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GCSE Chemistry

Full Specification - Spec Cheat Sheet · Papers 1 & 2

AQA Chemistry (8462) | Papers 1 & 2 | 242 spec points

Every specification point for **Full Specification**, in spec order - tick each one off as you revise. Codes match the official AQA specification. Nothing here is optional.

1. Atomic structure and the periodic table

4.1.1.1	All substances are made of atoms; an atom is the smallest part of an element that can exist.
4.1.1.1	Atoms of each element are represented by a chemical symbol, e.g. O for oxygen, Na for sodium.
4.1.1.1	There are about 100 different elements, shown in the periodic table.
4.1.1.2	A compound contains two or more elements chemically combined in fixed proportions, shown by a formula.
4.1.1.2	Compounds form via reactions that involve atoms combining; reactions can be represented by word and balanced symbol equations.
4.1.1.2	Compounds are separated into elements only by chemical reactions (or sometimes electricity).
4.1.1.3	A mixture contains two or more elements or compounds not chemically combined; properties are those of the individual substances.
4.1.1.3	Mixtures can be separated by physical processes: filtration, crystallisation, simple/fractional distillation and chromatography.
4.1.1.3	These physical processes do not involve chemical reactions and no new substances are made.
4.1.1.4	The model of the atom has changed over time as new experimental evidence was found.
4.1.1.4	Before the electron, atoms were thought to be tiny indivisible spheres.
4.1.1.4	The plum pudding model showed the atom as a ball of positive charge with embedded electrons.
4.1.1.4	The alpha particle scattering experiment led to the nuclear model: a small positively charged nucleus surrounded by electrons.
4.1.1.4	Bohr proposed that electrons orbit the nucleus at specific distances (energy levels), supported by calculations.
4.1.1.4	Later work showed positive charge is made of particles (protons); the existence of the neutron was confirmed by Chadwick.
4.1.1.5	Atoms have a radius of about 0.1 nm (1 x 10^-10 m); the nucleus radius is about 1/10000 of the atom (about 1 x 10^-14 m).
4.1.1.6	The nucleus contains protons and neutrons; electrons occupy energy levels (shells) around it.
4.1.1.6	Relative charges: proton +1, neutron 0, electron -1; relative masses: proton 1, neutron 1, electron very small.
4.1.1.6	Atoms have no overall charge because the number of protons equals the number of electrons.
4.1.1.6	Atomic number = number of protons; this defines the element.
4.1.1.7	Mass number = total number of protons and neutrons in an atom.
4.1.1.7	Isotopes are atoms of the same element with the same number of protons but different numbers of neutrons.
4.1.1.7	Relative atomic mass is the average mass of atoms of an element accounting for isotope abundance.

4.1.1.7	Calculate relative atomic mass from the relative abundances of isotopes.
4.1.1.8	Electrons occupy lowest available energy levels first; the arrangement can be shown as numbers or a diagram.
4.1.1.8	Energy level (shell) capacities build outwards: 2, 8, 8 for the first 20 elements.
4.1.2.1	The periodic table arranges elements in order of atomic (proton) number.
4.1.2.1	Elements with similar properties are placed in the same vertical column (group).
4.1.2.1	Rows are called periods; the table is named after the periodic recurrence of similar properties.
4.1.2.2	Mendeleev arranged elements largely by atomic weight but left gaps and switched some orders, predicting undiscovered elements.
4.1.2.2	Discovery of isotopes explained why ordering by atomic number (not weight) gives a consistent table.
4.1.2.3	Metals are found to the left and bottom; non-metals to the right and top.
4.1.2.3	Elements that react to form positive ions are metals; those that do not are non-metals.
4.1.2.4	Group 0 (noble gases) are unreactive because they have stable, full outer electron shells.
4.1.2.4	Boiling points of noble gases increase down the group as relative atomic mass increases.
4.1.2.5	Group 1 (alkali metals) have one outer electron and react readily, losing it to form +1 ions.
4.1.2.5	Reactivity of Group 1 increases down the group; they react with water to give hydrogen and a metal hydroxide.
4.1.2.6	Group 7 (halogens) are non-metals existing as diatomic molecules; they gain one electron to form -1 ions.
4.1.2.6	Halogen reactivity decreases down the group; a more reactive halogen displaces a less reactive one from solution.
4.1.2.6	Down Group 7, relative molecular mass, melting point and boiling point increase.
4.1.3.1	Transition metals (between Groups 2 and 3) are typical metals with high melting points and densities (except mercury/group 1 comparisons).
4.1.3.1	Compared with Group 1, transition metals are harder, stronger, denser and less reactive.
4.1.3.2	Transition metals can form ions with different charges, form coloured compounds, and act as catalysts.

2. Bonding, structure, and the properties of matter

4.2.1.1	There are three types of strong chemical bond: ionic, covalent and metallic.
4.2.1.1	Ionic bonding occurs between metals and non-metals; covalent between non-metals; metallic in metals and alloys.
4.2.1.2	Ions form when atoms gain or lose electrons to achieve a stable outer shell (noble gas structure).
4.2.1.2	Metals lose electrons to form positive ions; non-metals gain electrons to form negative ions.
4.2.1.2	Use dot-and-cross diagrams to show electron transfer for Groups 1, 2, 6 and 7 elements.
4.2.1.3	Ionic compounds are giant structures held by strong electrostatic forces between oppositely charged ions, acting in all directions.
4.2.1.3	These forces of attraction are called ionic bonds.
4.2.1.4	Work out the empirical formula of an ionic compound from a diagram or from ion charges.
4.2.1.5	Covalent bonding involves shared pairs of electrons between non-metal atoms.
4.2.1.5	Covalent substances may be small molecules, large molecules (polymers) or giant covalent structures.
4.2.1.5	Represent covalent bonds using dot-and-cross diagrams, line diagrams and 3D models, recognising limitations of each.
4.2.1.6	Metallic bonding: a giant lattice of positive ions in a sea of delocalised electrons, shared across the structure.
4.2.2.1	The three states of matter are solid, liquid and gas, represented by a simple particle model.

4.2.2.1	State changes are physical; the amount of energy needed depends on the strength of forces between particles.
4.2.2.1	Stronger forces between particles mean higher melting and boiling points; limitations of the particle model should be recognised.
4.2.2.2	Use state symbols (s), (l), (g) and (aq) in equations to show the state of each substance.
4.2.2.2	Predict states of substances at given temperatures using melting and boiling point data.
4.2.3.1	Ionic compounds have high melting and boiling points due to strong electrostatic forces in the giant lattice.
4.2.3.1	Ionic compounds conduct electricity when molten or dissolved because ions are free to move and carry charge.
4.2.3.2	Small covalent molecules have low melting and boiling points due to weak intermolecular forces (the bonds themselves are strong).
4.2.3.2	Intermolecular forces increase with molecular size; simple molecular substances do not conduct electricity.
4.2.3.3	Polymers are very large molecules with strong covalent bonds in the chain and relatively strong intermolecular forces; they are solids at room temperature.
4.2.3.4	Giant covalent structures (e.g. diamond, silicon dioxide) have very high melting points; most do not conduct electricity.
4.2.3.5	Metals are good conductors of heat and electricity because delocalised electrons can move and carry charge/energy.
4.2.3.5	Metals are malleable because layers of atoms can slide over each other.
4.2.3.6	Pure metals are often too soft; alloys are mixtures with other elements whose different-sized atoms distort the layers, making them harder.
4.2.4.1	Diamond: each carbon forms four covalent bonds in a rigid giant structure, making it hard with a very high melting point; it does not conduct electricity.
4.2.4.2	Graphite: each carbon forms three bonds in layers with no covalent bonds between layers; layers slide, and one delocalised electron per atom allows conduction.
4.2.4.3	Graphene is a single layer of graphite, one atom thick, with useful electrical and mechanical properties.
4.2.4.3	Fullerenes are molecules of carbon shaped as hollow tubes or spheres, based mainly on hexagonal rings.
4.2.4.3	Buckminsterfullerene (C60) was the first fullerene discovered, with a spherical shape.
4.2.4.3	Carbon nanotubes are cylindrical fullerenes with high tensile strength and useful electrical properties; uses include nanotechnology, electronics and materials.
4.2.5.1	Nanoparticles are 1-100 nm in size (a few hundred atoms); larger than atoms/molecules but smaller than fine and coarse particles.
4.2.5.1	Coarse particles (dust) are 1x10^-5 to 2.5x10^-6 m; fine particles are 100-2500 nm.
4.2.5.2	As particle size decreases by a factor of 10, surface area to volume ratio increases by a factor of 10.
4.2.5.2	High surface area to volume ratios mean smaller quantities of nanoparticulate materials may be effective (e.g. catalysts).
4.2.5.3	Nanoparticles have uses in medicine, electronics, cosmetics, sun creams, deodorants and as catalysts.
4.2.5.3	The new properties of nanoparticles may have risks to health that are not yet fully understood; evaluate benefits and risks.

3. Quantitative chemistry

4.3.1.1	Conservation of mass: no atoms are lost or made in a reaction, so the total mass of reactants equals the total mass of products.
4.3.1.1	Balanced symbol equations have equal numbers of each type of atom on both sides.
4.3.1.2	Relative formula mass (Mr) is the sum of the relative atomic masses of the atoms in a formula.

4.3.1.2	In a balanced equation, the sum of the Mr of reactants (in the stated ratios) equals the sum of the Mr of the products.
4.3.1.3	A mass change in a non-enclosed reaction may occur if a gas is given off or a gas from the air is a reactant.
4.3.1.4	Whenever a measurement is made there is uncertainty; estimate uncertainty from the range of a set of measurements (HT).
4.3.2.1	Chemical amounts are measured in moles; the symbol for the unit is mol.
4.3.2.1	The Avogadro constant is 6.02 x 10^23 per mole; one mole of a substance contains this number of particles.
4.3.2.1	The mass of one mole of a substance in grams equals its relative formula mass; moles = mass / Mr (HT).
4.3.2.2	Use a balanced equation to calculate the masses of reactants and products from the mole ratio (HT).
4.3.2.3	Balance an equation given the masses of reactants and products by converting to moles (HT).
4.3.2.4	The limiting reactant is the one not in excess; it is completely used up and limits the amount of product (HT).
4.3.2.4	The amount of product is directly proportional to the amount of limiting reactant (HT).
4.3.3.1	Concentration of a solution can be measured in mass per volume, e.g. grams per dm^3 (g/dm^3).
4.3.3.1	Calculate the mass of solute in a given volume of a solution of known concentration; equal masses in smaller volumes give higher concentration.
4.3.4.1	Percentage yield = (mass of product made / maximum theoretical mass) x 100 (HT).
4.3.4.1	Yield is below 100% because reactions may be reversible, products are lost in handling, or side reactions occur (HT).
4.3.4.2	Atom economy = (Mr of desired product / total Mr of all products) x 100; high atom economy is better for sustainability (HT).
4.3.5	Calculate the concentration in mol/dm^3 from mass concentration and convert between cm^3 and dm^3 (chemistry only, HT).
4.3.5	Use titration data and a balanced equation to find an unknown concentration (chemistry only, HT).
4.3.6	Equal amounts in moles of gases occupy the same volume under the same conditions of temperature and pressure (chemistry only, HT).
4.3.6	The volume of one mole of any gas at room temperature and pressure (rtp) is 24 dm^3; volume = moles x 24 (chemistry only, HT).

4. Chemical changes

4.4.1.1	Metals react with oxygen to form metal oxides; oxidation is the gain of oxygen and reduction is the loss of oxygen.
4.4.1.2	The reactivity series ranks metals (and hydrogen and carbon) by their tendency to form positive ions.
4.4.1.2	Potassium, sodium, lithium, calcium, magnesium, aluminium, (carbon), zinc, iron, (hydrogen), copper, by decreasing reactivity.
4.4.1.2	Metals react with water and/or acids; the more reactive the metal, the more vigorous the reaction.
4.4.1.2	A more reactive metal displaces a less reactive metal from a compound (displacement reaction).
4.4.1.3	Unreactive metals such as gold are found native; most metals are extracted from ores in the Earth's crust.
4.4.1.3	Metals less reactive than carbon can be extracted from their oxides by reduction with carbon.
4.4.1.4	Oxidation is loss of electrons and reduction is gain of electrons (OIL RIG) (HT).
4.4.1.4	Write ionic equations for displacement reactions, identifying which species is oxidised and which is reduced (HT).
4.4.2.1	Acids produce hydrogen ions (H+) in aqueous solution; aqueous alkalis contain hydroxide ions (OH-).
4.4.2.1	The pH scale (0-14) measures acidity/alkalinity; use universal indicator or a pH probe to measure pH.
4.4.2.1	In neutralisation, hydrogen ions react with hydroxide ions to make water: H+ + OH- -> H2O.

4.4.2.2	Acids react with metals, bases (metal oxides/hydroxides) and carbonates to form salts.
4.4.2.2	Acid + metal -> salt + hydrogen; acid + metal oxide/hydroxide -> salt + water; acid + carbonate -> salt + water + carbon dioxide.
4.4.2.3	Prepare a pure, dry sample of a soluble salt from an acid and an insoluble base by reaction, filtration, crystallisation.
4.4.2.4	A strong acid is fully ionised in water; a weak acid is only partially ionised (HT).
4.4.2.4	As pH decreases by one unit, hydrogen ion concentration increases by a factor of 10 (HT).
4.4.2.4	Distinguish acid strength (degree of ionisation) from concentration (amount of acid per volume) (HT).
4.4.3.1	Electrolysis uses electricity to break down ionic compounds (electrolytes) that are molten or in solution.
4.4.3.1	Positive ions move to the cathode and negative ions move to the anode; ions are discharged at the electrodes.
4.4.3.2	Binary molten ionic compounds produce the metal at the cathode and the non-metal at the anode.
4.4.3.3	Aluminium is extracted by electrolysis of molten aluminium oxide mixed with cryolite to lower the melting point.
4.4.3.3	Carbon anodes react with oxygen produced and must be replaced regularly.
4.4.3.4	In aqueous solutions, the less reactive of hydrogen and the metal is produced at the cathode; oxygen forms at the anode unless halide ions are present.
4.4.3.4	If a halide ion is present, the halogen is produced at the anode.
4.4.3.5	Half equations show what happens at each electrode, balancing electrons gained or lost (HT).

5. Energy changes

4.5.1.1	Exothermic reactions transfer energy to the surroundings (temperature rises), e.g. combustion, oxidation, neutralisation.
4.5.1.1	Endothermic reactions take in energy from the surroundings (temperature falls), e.g. thermal decomposition and some sports cold packs.
4.5.1.1	Everyday uses include self-heating cans, hand warmers (exothermic) and instant cold packs (endothermic).
4.5.1.2	Reaction profiles show the relative energies of reactants and products and the activation energy.
4.5.1.2	Activation energy is the minimum energy needed for reactant particles to react on collision.
4.5.1.3	Energy is needed to break bonds (endothermic) and released when bonds form (exothermic) (HT).
4.5.1.3	Overall energy change = energy to break bonds - energy released making bonds; calculate using bond energies (HT).
4.5.2.1	In a chemical cell, two electrodes in an electrolyte produce a voltage depending on the metals and conditions (chemistry only).
4.5.2.1	Batteries are cells joined in series; in non-rechargeable cells the reaction stops when a reactant is used up (chemistry only).
4.5.2.1	Rechargeable cells and batteries can be recharged because the reaction is reversed by an external current (chemistry only).
4.5.2.2	Fuel cells are supplied with fuel and oxygen and produce a voltage continuously while supplied (chemistry only).
4.5.2.2	The hydrogen fuel cell: hydrogen and oxygen react to form water, releasing energy; write electrode half equations (chemistry only, HT).
4.5.2.2	Evaluate the use of hydrogen fuel cells compared with rechargeable batteries (chemistry only).

6. The rate and extent of chemical change

4.6.1.1	Rate of reaction can be found from the mean rate of formation of a product or use of a reactant over time.
4.6.1.1	Calculate mean rate as quantity of reactant used or product formed divided by time.

4.6.1.1	Find the rate at a specific time from the gradient of a tangent to the concentration-time graph (HT).
4.6.1.2	Factors affecting rate: concentration of solutions, pressure of gases, surface area of solids, temperature and the presence of a catalyst.
4.6.1.3	Collision theory: reactions occur when particles collide with enough energy (the activation energy).
4.6.1.3	Increasing temperature, concentration, pressure or surface area increases the frequency (and energy) of collisions, raising rate.
4.6.1.4	Catalysts increase rate without being used up by providing an alternative pathway with lower activation energy.
4.6.1.4	Enzymes are biological catalysts; catalysts do not change the products of a reaction.
4.6.2.1	In a reversible reaction, products can react to re-form the original reactants; shown by the ⇌ symbol.
4.6.2.2	If a reversible reaction is exothermic in one direction it is endothermic in the other, transferring equal energy.
4.6.2.3	At equilibrium in a closed system, the forward and reverse reactions occur at the same rate (HT).
4.6.2.4	The relative amounts at equilibrium depend on the conditions (HT).
4.6.2.5	Le Chatelier's Principle: if a condition is changed, the equilibrium shifts to counteract the change (HT).
4.6.2.6	Changing concentration: increasing a reactant shifts equilibrium to make more product (HT).
4.6.2.7	Changing temperature: raising it favours the endothermic direction; lowering it favours the exothermic direction (HT).
4.6.2.8	Changing pressure of gases: increasing pressure shifts equilibrium toward the side with fewer molecules (HT).

7. Organic chemistry

4.7.1.1	Crude oil is a finite resource: a mixture of hydrocarbons formed from ancient biomass (mainly plankton).
4.7.1.1	Hydrocarbons are compounds of hydrogen and carbon only; alkanes have the general formula C_nH_{2n+2}.
4.7.1.1	The first four alkanes are methane, ethane, propane and butane.
4.7.1.2	Crude oil is separated by fractional distillation into fractions with similar boiling points (chain lengths).
4.7.1.2	Fractions provide fuels and feedstock for the petrochemical industry, e.g. solvents, lubricants, polymers and detergents.
4.7.1.3	As hydrocarbon chain length increases, boiling point, viscosity and the difficulty of ignition increase, while flammability decreases.
4.7.1.4	Complete combustion of hydrocarbon fuels releases energy and produces carbon dioxide and water; carbon and hydrogen are oxidised.
4.7.2.1	Cracking breaks long-chain hydrocarbons into smaller, more useful molecules using heat (catalytic or steam cracking).
4.7.2.1	Cracking produces alkanes and alkenes; alkenes are unsaturated and are used to make polymers and other chemicals.
4.7.2.1	Supply and demand for short-chain hydrocarbons (e.g. petrol) make cracking economically important.
4.7.3.1	Alkenes contain a carbon-carbon double bond and have the general formula C_nH_{2n} (chemistry only).
4.7.3.1	Alkenes (ethene, propene, butene, etc.) are more reactive than alkanes due to the double bond and react with bromine water, decolourising it (chemistry only).
4.7.3.2	In addition reactions, the double bond opens and atoms add across it (e.g. with hydrogen, water or halogens) (chemistry only).
4.7.4.1	Alcohols (methanol, ethanol, propanol, butanol) contain the -OH functional group (chemistry only).
4.7.4.1	Alcohols dissolve in water to form neutral solutions, react with sodium, burn in air, and are oxidised by microbes/oxidising agents to form carboxylic acids (chemistry only).

4.7.4.1	Uses of alcohols include fuels and solvents; ethanol is made industrially by fermentation of sugars (chemistry only).
4.7.4.2	Carboxylic acids (methanoic, ethanoic, propanoic, butanoic) contain the -COOH group and are weak acids (chemistry only).
4.7.4.2	Carboxylic acids dissolve to form acidic solutions, react with carbonates to give carbon dioxide, and form esters with alcohols (chemistry only).
4.7.5.1	Addition polymers form from alkene monomers; the double bond opens to join monomers into long chains (chemistry only).
4.7.5.1	Deduce the polymer repeating unit from a monomer (and vice versa) for simple addition polymers (chemistry only).
4.7.5.2	Condensation polymerisation involves monomers with two functional groups and loses a small molecule (e.g. water) each time (chemistry only, HT).
4.7.5.3	Amino acids have an amine and a carboxylic acid group; they polymerise to form polypeptides (e.g. proteins) (chemistry only).
4.7.5.4	DNA is a polymer made of two strands of nucleotides forming a double helix; starch and cellulose are natural polymers of sugars (chemistry only).

8. Chemical analysis

4.8.1.1	A pure substance in chemistry is a single element or compound, not mixed with anything else.
4.8.1.1	Pure substances melt and boil at specific, sharp temperatures; impurities lower the melting point and broaden the range.
4.8.1.2	A formulation is a useful mixture with each component in a measured quantity for a specific purpose (e.g. fuels, paints, medicines).
4.8.1.3	Chromatography separates mixtures; substances separate based on their relative attraction to the stationary and mobile phases.
4.8.1.3	Calculate the Rf value = distance moved by substance / distance moved by solvent; pure substances give one spot.
4.8.2.1	Test for hydrogen: a lit splint produces a squeaky pop.
4.8.2.2	Test for oxygen: a glowing splint relights.
4.8.2.3	Test for carbon dioxide: bubbling through limewater turns it milky (cloudy white).
4.8.2.4	Test for chlorine: it bleaches damp litmus paper, turning it white (after possibly turning red).
4.8.3.1	Flame tests identify metal ions: lithium crimson, sodium yellow, potassium lilac, calcium orange-red, copper green (chemistry only).
4.8.3.2	Adding sodium hydroxide forms coloured metal hydroxide precipitates: copper(II) blue, iron(II) green, iron(III) brown (chemistry only).
4.8.3.2	Aluminium, calcium and magnesium give white precipitates with NaOH; only the aluminium hydroxide redissolves in excess (chemistry only).
4.8.3.3	Test for carbonates: add dilute acid; effervescence of carbon dioxide turns limewater milky (chemistry only).
4.8.3.3	Test for halides: add nitric acid then silver nitrate; chloride gives white, bromide cream, iodide yellow precipitate (chemistry only).
4.8.3.3	Test for sulfate: add hydrochloric acid then barium chloride; a white precipitate confirms sulfate (chemistry only).
4.8.3.4	Flame emission spectroscopy is an instrumental method that identifies metal ions and measures their concentration from line spectra (chemistry only).
4.8.3.4	Instrumental methods are accurate, sensitive and rapid, and useful for very small samples (chemistry only).

9. Chemistry of the atmosphere

4.9.1.1	Today's atmosphere is about 80% nitrogen, 20% oxygen, with small amounts of carbon dioxide, water vapour and noble gases.
4.9.1.1	Theories about the early atmosphere are limited by lack of evidence from billions of years ago.
4.9.1.1	Early volcanic activity is thought to have released gases (mainly carbon dioxide), water vapour, and some nitrogen and other gases.
4.9.1.1	As the Earth cooled, water vapour condensed to form the oceans, and carbon dioxide dissolved into them.
4.9.1.1	Algae and plants produced oxygen by photosynthesis, increasing oxygen and decreasing carbon dioxide over time.
4.9.1.1	Carbon dioxide was reduced by being locked in sedimentary rocks (e.g. limestone) and in fossil fuels.
4.9.2.1	Greenhouse gases (carbon dioxide, methane and water vapour) maintain temperatures on Earth by absorbing and re-radiating heat.
4.9.2.2	Human activities (burning fossil fuels, deforestation, farming livestock, waste) increase greenhouse gas concentrations.
4.9.2.2	Peer-reviewed evidence links increasing greenhouse gases to global climate change; consider why models may be incomplete or misreported.
4.9.2.3	Global climate change may cause rising sea levels, more extreme weather, and changes to ecosystems and food supplies.
4.9.2.4	Carbon footprint is the total greenhouse gases (as carbon dioxide equivalent) emitted over the life of a product, service or event.
4.9.2.4	Carbon footprint can be reduced by using less energy/fossil fuel, carbon capture, taxes/quotas and reducing emissions; consider limitations.
4.9.3.1	Incomplete combustion of fuels can produce carbon monoxide (a toxic, colourless, odourless gas) and soot (carbon particulates).
4.9.3.1	Burning fuels also releases sulfur dioxide and oxides of nitrogen, which cause acid rain and respiratory problems.
4.9.3.1	Particulates can cause global dimming and health problems; sulfur dioxide and nitrogen oxides come from impurities and high combustion temperatures.

10. Using resources

4.10.1.1	The Earth's resources provide warmth, shelter, food and transport; chemistry helps improve agricultural and industrial processes.
4.10.1.1	Natural resources can be supplemented or replaced by synthetic (man-made) products.
4.10.1.1	Sustainable development meets present needs without compromising future generations; chemists help reduce resource use.
4.10.1.2	Potable water is safe to drink (low dissolved salts/microbes); it is not pure water.
4.10.1.2	In the UK, fresh water is treated by filtration to remove solids and sterilisation (chlorine, ozone or UV) to kill microbes.
4.10.1.2	Where fresh water is scarce, desalination of salty water by distillation or reverse osmosis is used but needs much energy.
4.10.1.2	Required practical: analyse and purify water samples by testing pH and distillation (chemistry only).
4.10.1.3	Sewage and agricultural waste water must be treated to remove organic matter and harmful microbes before release (chemistry only).
4.10.1.3	Sewage treatment involves screening, sedimentation, anaerobic digestion of sludge and aerobic treatment of effluent (chemistry only).

4.10.1.4	Copper-rich ores are limited; alternative extraction methods include phytomining and bioleaching (chemistry only).
4.10.1.4	Phytomining uses plants to absorb metal compounds; bioleaching uses bacteria to produce leachate solutions of metal compounds (chemistry only).
4.10.1.4	Copper can be obtained from solutions by displacement using scrap iron or by electrolysis (chemistry only).
4.10.2.1	Life Cycle Assessment (LCA) assesses the environmental impact of a product across raw materials, manufacture, use and disposal.
4.10.2.1	Some LCA factors (use of water, resources, energy, waste) are measurable; assessing pollutant effects requires judgement, so LCAs can be biased.
4.10.2.2	Reducing, reusing and recycling materials cuts the use of limited resources, energy, waste and environmental impact.
4.10.2.2	Metals can be recycled by melting and recasting; quantities of materials should be matched to product requirements.
4.10.3.1	Corrosion is the destruction of materials by reaction with the environment, e.g. rusting of iron (needs water and oxygen) (chemistry only).
4.10.3.1	Corrosion can be prevented by barriers (paint, oil, greasing) or by sacrificial protection and galvanising with a more reactive metal (chemistry only).
4.10.3.2	Alloys are usually mixtures of metals (or metal with carbon) designed to be harder or more useful (chemistry only).
4.10.3.2	Examples: bronze (copper-tin), brass (copper-zinc), gold alloys measured in carats, steels (iron-carbon, sometimes with chromium/nickel), aluminium alloys (chemistry only).
4.10.3.3	Soda-lime glass is made from sand, sodium carbonate and limestone; borosilicate glass melts at higher temperatures (chemistry only).
4.10.3.3	Clay ceramics (pottery, bricks) are made by shaping wet clay and firing it; composites consist of fibres or fragments in a matrix/binder (chemistry only).
4.10.3.3	Thermosoftening polymers melt on heating (tangled chains); thermosetting polymers have cross-links and do not melt (chemistry only).
4.10.3.3	The properties of a polymer depend on the monomers used and the conditions of polymerisation, e.g. LDPE and HDPE poly(ethene) (chemistry only).
4.10.4.1	The Haber process makes ammonia from nitrogen (from air) and hydrogen (from natural gas) in a reversible reaction (chemistry only).
4.10.4.1	Haber conditions: about 450 C, 200 atmospheres pressure, and an iron catalyst; ammonia is removed by cooling and unreacted gases recycled (chemistry only).
4.10.4.2	Explain the choice of Haber conditions in terms of rate, equilibrium yield and cost (HT) (chemistry only).
4.10.4.3	NPK fertilisers provide nitrogen, phosphorus and potassium for plant growth (chemistry only).
4.10.4.3	Industrial production of NPK uses ammonia, acids and mined minerals; compare lab and industrial preparation of salts (chemistry only).