

IGCSE Chemistry 0620: full syllabus summary notes

These notes cover the whole Cambridge IGCSE Chemistry 0620 syllabus in the order of the syllabus topics. Each section gives the key facts, equations and tests you need, with short explanations of the ideas behind them. Content marked **(Extended)** is supplement material for students taking the Extended papers. Learn the boxed definitions word for word: examiners award marks for the exact wording. A Quick check with full worked answers is at the end.

In calculations these notes use the relative atomic masses normally printed on the Periodic Table in the exam: H = 1, C = 12, N = 14, O = 16, Na = 23, Mg = 24, Al = 27, S = 32, Cl = 35.5, K = 39, Ca = 40, Fe = 56, Cu = 64, Zn = 65. Always use the values from the Periodic Table printed in your own paper.

1. States of matter

1.1 Solids, liquids and gases

The three states of matter differ in the arrangement, separation and motion of their particles.

Property	Solid	Liquid	Gas
Shape	Fixed	Takes the shape of the container	Fills the whole container
Volume	Fixed	Fixed	Not fixed, can be compressed
Arrangement	Regular, closely packed	Irregular, close together	Random, far apart
Motion	Vibrate about fixed positions	Move around each other (slide past)	Move quickly and randomly in all directions
Energy of particles	Lowest	Higher	Highest

1.2 Changes of state

Melting (solid to liquid), boiling and evaporation (liquid to gas), condensation (gas to liquid) and freezing (liquid to solid) are physical changes. No new substance is made.

- **Melting:** energy supplied makes the particles vibrate more until they overcome the forces holding them in fixed positions.
- **Boiling:** happens throughout the liquid at a fixed temperature, the boiling point. Bubbles of gas form inside the liquid.
- **Evaporation:** happens only at the surface of a liquid and at any temperature below the boiling point. The most energetic particles escape from the surface.
- **Condensation and freezing** release energy to the surroundings as the particles move closer together and attractive forces form.



Heating and cooling curves: when a pure solid is heated, its temperature rises until it reaches the melting point. The temperature then stays constant (a flat section on the graph) while the solid melts, because the energy supplied is used to overcome the forces between particles rather than to raise the temperature. A second flat section appears at the boiling point. A cooling curve shows the same flat sections at the same temperatures, where energy is released as the substance condenses and freezes.

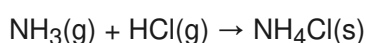
1.3 Effect of temperature and pressure on gases

- At constant pressure, increasing the temperature increases the volume of a gas. The particles move faster and hit the container walls harder and more often, so the gas expands.
- At constant temperature, increasing the pressure decreases the volume of a gas. The particles are pushed closer together.

1.4 Diffusion

Diffusion: the net movement of particles from a region of higher concentration to a region of lower concentration, as a result of their random motion.

Diffusion is fast in gases, slow in liquids and almost zero in solids. **(Extended)** At the same temperature, a gas with a lower relative molecular mass diffuses faster, because its particles move faster on average. In the classic experiment, cotton wool soaked in concentrated ammonia solution is placed at one end of a long glass tube and cotton wool soaked in concentrated hydrochloric acid at the other. A white ring of ammonium chloride forms nearer the hydrochloric acid end, because ammonia ($M_r = 17$) diffuses faster than hydrogen chloride ($M_r = 36.5$).



2. Atoms, elements and compounds

2.1 Elements, compounds and mixtures

Element: a substance made of only one type of atom.

Compound: a substance containing two or more elements chemically combined.

Mixture: two or more substances that are not chemically combined.

A compound has fixed proportions of its elements and different properties from them. For example, sodium is a reactive metal and chlorine is a toxic gas, but sodium chloride is table salt. A mixture can be separated by physical methods such as filtration and distillation, and its composition can vary.

2.2 Atomic structure

Particle	Relative mass	Relative charge	Where found
Proton	1	+1	Nucleus
Neutron	1	0	Nucleus

Electron	1/1840	-1	Shells around the nucleus
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Proton number (atomic number): the number of protons in the nucleus of an atom.

Mass number (nucleon number): the total number of protons and neutrons in the nucleus of an atom.

An atom is neutral, so the number of electrons equals the number of protons. Number of neutrons = mass number – proton number. For example, an atom of $^{23}_{11}\text{Na}$ has 11 protons, 11 electrons and $23 - 11 = 12$ neutrons.

2.3 Electronic configuration

Electrons fill shells from the nucleus outwards. For the first 20 elements: the first shell holds up to 2 electrons, the second up to 8, and the third holds 8 before the fourth shell starts filling. For example, oxygen is 2,6, chlorine is 2,8,7 and calcium is 2,8,8,2.

- The number of outer-shell electrons equals the group number (for Groups I to VII).
- The number of occupied shells equals the period number.
- Noble gases have full outer shells, which explains why they are unreactive. Helium has 2 outer electrons; the others have 8.

2.4 Isotopes

Isotopes: different atoms of the same element that have the same number of protons but different numbers of neutrons.

Isotopes of an element have the same chemical properties because they have the same number of outer-shell electrons (and the same electronic configuration). Their physical properties, such as density, can differ slightly.

(Extended) The relative atomic mass, A_r , is the average mass of the isotopes of an element compared with 1/12 of the mass of an atom of carbon-12. It is calculated from the abundances of the isotopes.

Worked example: chlorine is 75% chlorine-35 and 25% chlorine-37. Calculate its relative atomic mass.

$$A_r = (75 \times 35 + 25 \times 37) \div 100 = (2625 + 925) \div 100 = 3550 \div 100 = 35.5$$

2.5 Ions and ionic bonding

Ion: an electrically charged particle formed when an atom or group of atoms loses or gains electrons.

Ionic bond (Extended): a strong electrostatic attraction between oppositely charged ions.

Metal atoms lose their outer electrons to form positive ions (cations). Non-metal atoms gain electrons to form negative ions (anions). Both reach a noble gas electronic configuration. For example, in sodium chloride, sodium (2,8,1) loses one electron to become Na^+ (2,8), and chlorine (2,8,7) gains one electron to become Cl^- (2,8,8). Dot-and-cross diagrams show the transferred electron as a cross on

the chloride ion, with square brackets and the charge around each ion.

(Extended) Ionic compounds form a **giant lattice**: a regular three-dimensional arrangement of alternating positive and negative ions. Properties of ionic compounds:

- High melting and boiling points, because a lot of energy is needed to overcome the strong electrostatic attractions between the many ions in the lattice.
- They conduct electricity when molten or dissolved in water, because the ions are then free to move. They do not conduct when solid, because the ions are held in fixed positions.
- Many are soluble in water.

2.6 Simple molecules and covalent bonds

Covalent bond: a bond formed when a pair of electrons is shared between two atoms, leading to noble gas electronic configurations.

Covalent bonds form between non-metal atoms. You should be able to draw dot-and-cross diagrams (outer shells only) for H_2 , Cl_2 , H_2O , CH_4 , NH_3 and HCl , and **(Extended)** for N_2 (a triple bond), O_2 (a double bond), C_2H_4 (a $C=C$ double bond), CO_2 (two $C=O$ double bonds) and CH_3OH .

Simple molecular substances have:

- Low melting and boiling points. **(Extended)** The covalent bonds inside each molecule are strong, but the intermolecular forces between molecules are weak, so little energy is needed to separate the molecules.
- Poor electrical conductivity, because there are no ions or electrons free to move.

2.7 Giant covalent structures

In a giant covalent (macromolecular) structure, every atom is joined to others by covalent bonds in a continuous network.

- **Diamond**: each carbon atom forms four covalent bonds in a tetrahedral network. Very hard, very high melting point, does not conduct electricity (no free electrons). Used in cutting tools.
- **Graphite**: each carbon atom forms three covalent bonds in flat layers of hexagons. The layers are held together by weak forces, so they slide over each other: graphite is soft and slippery, and is used as a lubricant. Each carbon has one electron not used in bonding, which is delocalised and free to move along the layers, so graphite conducts electricity. It is used for electrodes.
- **(Extended) Silicon(IV) oxide, SiO_2** : each silicon atom is bonded to four oxygen atoms and each oxygen atom to two silicon atoms. Its structure is similar to diamond, so it is hard, has a high melting point and does not conduct.

2.8 Metallic bonding (Extended)

Metallic bonding: the electrostatic attraction between the positive ions in a giant metallic lattice and a "sea" of delocalised electrons.

- Good electrical conductivity: the delocalised electrons are free to move through the structure.

- Malleable and ductile: the layers of positive ions can slide over each other without the bonding being broken, because the delocalised electrons still hold them together.
- High melting and boiling points: the attraction between the ions and the delocalised electrons is strong.

3. Stoichiometry

3.1 Formulae and equations

A chemical formula shows the elements in a compound and the ratio of their atoms. The molecular formula gives the actual number of each type of atom in one molecule, for example C_2H_6 . The empirical formula is the simplest whole-number ratio of atoms, for example CH_3 .

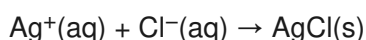
The formula of an ionic compound is worked out from the charges on its ions, so that the total charge is zero. Learn these ions:

Positive ions	Negative ions
H^+ , Na^+ , K^+ , Ag^+ , NH_4^+	Cl^- , Br^- , I^- , OH^- , NO_3^-
Mg^{2+} , Ca^{2+} , Cu^{2+} , Zn^{2+} , Fe^{2+} , Ba^{2+} , Pb^{2+}	O^{2-} , S^{2-} , SO_4^{2-} , CO_3^{2-} , SO_3^{2-}
Al^{3+} , Fe^{3+} , Cr^{3+}	PO_4^{3-}

For example, aluminium oxide is Al_2O_3 ($2 \times 3+ = 6+$ balances $3 \times 2- = 6-$), and calcium hydroxide is $Ca(OH)_2$.

Balanced equations have the same number of each type of atom on both sides. State symbols: (s) solid, (l) liquid, (g) gas, (aq) aqueous, meaning dissolved in water.

(Extended) Ionic equations show only the ions that change. Spectator ions, which are present on both sides unchanged, are left out. For example, when silver nitrate solution is added to sodium chloride solution:



3.2 Relative masses

Relative atomic mass, A_r : the average mass of the isotopes of an element compared with 1/12 of the mass of an atom of carbon-12.

Relative molecular mass, M_r : the sum of the relative atomic masses of all the atoms in a molecule (the term relative formula mass is used for ionic compounds, with the same symbol M_r).

For example, M_r of $CaCO_3 = 40 + 12 + (3 \times 16) = 100$. M_r of $Ca(OH)_2 = 40 + 2 \times (16 + 1) = 74$.

3.3 The mole (Extended)

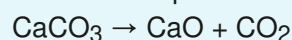
Mole: the unit of amount of substance. One mole contains 6.02×10^{23} particles (the Avogadro constant).

The key relationships:

- amount (mol) = mass (g) ÷ molar mass (g/mol)
- volume of gas (dm³) = amount (mol) × 24 dm³/mol, at room temperature and pressure (r.t.p.)
- concentration (mol/dm³) = amount (mol) ÷ volume (dm³)
- concentration (g/dm³) = concentration (mol/dm³) × molar mass
- 1 dm³ = 1000 cm³, so divide cm³ by 1000 to get dm³.

Use the balanced equation to find the mole ratio of reactants and products. The steps are always: convert what you know into moles, use the ratio, then convert back into the quantity asked for.

Worked example (reacting masses and gas volumes): 25.0 g of calcium carbonate is heated until it fully decomposes. Calculate (a) the mass of calcium oxide formed, (b) the volume of carbon dioxide at r.t.p.

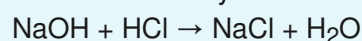


Moles of CaCO₃ = 25.0 ÷ 100 = 0.250 mol. The ratio is 1 : 1 : 1.

(a) Mass of CaO = 0.250 × 56 = 14.0 g

(b) Volume of CO₂ = 0.250 × 24 = 6.00 dm³

Worked example (titration): 25.0 cm³ of sodium hydroxide solution is neutralised by 20.0 cm³ of 0.100 mol/dm³ hydrochloric acid. Calculate the concentration of the sodium hydroxide.



Moles of HCl = 0.100 × 20.0 ÷ 1000 = 0.00200 mol

Ratio 1 : 1, so moles of NaOH = 0.00200 mol

Concentration of NaOH = 0.00200 ÷ (25.0 ÷ 1000) = 0.0800 mol/dm³

3.4 Empirical and molecular formulae (Extended)

To find an empirical formula from masses or percentages: divide each by A_r, divide by the smallest answer, and round to a whole-number ratio (multiply up if you get halves, such as 1.5). To find the molecular formula, divide M_r by the mass of the empirical formula and multiply.

Worked example: a compound contains 40.0% carbon, 6.7% hydrogen and 53.3% oxygen by mass. Its M_r is 60. Find its empirical and molecular formulae.

C: 40.0 ÷ 12 = 3.33; H: 6.7 ÷ 1 = 6.7; O: 53.3 ÷ 16 = 3.33

Divide by 3.33: C 1, H 2, O 1. Empirical formula CH₂O (mass 30).

60 ÷ 30 = 2, so the molecular formula is C₂H₄O₂.

3.5 Yield, purity and limiting reactants (Extended)

- percentage yield = (actual yield ÷ theoretical yield) × 100
- percentage purity = (mass of pure substance ÷ total mass of sample) × 100
- percentage composition by mass of an element = (A_r × number of atoms ÷ M_r) × 100

The **limiting reactant** is the one that is completely used up first; it decides how much product forms. The other reactant is in excess. To find it, work out the moles of each reactant and compare them with the ratio in the equation.

Worked example: 2.4 g of magnesium is added to 0.15 mol of hydrochloric acid. Mg + 2HCl →

$\text{MgCl}_2 + \text{H}_2$. Which reactant is limiting?

Moles of Mg = $2.4 \div 24 = 0.10$ mol. This needs 0.20 mol of HCl, but only 0.15 mol is present. So HCl is limiting and magnesium is in excess. Moles of H_2 formed = $0.15 \div 2 = 0.075$ mol.

4. Electrochemistry

4.1 Electrolysis

Electrolysis: the decomposition of an ionic compound, when molten or in aqueous solution, by the passage of an electric current.

- The **electrolyte** is the molten or aqueous ionic compound. It conducts because its ions are free to move.
- The **anode** is the positive electrode. Negative ions (anions) are attracted to it.
- The **cathode** is the negative electrode. Positive ions (cations) are attracted to it.
- Inert electrodes, made of graphite (carbon) or platinum, do not react.
- In the external circuit, electrons flow from the anode to the cathode through the wires. In the electrolyte, the charge is carried by moving ions.

Metals or hydrogen form at the cathode. Non-metals, other than hydrogen, form at the anode.

4.2 Products of electrolysis

Molten binary compounds: the metal forms at the cathode and the non-metal at the anode. For molten lead(II) bromide: lead at the cathode and red-brown bromine vapour at the anode.

Aqueous solutions: water provides small amounts of H^+ and OH^- ions as well.

- At the cathode: hydrogen is produced unless the metal is less reactive than hydrogen (for example copper or silver), in which case the metal is deposited.
- At the anode: a halogen is produced if a concentrated halide is present. Otherwise oxygen is produced from hydroxide ions (from water).

Electrolyte (inert electrodes)	Cathode product	Anode product
Molten lead(II) bromide	Lead	Bromine
Concentrated aqueous sodium chloride	Hydrogen	Chlorine
Dilute sulfuric acid	Hydrogen	Oxygen
Aqueous copper(II) sulfate	Copper	Oxygen

In the electrolysis of concentrated aqueous sodium chloride, the solution left behind is sodium hydroxide. In the electrolysis of aqueous copper(II) sulfate with graphite electrodes, the blue colour fades as copper ions are removed, and the solution becomes acidic.

Copper(II) sulfate with copper electrodes: the copper anode dissolves, forming Cu^{2+} ions, and copper is deposited on the cathode at the same rate. The colour of the solution does not change. This is how impure copper is purified: the impure copper is the anode and a thin strip of pure copper is the

cathode.

4.3 Half-equations (Extended)

Reduction (gain of electrons) happens at the cathode. Oxidation (loss of electrons) happens at the anode.

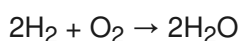
- $\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$
- $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$
- $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$
- $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$
- $4\text{OH}^- \rightarrow \text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^-$
- $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$, and at a copper anode $\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$

4.4 Electroplating

Electroplating coats an object with a thin layer of metal, to improve its appearance or its resistance to corrosion. The object to be plated is the cathode, the anode is made of the plating metal, and the electrolyte contains ions of the plating metal. For example, to copper-plate a key: key as cathode, copper anode, aqueous copper(II) sulfate electrolyte.

4.5 Hydrogen-oxygen fuel cells

A fuel cell uses hydrogen and oxygen to produce electricity, with water as the only chemical product.



(Extended) Compared with petrol engines in cars:

- Advantages: the only product is water, so no carbon dioxide or other pollutants are released by the vehicle; fuel cells are more efficient.
- Disadvantages: hydrogen is a flammable gas that is difficult to store and transport; there are few hydrogen filling stations; and most hydrogen is currently made from fossil fuels such as methane, which releases carbon dioxide.

5. Chemical energetics

5.1 Exothermic and endothermic reactions

Exothermic reaction: a reaction that transfers thermal energy to the surroundings, leading to an increase in the temperature of the surroundings.

Endothermic reaction: a reaction that takes in thermal energy from the surroundings, leading to a decrease in the temperature of the surroundings.

Examples of exothermic reactions: combustion, neutralisation, respiration, and many displacement reactions. Examples of endothermic reactions: thermal decomposition, photosynthesis, and dissolving ammonium nitrate in water.

5.2 Enthalpy change and reaction pathway diagrams

Enthalpy change, ΔH (Extended): the transfer of thermal energy during a reaction. ΔH is negative for exothermic reactions and positive for endothermic reactions.

Activation energy, E_a (Extended): the minimum energy that colliding particles must have in order to react.

A reaction pathway diagram plots energy (vertical axis) against the progress of the reaction. For an exothermic reaction the products are lower in energy than the reactants; for an endothermic reaction they are higher. The curve rises to a peak between them. The height from the reactants to the peak is the activation energy. The difference between the reactant and product levels is ΔH , shown by an arrow pointing down (exothermic) or up (endothermic).

5.3 Bond energies (Extended)

Breaking bonds is endothermic: energy is taken in. Making bonds is exothermic: energy is released.

ΔH = total energy needed to break bonds in the reactants – total energy released making bonds in the products

Worked example: calculate ΔH for $H_2 + Cl_2 \rightarrow 2HCl$, using these bond energies (in kJ/mol) as given in a question: H–H 436, Cl–Cl 242, H–Cl 431.

Bonds broken: $436 + 242 = 678$ kJ/mol

Bonds made: $2 \times 431 = 862$ kJ/mol

$\Delta H = 678 - 862 = -184$ kJ/mol. The reaction is exothermic because more energy is released making bonds than is taken in breaking them.

Bond energy values vary between data sources, so always use the values printed in the question.

6. Chemical reactions

6.1 Physical and chemical changes

A physical change, such as melting or dissolving, makes no new substance and is usually easy to reverse. A chemical change makes one or more new substances and is usually accompanied by an energy change. Signs of a chemical change include a colour change, effervescence (gas given off), a precipitate forming, or a temperature change.

6.2 Rate of reaction

The rate of a reaction increases when you increase:

- the concentration of a solution;
- the pressure of a gas;
- the surface area of a solid (use smaller pieces or a powder);
- the temperature;
- or when you add a catalyst.

Catalyst: a substance that increases the rate of a reaction and is unchanged at the end of the reaction.

(Extended) A catalyst works by providing an alternative reaction pathway with a lower activation

energy.

Enzymes: proteins that act as biological catalysts.

Explaining rate with collision theory: particles react only when they collide. Higher concentration, pressure or surface area increases the number of collisions per unit time (the collision frequency). **(Extended)** Raising the temperature makes particles move faster, so they collide more often, and also gives a greater proportion of particles energy equal to or greater than the activation energy, so a greater proportion of collisions are successful. The second effect is the more important one. A catalyst lowers the activation energy, so more collisions are successful.

Measuring rate:

- Collect a gas in a gas syringe (or an upturned measuring cylinder) and record its volume at regular time intervals.
- Stand the flask on a balance and record the loss in mass as a gas escapes. A cotton wool plug lets gas out but stops liquid spraying out.
- For reactions that form a precipitate (for example sodium thiosulfate with hydrochloric acid, which forms sulfur), time how long it takes for a cross under the flask to disappear. The shorter the time, the faster the rate.

Interpreting graphs: on a graph of volume of gas against time, the gradient is the rate. The curve is steepest at the start (highest concentration of reactants), becomes less steep as reactants are used up, and levels off when the reaction stops because one reactant has run out. A faster reaction gives a steeper curve that levels off sooner. The final volume depends only on the amount of limiting reactant, not on the rate.

6.3 Reversible reactions and equilibrium

A reversible reaction can go in both directions, shown by the symbol $\rightleftharpoons$. Examples:

- Heating hydrated copper(II) sulfate: blue crystals turn to white anhydrous copper(II) sulfate as water is lost. Adding water turns it blue again.

$$\text{CuSO}_4 \cdot 5\text{H}_2\text{O} \rightleftharpoons \text{CuSO}_4 + 5\text{H}_2\text{O}$$
- Heating hydrated cobalt(II) chloride: pink turns blue. Adding water turns it pink again.

(Extended) In a closed system, a reversible reaction reaches **equilibrium**, where the rate of the forward reaction equals the rate of the reverse reaction, and the concentrations of reactants and products no longer change.

(Extended) If a condition is changed, the position of equilibrium shifts to oppose the change:

- Increasing the temperature shifts it in the endothermic direction.
- Increasing the pressure shifts it towards the side with fewer moles of gas.
- Increasing the concentration of a reactant shifts it to the right (towards the products).
- A catalyst does not change the position of equilibrium. It only makes equilibrium reached faster.

6.4 The Haber process

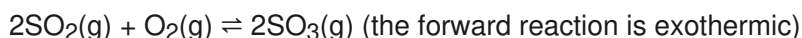
$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ (the forward reaction is exothermic)

Sources: nitrogen from the air; hydrogen from methane (natural gas). Typical conditions: 450 °C, 20 000 kPa (200 atm) and an iron catalyst. Ammonia is used to make fertilisers.

(Extended) Explaining the conditions:

- High pressure increases the yield, because the forward reaction makes fewer moles of gas (4 moles to 2). It also increases the rate. Very high pressure is not used because it is expensive and dangerous.
- A lower temperature would give a higher yield, because the forward reaction is exothermic, but the rate would be too slow. 450 °C is a compromise between a reasonable yield and a reasonable rate.
- The iron catalyst increases the rate but does not change the yield.

6.5 The Contact process



Sources: sulfur dioxide from burning sulfur or roasting sulfide ores; oxygen from the air. Typical conditions: 450 °C, 200 kPa (2 atm) and a vanadium(V) oxide catalyst, V_2O_5 . Sulfur trioxide is used to make sulfuric acid.

(Extended) The yield is already high at low pressure, so a high pressure is not worth the extra cost. The temperature is again a compromise between yield and rate.

6.6 Redox

Oxidation: gain of oxygen, loss of electrons, or an increase in oxidation number.

Reduction: loss of oxygen, gain of electrons, or a decrease in oxidation number.

Redox reaction: a reaction in which oxidation and reduction take place at the same time.

Roman numerals show the oxidation number of an element in a compound, as in iron(II) and iron(III).

(Extended) Rules for oxidation numbers: an uncombined element is 0; a simple ion has an oxidation number equal to its charge; in a compound the oxidation numbers add up to zero; in an ion containing more than one element they add up to the charge on the ion. Oxygen is usually -2 and hydrogen usually +1 in compounds.

Oxidising agent (Extended): a substance that oxidises another substance and is itself reduced.

Reducing agent (Extended): a substance that reduces another substance and is itself oxidised.

Two laboratory tests for redox:

- Acidified aqueous potassium manganate(VII) is an oxidising agent. It changes from purple to colourless when it reacts with a reducing agent.
- Aqueous potassium iodide is a reducing agent. It changes from colourless to brown (iodine forms) when it reacts with an oxidising agent.

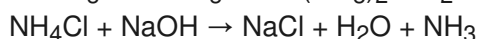
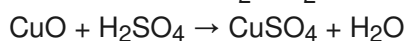
For example, in $\text{CuO} + \text{H}_2 \rightarrow \text{Cu} + \text{H}_2\text{O}$, copper(II) oxide is reduced (loses oxygen; copper goes from +2 to 0) and hydrogen is oxidised (gains oxygen; 0 to +1). Hydrogen is the reducing agent.

7. Acids, bases and salts

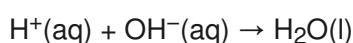
7.1 Properties of acids and bases

- Acids turn blue litmus red and have a pH below 7. They react with metals to give a salt and hydrogen, with bases to give a salt and water, and with carbonates to give a salt, water and carbon dioxide.
- Bases are metal oxides and hydroxides. A soluble base is an alkali. Alkalis turn red litmus blue and have a pH above 7. Bases react with acids to give a salt and water. Alkalis react with ammonium salts, on warming, to release ammonia.

Examples:



Acids contain H^+ ions in aqueous solution; alkalis contain OH^- ions. Neutralisation is the reaction:



(Extended) Acid: a proton (H^+) donor.

(Extended) Base: a proton (H^+) acceptor.

(Extended) Strong acid: an acid that is completely dissociated (ionised) in aqueous solution, for example hydrochloric acid.

(Extended) Weak acid: an acid that is partially dissociated in aqueous solution, for example ethanoic acid.

At the same concentration, a strong acid has a lower pH and reacts faster with magnesium or carbonates than a weak acid, because it has a higher concentration of H^+ ions.

7.2 The pH scale and indicators

The pH scale runs from 0 to 14. pH 7 is neutral; below 7 is acidic and above 7 is alkaline. The lower the pH, the higher the concentration of H^+ ions. **(Extended)** A decrease of 1 pH unit means a tenfold increase in H^+ concentration.

Indicator	In acid	In alkali
Litmus	Red	Blue
Methyl orange	Red	Yellow
Thymolphthalein	Colourless	Blue

Universal indicator shows a range of colours and is used with a colour chart to estimate pH: red for strong acid, through orange and yellow, green at neutral, to blue and purple for strong alkali.

7.3 Oxides

- Metal oxides are **basic**: they react with acids, for example CaO and CuO .
- Most non-metal oxides are **acidic**: they react with bases, for example CO_2 and SO_2 .
- **(Extended) Amphoteric** oxides react with both acids and bases to form a salt and water, for example aluminium oxide, Al_2O_3 , and zinc oxide, ZnO .
- **(Extended)** Some non-metal oxides are **neutral** and react with neither acids nor bases, for

example carbon monoxide, CO.

7.4 Preparation of salts

A salt is formed when the hydrogen ion of an acid is replaced by a metal ion or an ammonium ion. Hydrochloric acid gives chlorides, sulfuric acid gives sulfates and nitric acid gives nitrates.

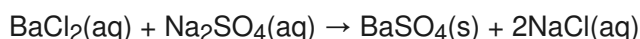
Solubility rules:

- All sodium, potassium and ammonium salts are soluble.
- All nitrates are soluble.
- Chlorides are soluble, except silver chloride and lead(II) chloride.
- Sulfates are soluble, except barium sulfate, calcium sulfate and lead(II) sulfate.
- Carbonates are insoluble, except sodium, potassium and ammonium carbonates.
- Hydroxides are insoluble, except sodium, potassium and ammonium hydroxides. Calcium hydroxide is partially soluble.

Method 1: soluble salt from an insoluble base, metal or carbonate. Warm the acid and add the solid a little at a time, stirring, until no more dissolves (excess solid remains), so that all the acid is used up. Filter to remove the excess solid. Heat the filtrate to evaporate some of the water until the solution is saturated (crystals form on a cold glass rod dipped in it). Leave it to cool and crystallise. Filter off the crystals, wash them with a little cold distilled water and dry them between filter papers.

Method 2: soluble salt from an alkali (titration). Used for sodium, potassium and ammonium salts. Use a pipette to put a known volume of alkali in a flask with a few drops of indicator. Add acid from a burette until the indicator changes colour. Repeat with the same volumes but no indicator, then crystallise as in method 1.

Method 3: insoluble salt by precipitation. Mix solutions of two soluble salts that contain the ions needed. Filter off the precipitate, wash it with distilled water and dry it. For example:



Hydrated substance: a substance that is chemically combined with water.

Anhydrous substance: a substance containing no water.

Water of crystallisation (Extended): the water molecules present in the crystals of a hydrated substance, as in $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.

8. The Periodic Table

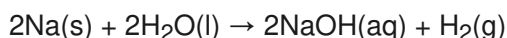
8.1 Arrangement

Elements are arranged in order of increasing proton number. Vertical columns are **groups** and horizontal rows are **periods**. Elements in the same group have the same number of outer-shell electrons and similar chemical properties. Across a period the elements change from metals on the left to non-metals on the right. **(Extended)** The Periodic Table can be used to predict the properties of an element from its position and the trends in its group, and to predict the charge on its ion: Group I gives 1+, Group II gives 2+, Group III gives 3+, Group VI gives 2– and Group VII gives 1–.

8.2 Group I: the alkali metals

Lithium, sodium and potassium are soft metals with relatively low melting points and low densities. Going down the group, the melting point decreases, the density generally increases (potassium is an exception, being less dense than sodium) and the reactivity increases.

They react with cold water to form an alkaline hydroxide solution and hydrogen:



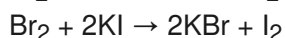
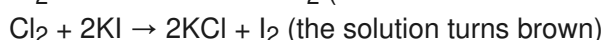
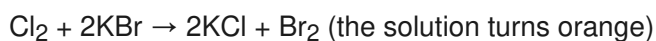
Lithium fizzes steadily. Sodium melts into a ball and moves around on the surface. Potassium reacts very vigorously, and the hydrogen burns with a lilac flame. The metals are stored under oil to keep air and water away. Reactivity increases down the group because the outer electron is further from the nucleus and more shielded, so it is lost more easily.

8.3 Group VII: the halogens

Halogen	Appearance at room temperature
Chlorine, Cl ₂	Pale yellow-green gas
Bromine, Br ₂	Red-brown liquid
Iodine, I ₂	Grey-black solid

Halogens exist as diatomic molecules. Going down the group, the density and the melting and boiling points increase, the colour gets darker, and the reactivity decreases.

A more reactive halogen displaces a less reactive halogen from a solution of its halide:



Iodine does not displace chlorine or bromine. **(Extended)** These are redox reactions. In ionic form, $\text{Cl}_2 + 2\text{Br}^- \rightarrow 2\text{Cl}^- + \text{Br}_2$: chlorine gains electrons (is reduced) and bromide ions lose electrons (are oxidised).

8.4 Transition elements

Transition elements, such as iron, copper and chromium, are metals with high densities and high melting points. They form coloured compounds and they and their compounds often act as catalysts (iron in the Haber process, vanadium(V) oxide in the Contact process). **(Extended)** They have ions with variable oxidation numbers, for example iron(II) and iron(III).

8.5 Noble gases

Group VIII (sometimes called Group 0) contains helium, neon, argon and the other noble gases. They are monatomic and unreactive because they have a full outer shell of electrons, so they have no tendency to gain, lose or share electrons. Uses: helium in balloons (low density, unreactive); argon in lamps (unreactive, stops the filament burning).

9. Metals

9.1 Properties of metals

Compared with non-metals, metals are good conductors of heat and electricity, are malleable and ductile, are usually shiny, and usually have high melting and boiling points. Non-metals are usually poor conductors (graphite is an exception) and are brittle when solid.

Chemical properties of metals: many react with dilute acids to give a salt and hydrogen, many react with oxygen to give basic oxides, and the more reactive ones react with water or steam.

9.2 Uses of metals

- **Aluminium:** aircraft parts, because of its low density and strength; food containers, because it resists corrosion; overhead power cables, because it conducts electricity well and has a low density.
- **Copper:** electrical wiring, because it is a good conductor and is ductile.

9.3 Alloys

Alloy: a mixture of a metal with other elements.

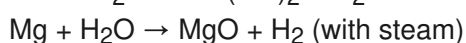
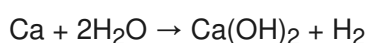
- **Brass** is a mixture of copper and zinc. It is harder than copper and resists corrosion; it is used for musical instruments, door fittings and electrical plugs.
- **Stainless steel** is a mixture of iron with other elements such as chromium, nickel and carbon. It is hard and resists rusting, so it is used for cutlery and chemical plant.
- **Mild steel** (iron with a small amount of carbon) is strong and cheap and is used for car bodies and machinery.

Alloys are often harder and stronger than the pure metals. **(Extended)** In a pure metal the ions are all the same size and are arranged in regular layers that slide over each other easily. In an alloy, atoms of a different size disrupt the regular layers, so the layers cannot slide as easily.

9.4 The reactivity series

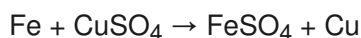
From most reactive to least reactive: potassium, sodium, calcium, magnesium, aluminium, carbon, zinc, iron, hydrogen, copper, silver, gold. (Carbon and hydrogen are non-metals included for comparison.)

Metal	Cold water	Steam	Dilute hydrochloric acid
Potassium, sodium, calcium	React to give the hydroxide and hydrogen	(Not used: too dangerous)	React violently (too dangerous to try)
Magnesium	Very slow	Reacts to give magnesium oxide and hydrogen	Reacts quickly to give a salt and hydrogen
Zinc, iron	No visible reaction	React when heated to give the oxide and hydrogen	React steadily to give a salt and hydrogen
Copper, silver, gold	No reaction	No reaction	No reaction





Displacement: a more reactive metal displaces a less reactive metal from a solution of its compound or from its oxide. For example, iron displaces copper from copper(II) sulfate solution: the blue colour fades and a pink-brown coating of copper forms.



(Extended) The reactivity series reflects how easily a metal forms its positive ion. Displacement reactions are redox: in ionic form, $\text{Fe} + \text{Cu}^{2+} \rightarrow \text{Fe}^{2+} + \text{Cu}$.

(Extended) Aluminium appears unreactive, even though it is high in the series, because it is covered by a thin, tough layer of aluminium oxide that protects it from further attack.

9.5 Corrosion of metals

Rusting is the corrosion of iron and steel. Both **water and oxygen** are needed. Rust is hydrated iron(III) oxide. Salt water speeds up rusting.

Barrier methods stop water and oxygen reaching the iron: painting, greasing or oiling, and coating with plastic.

Galvanising coats iron with zinc. The zinc acts as a barrier and also protects the iron sacrificially if the coating is scratched.

(Extended) Sacrificial protection: a more reactive metal, such as zinc or magnesium, is attached to the iron. The more reactive metal loses electrons and is oxidised in preference to the iron, so it corrodes instead. Blocks of zinc are fixed to the hulls of ships for this reason.

9.6 Extraction of metals

The method used depends on the position of the metal in the reactivity series. Metals below carbon can be extracted by heating their oxides with carbon (reduction). Metals above carbon must be extracted by electrolysis of the molten compound, which needs a lot of electrical energy. Gold and silver are so unreactive that they are found uncombined.

Iron from hematite in the blast furnace. The raw materials are hematite (iron ore, mainly Fe_2O_3), coke (carbon), limestone (calcium carbonate) and hot air.

- The coke burns in the hot air, releasing heat: $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$
- Carbon dioxide is reduced by more coke to carbon monoxide: $\text{CO}_2 + \text{C} \rightarrow 2\text{CO}$
- Carbon monoxide reduces the iron(III) oxide to molten iron: $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$
- Limestone decomposes in the heat: $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
- Calcium oxide reacts with the impurity silicon(IV) oxide to form slag: $\text{CaO} + \text{SiO}_2 \rightarrow \text{CaSiO}_3$

The molten slag floats on the molten iron and both are tapped off separately.

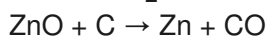
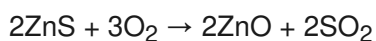
Aluminium from bauxite by electrolysis. The ore is purified to aluminium oxide, which is dissolved in molten cryolite. **(Extended)** Cryolite lowers the operating temperature (saving energy) and improves the conductivity of the electrolyte. The electrodes are graphite.

- Cathode: $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$ (molten aluminium collects at the bottom)
- Anode: $2\text{O}^{2-} \rightarrow \text{O}_2 + 4\text{e}^-$

The oxygen reacts with the hot graphite anodes to form carbon dioxide, so the anodes burn away and

must be replaced regularly.

(Extended) Zinc from zinc blende (ZnS). The ore is roasted in air to form zinc oxide, which is then reduced with carbon:



10. Chemistry of the environment

10.1 Water

Chemical tests for water: anhydrous copper(II) sulfate turns from white to blue; anhydrous cobalt(II) chloride turns from blue to pink. These tests show water is present, not that it is pure. **Physical test for purity:** pure water boils at exactly 100 °C and melts at exactly 0 °C (at standard atmospheric pressure).

Distilled water is used in practical chemistry rather than tap water, because it contains fewer chemical impurities that could interfere with results.

Water from natural sources may contain substances that are beneficial, such as dissolved oxygen for aquatic life and some mineral salts, and substances that are harmful, such as some metal compounds, plastics, sewage, harmful microbes, and nitrates and phosphates from fertilisers and detergents. Nitrates and phosphates cause excessive growth of algae, which reduces dissolved oxygen when they die and decompose.

Treatment of the domestic water supply:

- Sedimentation and filtration remove solid particles.
- Carbon (activated charcoal) removes tastes and odours.
- Chlorination kills microbes.

10.2 Fertilisers

Ammonium salts and nitrates are used as fertilisers. NPK fertilisers supply the three main elements plants need: nitrogen (N), phosphorus (P) and potassium (K), to improve plant growth and crop yield. Ammonium nitrate is made by neutralising nitric acid with ammonia: $\text{NH}_3 + \text{HNO}_3 \rightarrow \text{NH}_4\text{NO}_3$.

10.3 Air quality

Clean, dry air is approximately 78% nitrogen, 21% oxygen, and the remainder is a mixture of noble gases (mainly argon) and carbon dioxide.

Pollutant	Main source	Adverse effect
Carbon dioxide	Complete combustion of carbon-containing fuels	Increased global warming, leading to climate change
Carbon monoxide	Incomplete combustion of carbon-containing fuels	Toxic gas: reduces the blood's ability to carry oxygen
Particulates (soot, carbon)	Incomplete combustion of carbon-containing fuels	Increased risk of respiratory problems and cancer

Methane	Decomposition of vegetation and waste gases from digestion in animals	Increased global warming, leading to climate change
Oxides of nitrogen	Car engines (nitrogen and oxygen from the air react at high temperature)	Acid rain, photochemical smog and respiratory problems
Sulfur dioxide	Combustion of fossil fuels that contain sulfur compounds	Acid rain

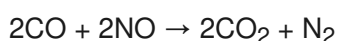
Acid rain damages buildings made of limestone and marble, corrodes metal structures, and harms plants and aquatic life.

(Extended) The greenhouse effect: the Earth's surface absorbs energy from the Sun and emits thermal (infrared) radiation. Greenhouse gases such as carbon dioxide and methane absorb this thermal energy and re-emit it in all directions, including back towards the Earth. This keeps the Earth warmer. Increasing their concentration increases global warming and leads to climate change.

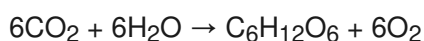
Reducing pollution:

- Climate change: plant trees, reduce livestock farming, use fewer fossil fuels and more hydrogen and renewable energy such as wind and solar.
- Acid rain: use low-sulfur fuels, and remove sulfur dioxide from power station waste gases (flue gas desulfurisation, for example using calcium oxide).
- Car exhausts: catalytic converters remove carbon monoxide, oxides of nitrogen and unburnt hydrocarbons.

(Extended) In a catalytic converter, carbon monoxide and nitrogen monoxide react to give less harmful gases:



Photosynthesis removes carbon dioxide from the air. Plants use carbon dioxide and water, with energy from light and chlorophyll, to make glucose and oxygen:



11. Organic chemistry

11.1 Formulae, functional groups and homologous series

Organic compounds can be shown by a molecular formula ($\text{C}_2\text{H}_6\text{O}$), a displayed formula (every atom and bond drawn) or **(Extended)** a structural formula, which shows each carbon atom and its attached atoms without drawing the bonds, for example $\text{CH}_3\text{CH}_2\text{OH}$.

Functional group: an atom or group of atoms that determines the chemical properties of a homologous series.

Homologous series: a family of similar compounds with similar chemical properties due to the presence of the same functional group.

Saturated compound: a compound in which molecules have only single covalent bonds between carbon atoms.

Unsaturated compound: a compound in which molecules have one or more double or triple covalent bonds between carbon atoms.

Members of a homologous series have the same general formula, differ from one member to the next by a CH_2 unit, show a trend in physical properties (for example, boiling point increases with chain length) and have similar chemical properties.

Homologous series	Functional group	General formula	Examples
Alkanes	C–C single bonds only	$\text{C}_n\text{H}_{2n+2}$	Methane, ethane, propane, butane
Alkenes	C=C	C_nH_{2n}	Ethene, propene, butene
Alcohols	–OH	$\text{C}_n\text{H}_{2n+1}\text{OH}$	Methanol, ethanol, propanol, butanol
Carboxylic acids	–COOH	$\text{C}_n\text{H}_{2n+1}\text{COOH}$	Methanoic acid, ethanoic acid, propanoic acid

Names: the stem shows the number of carbon atoms (meth = 1, eth = 2, prop = 3, but = 4) and the ending shows the series (-ane, -ene, -anol, -anoic acid). Note that in carboxylic acids the carbon in –COOH is counted: ethanoic acid is CH_3COOH .

Structural isomers (Extended): compounds with the same molecular formula but different structural formulae.

For example, butane, $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3$, and methylpropane, $\text{CH}_3\text{CH}(\text{CH}_3)\text{CH}_3$, are both C_4H_{10} . But-1-ene and but-2-ene are both C_4H_8 , with the double bond in different positions.

11.2 Fuels and petroleum

Coal, natural gas and petroleum are fossil fuels. Methane is the main constituent of natural gas. Petroleum is a mixture of hydrocarbons, compounds containing hydrogen and carbon only. It is separated by **fractional distillation**: the petroleum is heated and the vapour rises up a fractionating column, which is hotter at the bottom and cooler at the top. Each fraction condenses at a different height according to its boiling point. Fractions with small molecules have low boiling points and are collected at the top; fractions with large molecules have high boiling points and are collected at the bottom.

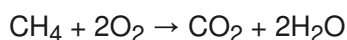
Going down the column (towards larger molecules): chain length increases, boiling point increases, and viscosity increases. Volatility decreases.

Fraction (top to bottom)	Use
Refinery gas	Gas for heating and cooking
Gasoline (petrol)	Fuel for cars
Naphtha	Chemical feedstock
Kerosene (paraffin)	Jet fuel
Diesel oil (gas oil)	Fuel for diesel engines
Fuel oil	Fuel for ships and home heating systems
Lubricating oil	Lubricants, waxes and polishes

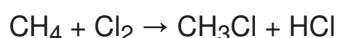
Bitumen	Making roads
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11.3 Alkanes

Alkanes are saturated hydrocarbons and are generally unreactive. Their main reaction is combustion. Complete combustion gives carbon dioxide and water; incomplete combustion, in a limited supply of oxygen, gives carbon monoxide and soot.



Alkanes undergo **substitution** reactions with chlorine in the presence of ultraviolet light. A hydrogen atom is replaced by a chlorine atom:



(Extended) A substitution reaction is one in which one atom or group of atoms is replaced by another. The UV light provides the energy to start this photochemical reaction.

11.4 Alkenes and cracking

Alkenes are unsaturated hydrocarbons made by **cracking** larger alkane molecules from petroleum, using a high temperature and a catalyst. Cracking produces alkenes and hydrogen, or alkenes and smaller alkanes. For example:



Cracking is useful because there is more demand for the smaller molecules (such as petrol) than there is supply of them from fractional distillation, and because alkenes are used to make polymers and other chemicals.

Test for unsaturation: add aqueous bromine (bromine water). An alkene decolourises it, from orange to colourless. An alkane leaves it orange.

Addition reactions of alkenes (Extended): the C=C double bond opens and only one product forms.

- With bromine: $\text{C}_2\text{H}_4 + \text{Br}_2 \rightarrow \text{C}_2\text{H}_4\text{Br}_2$ (1,2-dibromoethane)
- With hydrogen (hydrogenation), using a nickel catalyst: $\text{C}_2\text{H}_4 + \text{H}_2 \rightarrow \text{C}_2\text{H}_6$
- With steam (hydration), using an acid catalyst: $\text{C}_2\text{H}_4 + \text{H}_2\text{O} \rightarrow \text{C}_2\text{H}_5\text{OH}$

11.5 Alcohols

Ethanol, $\text{C}_2\text{H}_5\text{OH}$, is made in two ways.

- **Fermentation** of aqueous glucose by yeast, at 25 to 35 °C, in the absence of oxygen: $\text{C}_6\text{H}_{12}\text{O}_6 \rightarrow 2\text{C}_2\text{H}_5\text{OH} + 2\text{CO}_2$
- **Catalytic addition of steam to ethene**, at about 300 °C and 6000 kPa (60 atm), with a phosphoric acid catalyst.

(Extended) Comparison	Fermentation	Hydration of ethene
Raw material	Renewable (sugar from crops)	Non-renewable (ethene from petroleum)
Rate	Slow	Fast
Process	Batch	Continuous

Purity of product	Low: needs fractional distillation	High
Energy and conditions	Low temperature, atmospheric pressure	High temperature and pressure, more energy

Ethanol burns cleanly: $\text{C}_2\text{H}_5\text{OH} + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O}$. It is used as a solvent and as a fuel.

11.6 Carboxylic acids

Aqueous ethanoic acid is a weak acid and shows typical acid reactions: with metals (salt and hydrogen), bases (salt and water) and carbonates (salt, water and carbon dioxide). The salts are ethanoates, for example sodium ethanoate.

(Extended) Ethanoic acid is made by oxidising ethanol, either with acidified aqueous potassium manganate(VII) or by bacterial oxidation in the air (which is how vinegar is made).

(Extended) Esterification: a carboxylic acid reacts with an alcohol, in the presence of an acid catalyst (concentrated sulfuric acid), to form an ester and water.



Esters are named with the alcohol part first: ethanol and propanoic acid give ethyl propanoate. Esters have fruity smells and are used in flavourings and perfumes.

11.7 Polymers

Polymer: a large molecule built up from many small units called monomers.

Addition polymerisation: many alkene monomers join together, the $\text{C}=\text{C}$ bonds opening, to form a polymer with no other product. Ethene forms poly(ethene), used for plastic bags and cling film.

(Extended) You should be able to draw the repeat unit of an addition polymer from its monomer, and the monomer from a given repeat unit: the repeat unit has two carbon atoms joined by a single bond, with bonds extending out of brackets on each side and the letter n after the brackets.

(Extended) Condensation polymerisation: monomers with two functional groups join, and a small molecule, such as water, is released at each link.

- A **polyamide** (such as nylon) forms from a dicarboxylic acid and a diamine. The monomers are joined by amide links, $-\text{CONH}-$.
- A **polyester** (such as PET) forms from a dicarboxylic acid and a diol. The monomers are joined by ester links, $-\text{COO}-$. PET can be converted back into its monomers and re-polymerised, which helps recycling.
- **Proteins** are natural polyamides formed from amino acid monomers. They have the same amide linkage as nylon.

Plastics and the environment: plastics are made from petroleum. Most are non-biodegradable, so they persist in landfill sites and accumulate in oceans, harming wildlife. Burning them releases carbon dioxide and can produce toxic gases. Plastics can be reused or recycled to reduce these problems.

12. Experimental techniques and chemical analysis

12.1 Measurement and apparatus

Quantity	Apparatus
Time	Stop-watch
Temperature	Thermometer
Mass	Balance
Volume of liquid	Burette, volumetric pipette, measuring cylinder
Volume of gas	Gas syringe

Solvent: a substance that dissolves a solute.

Solute: a substance that is dissolved in a solvent.

Solution: a mixture of one or more solutes dissolved in a solvent.

Saturated solution: a solution containing the maximum concentration of a solute dissolved in the solvent at a specified temperature.

Residue: a substance that remains after evaporation, distillation, filtration or any similar process.

Filtrate: a liquid or solution that has passed through a filter.

12.2 Acid-base titrations

Use a volumetric pipette to put an exact volume (for example 25.0 cm^3) of one solution into a conical flask and add a few drops of a suitable indicator. Fill a burette with the other solution and record the starting reading. Add the solution from the burette, swirling the flask, adding dropwise near the end, until the indicator just changes colour (the end-point). Record the final reading. Repeat until you have concordant results, that is titres that agree closely (typically within 0.10 cm^3), and use their mean.

12.3 Chromatography

Paper chromatography separates a mixture of soluble substances. Draw a pencil base line (pencil does not dissolve in the solvent) and put small spots of the samples on it. Stand the paper in a solvent with the solvent level below the base line, so the spots do not dissolve into the solvent in the beaker. As the solvent moves up the paper, each substance moves a different distance depending on its solubility in the solvent.

- A pure substance gives a single spot. A mixture gives two or more spots.
- Substances can be identified by comparing their spots with those of known substances, run on the same paper with the same solvent.
- Colourless substances are made visible by spraying with a **locating agent**.

(Extended) Retention factor: $R_f = \text{distance travelled by the substance} \div \text{distance travelled by the solvent front}$. Both distances are measured from the base line. R_f is always less than 1, and a substance has a particular R_f value for a given solvent.

12.4 Separation and purification

- **Filtration:** separates an insoluble solid (the residue) from a liquid (the filtrate), for example sand from salt solution.
- **Crystallisation:** obtains a soluble solid from its solution. Heat to evaporate some of the solvent

until the solution is saturated, then cool so that crystals form, then filter and dry.

- **Simple distillation:** separates a solvent from a solution, for example pure water from sea water. The water boils, the vapour passes into a condenser and condenses, and the salt stays in the flask.
- **Fractional distillation:** separates miscible liquids with different boiling points, for example ethanol and water, or the fractions of petroleum. A fractionating column is placed between the flask and the condenser. The liquid with the lowest boiling point distils over first.

Assessing purity: a pure substance melts and boils at fixed, sharp temperatures. Impurities lower the melting point and raise the boiling point, and make them occur over a range of temperatures. Purity matters in substances people consume, such as foods and medicines, because impurities could be harmful.

12.5 Tests for cations

Cation	Aqueous sodium hydroxide	Aqueous ammonia
Aluminium, Al^{3+}	White precipitate, soluble in excess, giving a colourless solution	White precipitate, insoluble in excess
Ammonium, NH_4^+	Ammonia produced on warming (turns damp red litmus blue)	(Not applicable)
Calcium, Ca^{2+}	White precipitate, insoluble in excess	No precipitate, or a very slight white precipitate
Chromium(III), Cr^{3+}	Green precipitate, soluble in excess	Grey-green precipitate, insoluble in excess
Copper(II), Cu^{2+}	Light blue precipitate, insoluble in excess	Light blue precipitate, soluble in excess, giving a dark blue solution
Iron(II), Fe^{2+}	Green precipitate, insoluble in excess	Green precipitate, insoluble in excess
Iron(III), Fe^{3+}	Red-brown precipitate, insoluble in excess	Red-brown precipitate, insoluble in excess
Zinc, Zn^{2+}	White precipitate, soluble in excess, giving a colourless solution	White precipitate, soluble in excess, giving a colourless solution

Aluminium and zinc both give a white precipitate that dissolves in excess sodium hydroxide; aqueous ammonia tells them apart.

Flame tests: dip a clean wire in the sample and hold it in a blue Bunsen flame.

Ion	Flame colour
Lithium, Li^+	Red
Sodium, Na^+	Yellow
Potassium, K^+	Lilac
Calcium, Ca^{2+}	Orange-red
Barium, Ba^{2+}	Light green

Copper(II), Cu ²⁺	Blue-green
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12.6 Tests for anions

Anion	Test	Positive result
Carbonate, CO ₃ ²⁻	Add dilute acid, bubble the gas through limewater	Effervescence; carbon dioxide turns limewater milky
Chloride, Cl ⁻	Acidify with dilute nitric acid, then add aqueous silver nitrate	White precipitate
Bromide, Br ⁻	Acidify with dilute nitric acid, then add aqueous silver nitrate	Cream precipitate
Iodide, I ⁻	Acidify with dilute nitric acid, then add aqueous silver nitrate	Yellow precipitate
Nitrate, NO ₃ ⁻	Add aqueous sodium hydroxide, then aluminium foil; warm carefully	Ammonia produced (turns damp red litmus blue)
Sulfate, SO ₄ ²⁻	Acidify with dilute nitric acid, then add aqueous barium nitrate	White precipitate
Sulfite, SO ₃ ²⁻	Add acidified aqueous potassium manganate(VII)	Turns from purple to colourless

The acid is added first in the halide and sulfate tests to remove carbonate ions, which would otherwise also give a precipitate.

12.7 Tests for gases

Gas	Test and result
Ammonia, NH ₃	Turns damp red litmus paper blue
Carbon dioxide, CO ₂	Turns limewater milky
Chlorine, Cl ₂	Bleaches damp litmus paper
Hydrogen, H ₂	"Pops" with a lighted splint
Oxygen, O ₂	Relights a glowing splint
Sulfur dioxide, SO ₂	Turns acidified aqueous potassium manganate(VII) from purple to colourless

Exam tip: when a question asks for an observation, give what you would see, such as "white precipitate" or "fizzing", not the name of the product. When a test needs two reagents, give both and the order in which they are added.

Quick check

- Magnesium has proton number 12. (a) State the electronic configuration of a magnesium atom and of a magnesium ion. (b) Explain why magnesium forms an ion with a 2+ charge.
- Explain why graphite conducts electricity but diamond does not.

3. Calcium carbonate reacts with excess dilute hydrochloric acid: $\text{CaCO}_3 + 2\text{HCl} \rightarrow \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2$. Calculate the volume of carbon dioxide, measured at r.t.p., produced from 10.0 g of calcium carbonate.
4. 25.0 cm^3 of aqueous potassium hydroxide is neutralised by 18.0 cm^3 of 0.125 mol/dm^3 sulfuric acid. Calculate the concentration of the potassium hydroxide in mol/dm^3 .
5. Concentrated aqueous sodium chloride is electrolysed using inert electrodes. Name the product at each electrode and write the half-equation for each.
6. Use the bond energies given (in kJ/mol) to calculate the enthalpy change for $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$, and state whether the reaction is exothermic or endothermic. C–H 412, O=O 496, C=O 805, O–H 463.
7. In the Haber process, explain the effect of increasing the pressure on (a) the yield of ammonia, (b) the rate of reaction.
8. Describe how you would prepare pure, dry crystals of copper(II) sulfate from copper(II) oxide and dilute sulfuric acid.
9. Chlorine water is added to aqueous potassium iodide. Describe what you would see, write a balanced equation and explain the reaction in terms of reactivity.
10. In the blast furnace: (a) write the equation for the reduction of iron(III) oxide by carbon monoxide; (b) explain the purpose of adding limestone.
11. Ethanol can be made by fermentation or by the hydration of ethene. State the conditions for the hydration of ethene and give one advantage and one disadvantage of fermentation compared with hydration.
12. A solution of a salt gives a green precipitate with aqueous sodium hydroxide, which does not dissolve in excess. A separate sample, acidified with dilute nitric acid, gives a white precipitate with aqueous barium nitrate. Identify the salt, giving your reasoning.
13. In a chromatogram, a dye moves 3.6 cm from the base line and the solvent front moves 8.0 cm. Calculate the R_f value of the dye.

1. (a) Atom: 2,8,2. Ion: 2,8.

(b) A magnesium atom has two electrons in its outer shell. It loses these two electrons to reach the stable, full-shell configuration of a noble gas (neon). It then has 12 protons and 10 electrons, so its charge is $12 - 10 = 2+$.

2. In graphite each carbon atom forms only three covalent bonds, so one outer electron per atom is delocalised. These delocalised electrons are free to move along the layers and carry charge. In diamond each carbon atom forms four covalent bonds, so all outer electrons are used in bonding and there are no free electrons or ions to carry charge.

3. M_r of $\text{CaCO}_3 = 40 + 12 + 48 = 100$.

Moles of $\text{CaCO}_3 = 10.0 \div 100 = 0.100 \text{ mol}$.

The ratio $\text{CaCO}_3 : \text{CO}_2$ is 1 : 1, so moles of $\text{CO}_2 = 0.100 \text{ mol}$.

Volume = $0.100 \times 24 = 2.40 \text{ dm}^3$ (2400 cm^3).

4. $\text{H}_2\text{SO}_4 + 2\text{KOH} \rightarrow \text{K}_2\text{SO}_4 + 2\text{H}_2\text{O}$

Moles of $\text{H}_2\text{SO}_4 = 0.125 \times 18.0 \div 1000 = 0.00225 \text{ mol}$.

The ratio is 1 : 2, so moles of $\text{KOH} = 2 \times 0.00225 = 0.00450 \text{ mol}$.

Concentration of $\text{KOH} = 0.00450 \div (25.0 \div 1000) = 0.180 \text{ mol/dm}^3$.

5. Cathode: hydrogen. $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$. Hydrogen forms rather than sodium because sodium is more reactive than hydrogen.

Anode: chlorine. $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$. Chlorine forms because the chloride ions are concentrated. (Sodium hydroxide solution is left.)

6. Bonds broken: $4 \text{C-H} + 2 \text{O=O} = (4 \times 412) + (2 \times 496) = 1648 + 992 = 2640 \text{ kJ/mol}$.

Bonds made: $2 \text{C=O} + 4 \text{O-H} = (2 \times 805) + (4 \times 463) = 1610 + 1852 = 3462 \text{ kJ/mol}$.

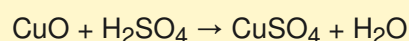
$\Delta H = 2640 - 3462 = -822 \text{ kJ/mol}$.

The reaction is exothermic, because ΔH is negative: more energy is released making the new bonds than is needed to break the old ones.

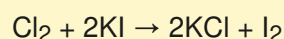
7. (a) The yield increases. The forward reaction produces fewer moles of gas (4 moles of reactant gases form 2 moles of ammonia), so increasing the pressure shifts the position of equilibrium to the right, the side with fewer gas molecules.

(b) The rate increases. The gas particles are closer together, so there are more collisions per unit time.

8. Warm the dilute sulfuric acid in a beaker. Add copper(II) oxide a little at a time, stirring, until some black solid remains undissolved (the oxide is in excess, so all the acid has reacted). Filter to remove the excess copper(II) oxide. Heat the blue filtrate to evaporate some of the water until the solution is saturated. Leave it to cool so that blue crystals form. Filter off the crystals, wash them with a little cold distilled water and dry them between filter papers.



9. The colourless solution turns brown, because iodine forms.



Chlorine is more reactive than iodine (reactivity decreases down Group VII), so chlorine displaces iodine from its compound. (Extended: chlorine is reduced to chloride ions and iodide ions are oxidised to iodine.)

10. (a) $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$

(b) Limestone decomposes to calcium oxide: $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$. The calcium oxide reacts with the acidic impurity silicon(IV) oxide to form slag: $\text{CaO} + \text{SiO}_2 \rightarrow \text{CaSiO}_3$. This removes the impurity; the molten slag floats on the iron and is run off separately.

11. Conditions: about 300°C , 6000 kPa (60 atm) and a phosphoric acid catalyst.

Advantage of fermentation: it uses a renewable raw material (sugar from crops), or it needs less energy because it runs at low temperature and atmospheric pressure.

Disadvantage of fermentation: it is slow, it is a batch process, or it produces impure ethanol that must be fractionally distilled.

12. Iron(II) sulfate, FeSO_4 .

A green precipitate that is insoluble in excess sodium hydroxide shows iron(II) ions, Fe^{2+} . A white precipitate with acidified barium nitrate shows sulfate ions, SO_4^{2-} (the precipitate is barium sulfate).

13. $R_f = \text{distance moved by the dye} \div \text{distance moved by the solvent front} = 3.6 \div 8.0 = 0.45$.