate.
3. Heat the solution to evaporate some water; cool. Crystals of zinc sulphate form.

### METHOD 2 – WITH AN INSOLUBLE BASE

Same procedure as Method 1 but using a metal oxide (the metal itself wouldn't react with the acid).

### METHOD 3 – WITH AN ALKALI (SOLUBLE BASE)

## CHAPTER 06 / Acids & Alkalis · continued

- Put the alkali into a flask and add a few drops of an indicator.
- Add acid from a burette – just a little at a time – swirling the flask to mix.
- When the indicator changes colour (e.g. methyl orange turns green), stop adding acid.
- Calculate how much acid was used.
- Repeat the experiment without indicator**, using the same volumes (the indicator would make the salt impure).
- Heat the solution; crystals form.

## INSOLUBLE SALTS BY PRECIPITATION

Not all salts are soluble. **Insoluble salts** can be made by **precipitation** – mixing two soluble salts that react to form an insoluble product.

### WORKED EXAMPLE – PREPARING BARIUM SULPHATE

- Make solutions of barium chloride and magnesium sulphate.
- Mix them. A white precipitate of barium sulphate forms at once.
- Filter the mixture – the precipitate is trapped in the filter paper.
- Rinse the precipitate by running distilled water through it.
- Place it in a warm oven to dry.

#### PICK THE RIGHT PAIR

To precipitate an insoluble salt MX, you need a soluble salt of  $M^+$  (use a nitrate – always soluble) AND a soluble salt of  $X^-$  (use a Group 1 or ammonium salt – always soluble). Mix and the MX precipitates.

	SOLUBLE	INSOLUBLE
Group 1, ammonium salts	All	–
Nitrates	All	–
Chlorides	Most	Silver, lead
Sulphates	Most	Calcium, barium, lead
Carbonates	Group 1 + ammonium only	All others

## Unit 10 – How Fast Reactions Are

Rates of reaction, collision theory, factors, photochemistry, catalysts and enzymes.

### RATES OF REACTION

Some reactions are **fast**, some are **slow**. **Rate** is a measure of how fast or slow something is.

#### WHAT IS RATE?

**Rate** is a measure of the change that happens in a single unit of time. To find the rate of a reaction you measure either: the amount of **reactant used up** per unit time, OR the amount of **product produced** per unit time.

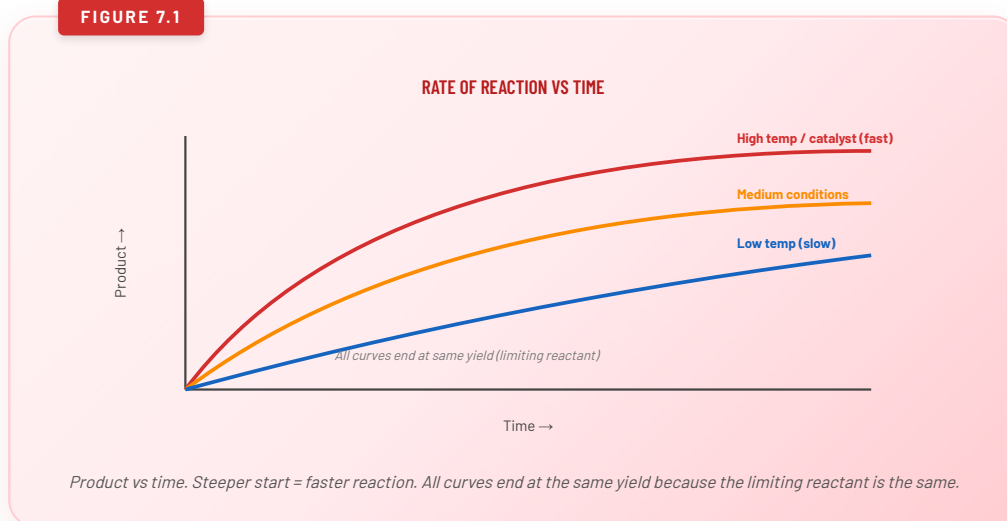
### A REACTION THAT PRODUCES A GAS

When you react magnesium and hydrochloric acid, hydrogen gas is produced. To measure the rate of this reaction, you collect the hydrogen in a syringe and time how much is produced. Using a **stop clock**, you can measure the amount of  $H_2$  produced in a given period of time.

### COLLISIONS

For a chemical reaction to occur, the reactant particles must **collide**. But collisions with too little energy do not produce a reaction. Particles must have enough energy for the collision to be **successful** in producing a reaction. The rate of reaction depends on the rate of **successful collisions** – the more there are per second, the faster the rate.

FIGURE 7.1



#### HOW TO READ FIGURE 7.1

Three curves: red (high temp / catalyst / large surface area – fastest), orange (medium), blue (low temp – slowest). Slope at any point = rate. All three plateau at the same height because the same amount of limiting reactant gives the same final product.

### CHANGING THE TEMPERATURE

If the temperature is increased:

- the reactant particles move more quickly
- they have more energy
- particles collide more often, and more collisions result in a reaction
- the rate of reaction **increases**

## CHANGING THE CONCENTRATION OR PRESSURE

If the concentration of a dissolved reactant is increased, or the pressure of a reacting gas is increased:

- the reactant particles become more crowded
- there is a greater chance of particles colliding
- the rate of reaction **increases**

## CHANGING THE SURFACE AREA

If a solid reactant is broken into small pieces or ground into a powder:

- its surface area increases
- more particles are exposed to the other reactant
- there are more collisions
- the rate of reaction **increases**

## THE EFFECT OF LIGHT

Some chemical reactions obtain the energy from **light**. They are called **photochemical reactions**.

1. **Silver bromide** is pale yellow but darkens on exposure to light because the light causes it to decompose to silver:  $2AgBr \rightarrow 2Ag + Br_2$  (used in old photographic film).
2. **Photosynthesis**: plants use carbon dioxide from air to make glucose, using sunlight energy. The green substance **chlorophyll** in leaves catalyses the reaction:  $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$

In both reactions, the stronger the light, the more energy it provides – the faster the reaction goes.

## EFFECT OF CATALYSTS

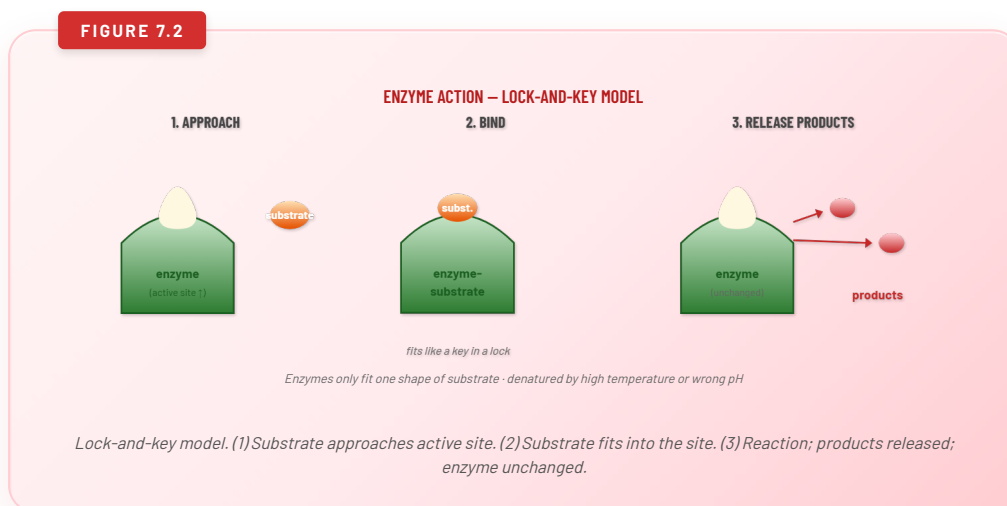
A **catalyst** is a substance that increases the rate of a reaction. The catalyst itself remains **unchanged** at the end of the reaction. Only a very small amount is needed to increase the rate of a reaction between large amounts of reactants. A catalyst works by lowering the **activation energy** – providing an alternative pathway.

## ENZYMES – BIOLOGICAL CATALYSTS

**Enzymes** are proteins that act as catalysts. They are often called **biological catalysts**.

## HOW ENZYMES WORK

First the enzyme and the reactant molecule fit together like jigsaw pieces. The reactant molecule has to be the **right shape**. The enzyme breaks the molecule down into smaller pieces (or builds two pieces into a larger one), then releases the products and is ready for another reactant.



### HOW TO READ FIGURE 7.2

Three stages left-to-right. The green enzyme has a notch – the *active site*. Only the right-shaped orange substrate locks into it. Once locked, the enzyme breaks the substrate apart. The red products drop away; the enzyme returns to its original shape, ready for the next molecule.

## IMPORTANT NOTES ABOUT ENZYMES

- An enzyme works best in conditions that match those in the living cells it came from.
- Most enzymes work best in the temperature range **25–45 °C**.
- If the temperature is too high, the enzyme loses its 3D shape and is **denatured** – the active site no longer fits the substrate.
- An enzyme also works best in a particular **pH range**.

## USES OF ENZYMES

1. **Making ethanol** – yeast contains zymase that catalyses fermentation.
2. **Making bread** – yeast enzymes break sugar into CO<sub>2</sub>, which makes the bread rise.
3. **Biological detergents** – protease and lipase enzymes break protein and fat stains at low wash temperatures.

## Unit 12 – The Behaviour of Metals

*Properties, reactivity, the reactivity series, thermal decomposition, uses.*

Most elements are metals.

### PROPERTIES OF METALS

1. They are strong
2. They are malleable
3. They are ductile
4. They are **sonorous** (make a ringing noise when struck)
5. They are shiny when polished
6. They are good conductors of heat and electricity
7. They have high m.p. and b.p. (all solid at room temperature except mercury)
8. They have high density (heavy)
9. They react with oxygen to form oxides
10. When they react, metals form **positive ions**

Properties 9 and 10 are **chemical**; the rest are physical.

### METALS REACTIVITY

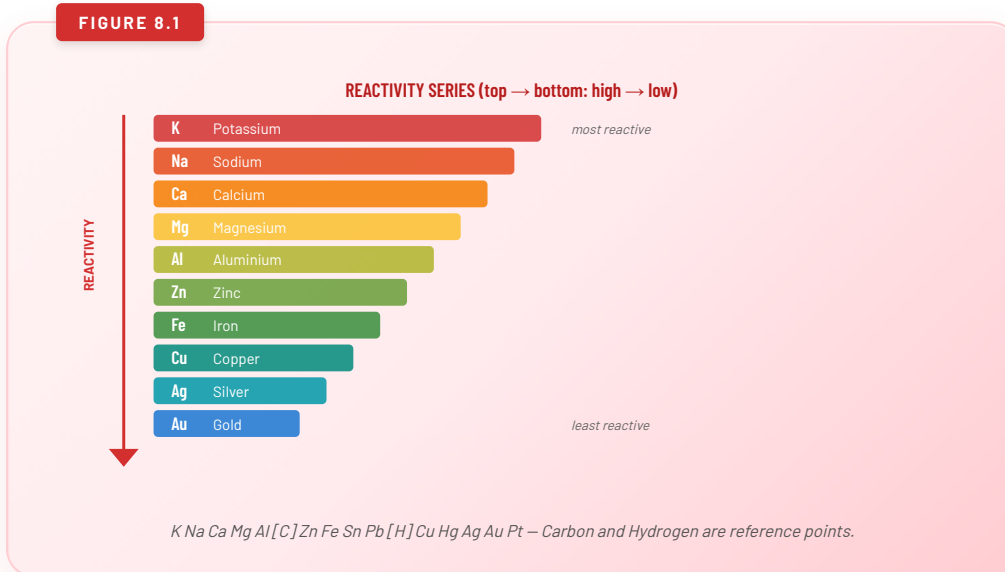
A reactive element has a strong drive to become a compound. It reacts readily with other elements. If a metal is more reactive than another, it **displaces** the less reactive metal from a compound.

When a metal is heated with the oxide of a less reactive metal, it acts as a **reducing agent**. The reaction always gives out heat – it is **exothermic**. (A reducing agent is a substance that brings about the reduction of another substance.)

A metal will always displace a less reactive metal from solutions of its compounds.

### THE REACTIVITY SERIES

FIGURE 8.1

**HOW TO READ FIGURE 8.1**

Most reactive at top, least at bottom. **Above hydrogen:** metal displaces H from acids ( $H_2$  bubbles released). **Below hydrogen:** metal cannot – copper, silver, gold won't bubble in acid. **Above carbon:** metal's oxide cannot be reduced by carbon – you must use electrolysis. **Below carbon:** cheaper carbon-reduction (blast furnace) extracts the metal.

**THINGS TO REMEMBER ABOUT THE REACTIVITY SERIES**

- The reactivity series lists metals in order of their drive to form positive ions. The more reactive the metal, the more easily it gives up electrons.
- A metal will react with a compound of a less reactive metal (e.g. an oxide) by pushing the less reactive metal out and taking its place as ions.
- The more reactive the metal, the more **stable** its compounds.
- The more reactive the metal, the **harder** it is to extract from ores.
- The less reactive the metal, the less it likes to form compounds – gold and silver occur as the pure metal in nature.

**THE STABILITY OF SOME METAL COMPOUNDS (THERMAL DECOMPOSITION)****THERMAL DECOMPOSITION**

**Thermal decomposition** is the breakdown of a compound by heating.

**1. CARBONATES**

Most decompose to oxide and  $CO_2$  on heating.  $CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$ . **But** the carbonates of potassium and sodium do **not** decompose. Strong heating is needed for calcium carbonate (the reaction is also reversible). The further down the reactivity series, the more easily the carbonate breaks down.

**2. HYDROXIDES**

Most decompose to oxide and water:  $Zn(OH)_2(s) \rightarrow ZnO(s) + H_2O(l)$ . **But** potassium and sodium hydroxides do **not** decompose. The further down, the more easily others break down.

**3. NITRATES**

All nitrates decompose on heating – but not all to the same products. Potassium and sodium nitrates  $\rightarrow$  nitrites + oxygen only:  $2NaNO_3(s) \rightarrow 2NaNO_2(s) + O_2(g)$ . Other metal nitrates  $\rightarrow$  metal oxides + brown  $NO_2$  gas + oxygen:  $2Pb(NO_3)_2 \rightarrow 2PbO(s) + 4NO_2(g) + O_2(g)$ . The further down, the more easily.

## USES OF REACTIVITY SERIES

- **The thermite process** – aluminium reduces iron oxide:  $2\text{Al} + \text{Fe}_2\text{O}_3 \rightarrow 2\text{Fe} + \text{Al}_2\text{O}_3$  (welding railway tracks)
- **Sacrificial protection of iron** – bolt blocks of more reactive metal (Mg or Zn) onto iron; they corrode in preference
- **Galvanizing** – coating iron with zinc; zinc is more reactive so it rusts first and protects the iron
- **Making cells (batteries)** – pair of metals far apart on the series gives the largest voltage

## Unit 13 – Making Use of Metals

Ores, extraction, blast furnace, steel, alloys, corrosion and rust prevention.

### METAL ORES

METAL	ORE
Sodium	Rock salt
Aluminium	Bauxite
Iron	Haematite

Where a metal sits on the reactivity series determines how it's extracted:

- **Above carbon** (K, Na, Ca, Mg, Al) – **electrolysis** – ores difficult to decompose, expensive method
- **Below carbon** (Zn, Fe, Pb, Sn) – **heating with a reducing agent** (carbon or carbon monoxide)
- **Bottom** (Cu, Ag, Au) – often occur naturally as the pure element – only separation from impurities is needed

### EXTRACTION OF ZINC FROM ZINC BLENDE

Zinc blende is mainly **zinc sulphide (ZnS)**. First it is roasted in air, giving zinc oxide:

#### STEP 1 – ROASTING



Then the oxide is reduced by carbon (or carbon monoxide) in a furnace:

#### STEP 2 – REDUCTION

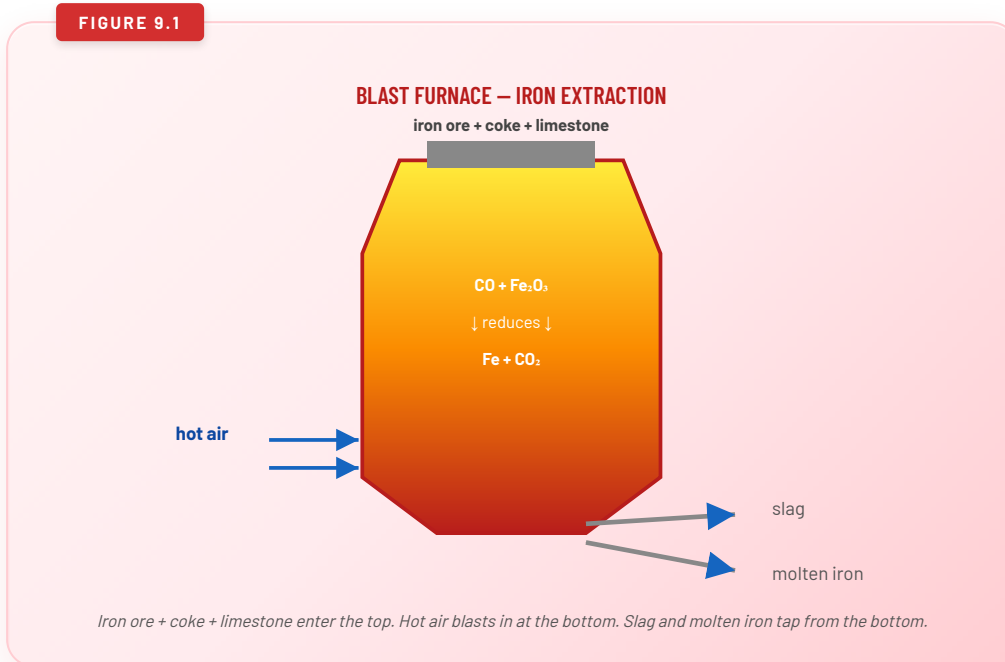


The mixture contains zinc and a slag of impurities. The zinc is separated by **fractional distillation** – it boils at 907 °C, so vaporises and is condensed at the top of the furnace.

Alternatively, zinc can be extracted by **electrolysis**.

### EXTRACTION OF IRON – THE BLAST FURNACE

FIGURE 9.1



**HOW TO READ FIGURE 9.1**

Three stages happen at different heights in the furnace. Stage 1 at the bottom (very hot): coke burns. Stage 2 going up: CO is made. Stage 3 (mid-height): CO reduces the iron ore. Iron sinks to the bottom; slag floats on top of it – both are tapped off.

**STAGE 1 – THE COKE BURNS GIVING OFF HEAT**

Hot air starts the coke burning. It reacts with oxygen, giving carbon dioxide. Combustion is a redox reaction: the carbon is oxidised to carbon dioxide. The reaction is **exothermic** – provides heat that drives the rest of the furnace.

**STAGE 1**



**STAGE 2 – CARBON MONOXIDE IS MADE**

CO<sub>2</sub> reacts with more coke, giving carbon monoxide. CO<sub>2</sub> loses oxygen – it is reduced. The reaction is **endothermic** – takes heat from the furnace, which is good because Stage 3 needs a lower temperature.

**STAGE 2**



**STAGE 3 – IRON(III) OXIDE IS REDUCED**

This is the actual extraction. CO reacts with the iron ore, giving liquid iron. CO acts as the **reducing agent**; it reduces the iron(III) oxide while it itself is oxidised to CO<sub>2</sub>. The iron trickles to the bottom of the furnace.

**STAGE 3**



**WHAT IS THE LIMESTONE FOR?**

CHEMISTRY · 4CH1

EXAM FREQUENCY: HIGH

Syllabus ref: Unit 13

## CHAPTER 09 / Making Use of Metals · continued

The limestone reacts with sand (silica, SiO<sub>2</sub>) in the ore to form calcium silicate – **slag**. The slag runs down and floats on the iron. Silica is an **acidic oxide**; its reaction with limestone (a base) is **neutralisation**, giving calcium silicate – a salt.

### SLAG FORMATION



## THE WASTE GASES

Carbon dioxide and nitrogen come out of the top of the furnace. The CO<sub>2</sub> comes from the reduction reactions; the nitrogen is from the air blast and has not reacted.

Molten iron is tapped from the bottom. It is impure with carbon as the main impurity. Some is run into moulds to give **cast iron** (hard but brittle). Most is turned into **steel**.

## USES OF SOME METALS

METAL	USE	REASON
Aluminium	Kitchen foil	Resistant to corrosion (oxide layer protects it)
Aluminium	Cooking utensils	Resists corrosion + good heat conductor
Aluminium	Aircraft and overhead power cables	Low density and high strength
Copper	Electrical cables	Excellent electrical conductor, ductile
Zinc	Galvanising iron (coating)	More reactive than iron – protects iron from rust

## STEEL AND OTHER ALLOYS

### ALLOY

An **alloy** is a mixture of metals (or with a small amount of non-metal) that changes their properties or increases their range of uses.

Pure iron is too soft, stretches easily, and rusts. Adding 0.5% carbon gives **mild steel** – hard and strong. Adding nickel and chromium gives **stainless steel** – hard and rustproof.

ALLOY	USES (MUST KNOW)
Mild steel	Buildings, ships, car bodies, machinery
Stainless steel	Car parts, kitchen sinks, cutlery

## MAKING STEELS

## CHAPTER 09 / Making Use of Metals · continued

- Remove impurities:** pour the molten iron from the blast furnace into an oxygen converter. Add calcium oxide. Blast oxygen through the molten iron. The CaO neutralises acidic impurities (Si, P, S) to form a slag that is skimmed off. The oxygen burns away the others.
- Add other elements:** nickel, chromium, manganese, etc., are measured carefully to give steels with the required properties.

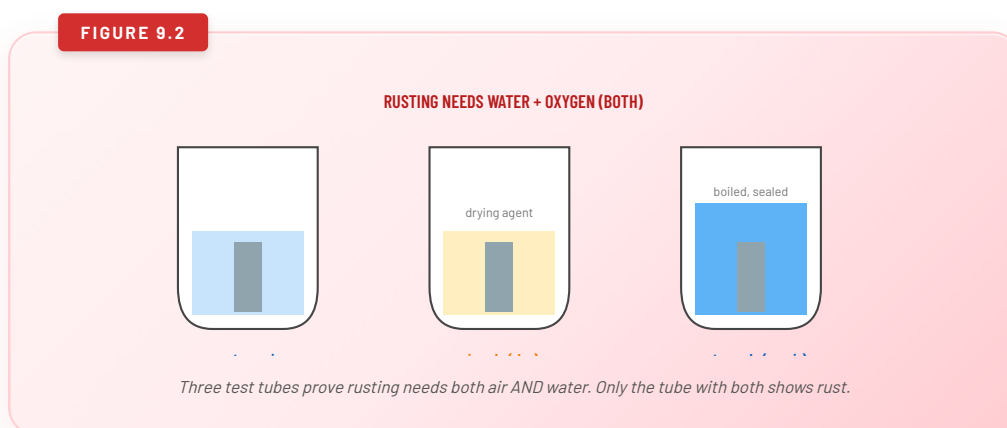
## CORROSION

### CORROSION

**Corrosion** is when a metal is attacked by air, water, or other substances in its surroundings. The more reactive the metal, the more readily it corrodes.

### WHAT DOES RUSTING INVOLVE?

Rusting needs **both air AND water**. The classic three-tube experiment proves it.



### HOW TO READ FIGURE 9.2

Tube 1: nail in tap water + air → **rusts**. Tube 2: nail in air with drying agent (no water) → no rust. Tube 3: nail in boiled water sealed under oil (no air) → no rust. Both air AND water are needed; either alone is not enough.

## HOW TO PREVENT RUSTING

- Coat the metal** to keep out air and moisture. Options: paint, grease, plastic, another metal.
- Sacrificial protection** – attach a more reactive metal to the iron; it corrodes *instead of* the iron.

### COATING WITH OTHER METALS

- Zinc** – iron is dipped into molten zinc. Called **galvanizing**.
- Tin** – deposited on steel by electrolysis. Called **tin plating**.
- Chromium** – deposited by electrolysis.

## DOES ALUMINIUM CORRODE?

**No.** A coat of aluminium oxide forms on the surface and acts as a seal preventing further corrosion.

## Unit 14 – Air & Water

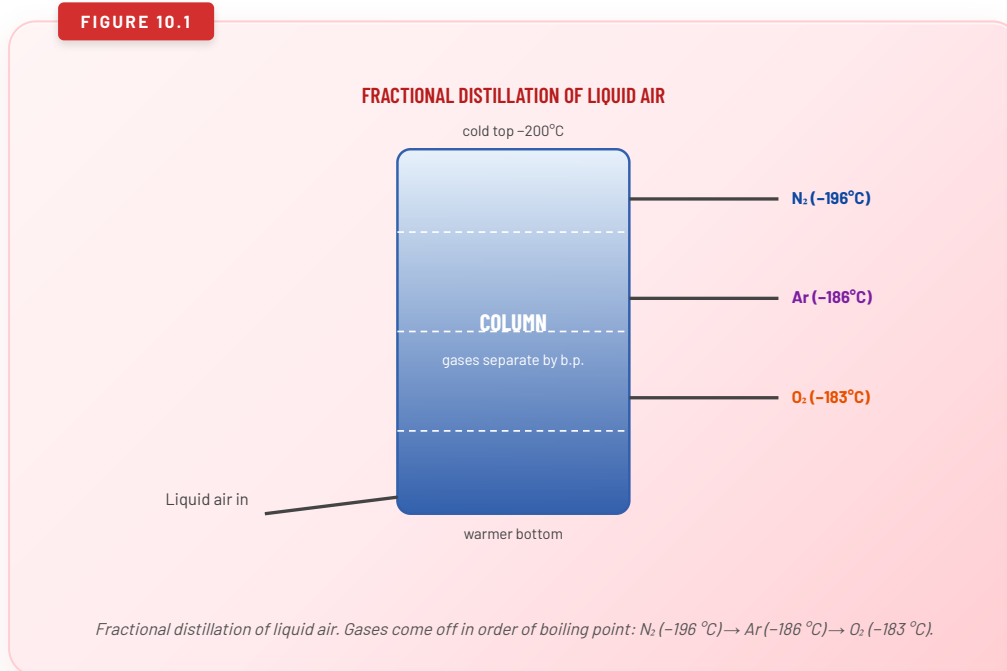
Air composition, oxygen uses, pollution, catalytic converters, water and its purification.

### AIR

Air is a mixture of gases. They can be separated by **fractional distillation** because the gases have different boiling points.

GAS	% IN AIR
Nitrogen N <sub>2</sub>	78%
Oxygen O <sub>2</sub>	21%
Argon Ar	~0.9-1%
Carbon dioxide CO <sub>2</sub>	~0.04%
Other gases (water vapour, He, Ne, Kr...)	Trace

FIGURE 10.1



#### HOW TO READ FIGURE 10.1

Air is first liquefied (cooled and compressed). The liquid feeds the bottom of a tall column – cold at top, warmer at bottom. Each gas boils off at its own boiling point and is drawn off through a tap at the matching height: N<sub>2</sub> first (lowest b.p.), then Ar, then O<sub>2</sub>.

### USES OF OXYGEN

1. **In hospitals** – people with breathing problems are given oxygen via masks or oxygen tents
2. **Welding metals** – oxygen and ethyne (acetylene) in oxy-acetylene torches

### USES OF NITROGEN

## CHAPTER 10 / Air &amp; Water · continued

- Making **ammonia** in the Haber process
- Making **nitric acid**
- Making **fertilisers**

## AIR POLLUTANTS

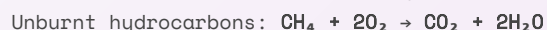
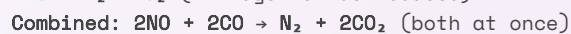
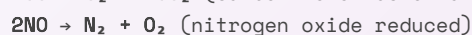
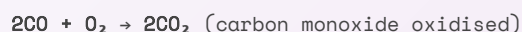
POLLUTANT	SOURCE	HARM
<b>Carbon monoxide (CO)</b>	Incomplete combustion of fuels	Poisonous – binds to haemoglobin and stops blood carrying O <sub>2</sub>
<b>Oxides of nitrogen (NO<sub>x</sub>)</b>	High-temperature combustion in car engines	Acid rain; brown haze
<b>Sulphur dioxide (SO<sub>2</sub>)</b>	Burning fossil fuels containing sulphur (coal)	Main cause of acid rain
<b>Lead compounds</b>	Leaded petrol	Brain damage
<b>Unburnt hydrocarbons</b>	Inefficient combustion of petrol	Carcinogenic

## WAYS TO REDUCE POLLUTION

- Use less fossil fuels
- Switch to clean sources of power
- Find ways to store CO<sub>2</sub> rather than letting it escape to the atmosphere

## CATALYTIC CONVERTERS

An exhaust pipe disposes of waste gases – including the harmful ones above. A **catalytic converter** sits in the exhaust and uses platinum / rhodium catalysts to react them away.

 CATALYTIC CONVERTER REACTIONS


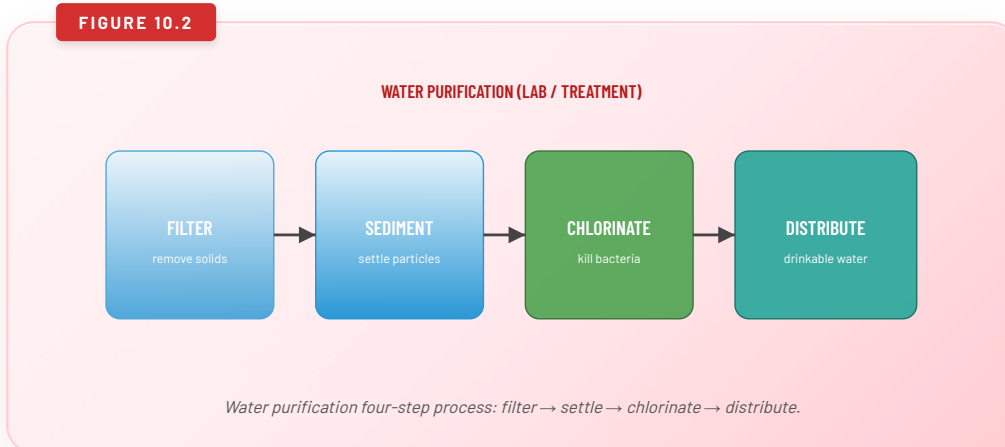
## WATER

## USES OF WATER

- **Home:** drinking, cooking, washing, flushing toilets
- **Farms:** drink for animals, watering crops
- **Industry:** solvent, washing, coolant in reaction tanks
- **Power stations:** heated to steam to drive turbines that generate electricity

## PURIFYING WATER

FIGURE 10.2



**HOW TO READ FIGURE 10.2**

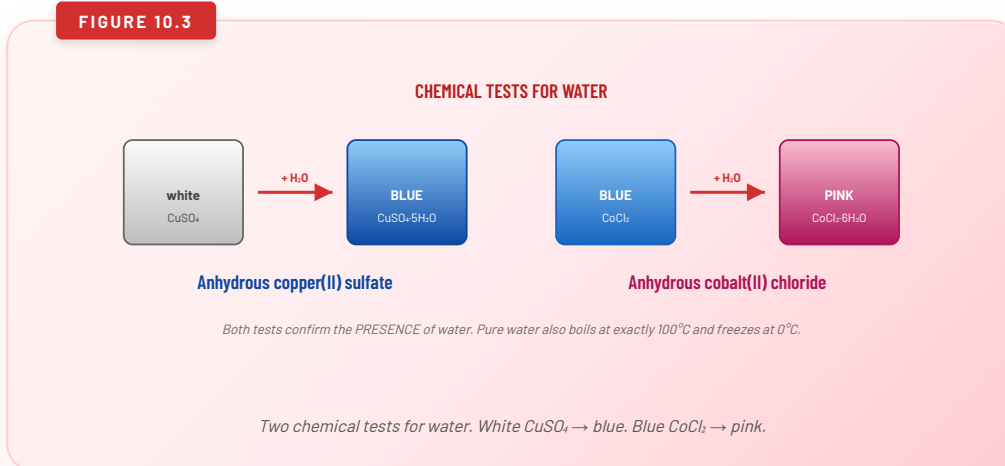
(1) Raw water is **filtered** – removes insoluble impurities. (2) Particles **settle** in tanks. (3) **Chlorination** kills bacteria. (4) Water is **distributed** as drinking water.

1. **Filtration** – remove insoluble impurities
2. **Chlorination** – kill bacteria

**CHEMICAL TESTS FOR WATER**

1. **Anhydrous copper(II) sulphate**: white → **blue** in the presence of water
2. **Anhydrous cobalt(II) chloride** paper: blue → **pink** in the presence of water

FIGURE 10.3



**TEST FOR PURE WATER**

Both above tests prove the PRESENCE of water. To prove water is PURE, also confirm it boils at exactly 100 °C and freezes at exactly 0 °C – sharp m.p./b.p. is the test for purity.

## Unit 17 – Organic Chemistry

Hydrocarbons, naming organic substances, alkanes and the homologous series.

**Organic chemistry** is the branch of chemistry connected with compounds of hydrogen and carbon (hydrocarbons).

### HYDROCARBONS

#### HYDROCARBON

A **hydrocarbon** is a class of compounds containing only hydrogen and carbon, bonded together covalently.

Hydrocarbons can form very long chains and can include double bonds (C=C) or triple bonds. They divide into two main groups:

- **Alkanes** – only C–C single bonds
- **Alkenes** – contain at least one C=C double bond

### NAMING OF ORGANIC SUBSTANCES

Naming an organic compound is a 4-step recipe:

1. **Number of carbon atoms:** 1=meth, 2=eth, 3=prop, 4=but, 5=pent, 6=hex, 7=hept, 8=oct
2. **Functional group:** -ane (single C–C), -ene (C=C double), -ol (alcohol with –OH)
3. **Position of functional group:** e.g. but-2-ene means the double bond is on the 2nd carbon
4. **Side groups:** chloro-, bromo-, methyl-, etc.

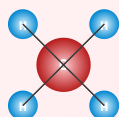
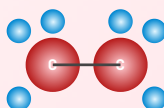
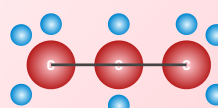
### ALKANES

#### ALKANES – GENERAL MOLECULAR FORMULA



FIGURE 11.1

## ALKANES – METHANE → PROPANE

Methane (CH<sub>4</sub>)Ethane (C<sub>2</sub>H<sub>6</sub>)Propane (C<sub>3</sub>H<sub>8</sub>)

First three alkanes: methane (CH<sub>4</sub>) → ethane (C<sub>2</sub>H<sub>6</sub>) → propane (C<sub>3</sub>H<sub>8</sub>). Each step adds one CH<sub>2</sub> unit.

**HOW TO READ FIGURE 11.1**

Each carbon (red) has 4 lines – 4 covalent bonds. Adding one carbon adds 2 hydrogens (the formula C<sub>n</sub>H<sub>2n+2</sub> emerges from this). Molecules are NOT flat; the bonds spread out tetrahedrally because each electron pair repels the others.

**PROPERTIES OF A HOMOLOGOUS SERIES**

The alkanes are a **homologous series** – meaning all members share these properties:

- Same **functional group**
- Same **general molecular formula**
- Same **chemical properties**
- Same **way of preparation**
- **Trend** in physical properties (e.g. boiling point rises with chain length)

## Synopsis, Glossary & Examiner's Tips

A final pass over the most exam-critical concepts across all 11 units of the source notebook.

### SYNOPSIS

IGCSE Chemistry at this level is built from the simplest idea – everything is made of **tiny particles** – and grows into the rich behaviour of atoms, ions, and reactions. Unit 1 trains your particle thinking through states of matter, diffusion, separation. Unit 2 opens the atom and explains isotopes. Unit 3 covers all four bonding families: ionic, covalent, giant covalent, and metallic. Unit 4 organises elements into the periodic table with predictable group trends. Units 6–8 give you the grammar of chemistry: equations, acids and bases, salt preparation. Unit 10 explains rates and the lock-and-key model of enzymes. Units 12–13 walk through the metals story – reactivity, extraction, alloys, corrosion. Unit 14 covers air and water. Unit 17 is the gateway to organic chemistry. 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.
- **Energy-level diagrams:** arrow points down for exothermic, up for endothermic. Always label reactants, products,  $\Delta H$ .
- **Metal extraction:** above carbon  $\rightarrow$  electrolysis; below carbon  $\rightarrow$  carbon reduction.
- **Rusting needs both air AND water.** The three-tube experiment proves it.

### COMPACT GLOSSARY

#### Atom

The smallest particle of an element that can take part in a chemical reaction without changing.

#### 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 physically together but not chemically combined.

#### Ion

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

#### Molecule

A group of atoms held together by covalent bonds.

#### Allotrope

Different physical forms of the same element (diamond & graphite are allotropes of carbon).

#### Oxidation

Loss of electrons (or gain of oxygen).

#### Reduction

Gain of electrons (or loss of oxygen).

#### Catalyst

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

#### Enzyme

A biological catalyst – a protein with a specifically-shaped active site.

#### Acid

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

#### Alkali

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

#### Salt

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

#### Hydrocarbon

A compound containing only carbon and hydrogen.

#### Homologous series

A family of organic compounds sharing the same general formula and gradual physical-property change.

#### Isotope

Atom of an element with a different number of neutrons (and therefore mass).

STUDENT NAME

STUDENT ID

SUBJECT

Chemistry

SUBJECT CODE

4CH1

CURRICULUM

Edexcel Pearson IGCSE / O-Level

ACADEMIC SESSION

## DECLARATION OF AUTHENTICITY

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

STUDENT SIGNATURE

INSTRUCTOR SIGNATURE

DATE