GCSE Combined Science: Chemistry

Curriculum Overview, Years 9–11

AQA GCSE Combined Science: Trilogy (8464) · Teaching sequence from the Curriculum Road Map (Chemistry: Combined Science)

Course at a glance

This overview sets out the chemistry content of AQA GCSE Combined Science: Trilogy (8464) in the order taught on the Combined Science route, with the specification content for each unit. Unit codes follow the textbook chapters; C10, C11 and C15 are Triple only and are not part of this course. Specification references are to the AQA 8464 specification.

Assessment

Paper	Specification content	Road map units	Format
Chemistry Paper 1	5.1 Atomic structure and the periodic table 5.2 Bonding, structure and the properties of matter 5.3 Quantitative chemistry 5.4 Chemical changes 5.5 Energy changes	C1, C2, C3, C4, C5 (Parts 1 and 2), C6, C7	Written exam, 1 h 15 min, 70 marks, 16.7% of the qualification
Chemistry Paper 2	5.6 The rate and extent of chemical change 5.7 Organic chemistry 5.8 Chemical analysis 5.9 Chemistry of the atmosphere 5.10 Using resources *Questions may also draw on fundamental concepts from 5.1–5.3.*	C8, C9, C12, C13, C14	Written exam, 1 h 15 min, 70 marks, 16.7% of the qualification

- Combined Science: Trilogy is a double award worth two GCSEs, assessed by six papers (two each in biology, chemistry and physics). Grades are on a 17-point scale from 1–1 to 9–9: Foundation tier 1–1 to 5–5, Higher tier 4–4 to 9–9.
- Papers include multiple choice, structured, closed short answer and open response questions.
- At least 15% of marks draw on the 21 required practicals; RP8–RP13 are chemistry. Across the qualification at least 20% of marks assess mathematical skills (weighted 10% biology, 20% chemistry, 30% physics).
- Assessment objectives: AO1 knowledge and understanding 40%, AO2 application 40%, AO3 analysis, interpretation and evaluation 20%.

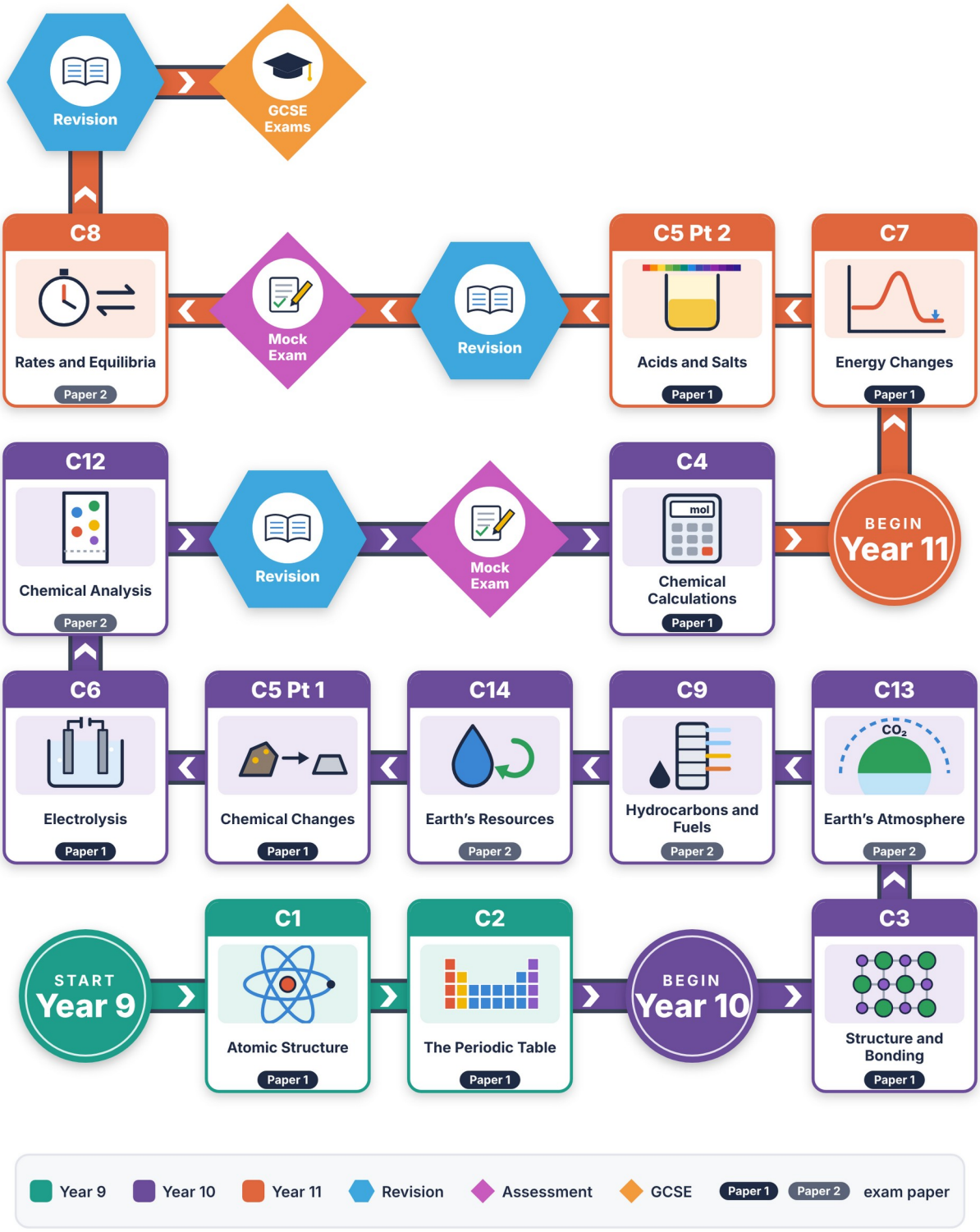
Key to labels

HT only Assessed on the Higher tier paper only. A label beside a topic heading applies to the whole subsection; a label at the start of a bullet applies to that point only.

Teaching sequence

Chemistry: Combined Science Road Map

AQA GCSE Combined Science: Trilogy (8464) · Years 9–11 · start at the bottom and follow the road up



#	Unit	AQA spec ref	Paper	Required practicals
Year 9				
1	**C1 Atomic Structure**	5.1.1	1	–
2	**C2 The Periodic Table**	5.1.2	1	–
Year 10				
3	**C3 Structure and Bonding**	5.2	1	–
4	**C13 Earth's Atmosphere**	5.9	2	–
5	**C9 Hydrocarbons and Fuels**	5.7.1	2	–
6	**C14 Earth's Resources**	5.10	2	RP13
7	**C5 Part 1 Chemical Changes (Reactivity of Metals)**	5.4.1	1	–
8	**C6 Electrolysis**	5.4.3	1	RP9
9	**C12 Chemical Analysis**	5.8	2	RP12
Revision → Year 10 mock exam				
10	**C4 Chemical Calculations**	5.3	1	–
Year 11				
11	**C7 Energy Changes**	5.5	1	RP10
12	**C5 Part 2 Acids and Salts**	5.4.2	1	RP8
Revision → Year 11 mock exam				
13	**C8 Rates and Equilibria**	5.6	2	RP11
Revision → GCSE exams (Chemistry Paper 1 and Paper 2)				

Year 9

Year 9 introduces the atomic model and the periodic table that the rest of the course builds on, with working scientifically skills developed throughout.

C1 Atomic Structure

Specification: 5.1.1 **Exam:** Chemistry Paper 1 **Required practicals:** None

Establishes the particle and atomic model that every later unit depends on: what atoms, elements, compounds and mixtures are, how the model of the atom developed, and how electrons are arranged. Leads directly into C2.

Spec ref and topic	Content to be covered
5.1.1.1 **Atoms, elements and compounds**	• All substances are made of atoms; an element contains one type of atom and is shown by a symbol (about 100 elements in the periodic table). • Compounds contain two or more elements chemically combined in fixed proportions; they can only be separated by chemical reactions. • Recall names and symbols of the first 20 elements and of Groups 1 and 7; name compounds from formulae. • Write word equations, formulae and balanced symbol equations for reactions in the specification. • **HT only** Write balanced half equations and ionic equations where appropriate.
5.1.1.2 **Mixtures**	• A mixture is two or more substances not chemically combined; each keeps its own properties. • Separation by physical processes: filtration, crystallisation, simple distillation, fractional distillation and chromatography. • Describe, explain and suggest suitable separation and purification techniques.
5.1.1.3 **Development of the model of the atom**	• From indivisible spheres, to the discovery of the electron and the plum pudding model. • Alpha particle scattering experiment leading to the nuclear model; Bohr's electrons in fixed shells; discovery of the proton; Chadwick's evidence for the neutron. • Explain why new experimental evidence leads to a scientific model being changed or replaced (shared content with Physics).
5.1.1.4 **Relative electrical charges of subatomic particles**	• Relative charges: proton +1, neutron 0, electron −1; atoms have equal numbers of protons and electrons so are neutral. • Atomic number = number of protons; all atoms of an element have the same number of protons.
5.1.1.5 **Size and mass of atoms**	• Atoms have a radius of about 0.1 nm (1×10^{-10} m); the nucleus is less than 1/10 000 of the atom's radius and holds almost all its mass. • Relative masses: proton 1, neutron 1, electron very small. Mass number = protons + neutrons. • Isotopes: atoms of the same element with different numbers of neutrons. • Calculate numbers of protons, neutrons and electrons from atomic and mass numbers; use standard form and orders of magnitude.
5.1.1.6 **Relative atomic mass**	• Relative atomic mass is a weighted average of the masses of an element's isotopes, taking account of their abundance. • Calculate relative atomic mass from isotope masses and percentage abundances.
5.1.1.7 **Electronic structure**	• Electrons occupy the lowest available energy levels (shells). • Represent electronic structures of the first 20 elements as diagrams and numbers (e.g. sodium 2,8,1).

C2 The Periodic Table

Specification: 5.1.2 **Exam:** Chemistry Paper 1 **Required practicals:** None

Uses electronic structure from C1 to explain the arrangement of the periodic table and the trends within groups.

Spec ref and topic	Content to be covered
5.1.2.1 **The periodic table**	• Elements arranged in order of atomic (proton) number; elements with similar properties fall in columns (groups). • Elements in the same group have the same number of outer-shell electrons, which gives them similar chemical properties. • Explain how an element's position relates to its atomic structure and use this to predict reactions.
5.1.2.2 **Development of the periodic table**	• Early tables ordered elements by atomic weight; they were incomplete and some elements were placed in the wrong groups. • Mendeleev left gaps for undiscovered elements and switched the order of some pairs; his predictions were confirmed by later discoveries. • Knowledge of isotopes explained why ordering strictly by atomic weight was not always correct.
5.1.2.3 **Metals and non-metals**	• Metals form positive ions; most elements are metals. Metals lie to the left and bottom of the table, non-metals to the right and top. • Explain the differences between metals and non-metals in terms of their atomic structure.
5.1.2.4 **Group 0**	• Noble gases are unreactive because they have stable arrangements of electrons (eight outer electrons; helium two). • Boiling points increase with relative atomic mass going down the group.
5.1.2.5 **Group 1**	• Alkali metals have one outer electron and react readily. • Describe the reactions of lithium, sodium and potassium with oxygen, chlorine and water. • Reactivity increases down the group; explain properties in terms of outer-shell electrons and predict properties from trends.
5.1.2.6 **Group 7**	• Halogens have seven outer electrons and exist as diatomic molecules. • React with metals to form ionic compounds (halide ions, 1− charge) and with non-metals to form covalent molecular compounds. • Down the group: relative molecular mass, melting point and boiling point increase; reactivity decreases. • A more reactive halogen displaces a less reactive one from an aqueous solution of its salt.

Year 10

Year 10 opens with structure and bonding, completing the building blocks started in Year 9. It then turns to chemistry in the natural world: the atmosphere, hydrocarbons and the Earth's resources, before moving to the extraction of metals, electrolysis and chemical analysis. Revision and the Year 10 mock follow, then chemical calculations, which carries on into Year 11.

C3 Structure and Bonding

Specification: 5.2 **Exam:** Chemistry Paper 1 **Required practicals:** None

Opens Year 10 by building on the atomic structure and electron arrangements from C1 and C2: the three types of strong chemical bond and how structure determines the properties of substances, including the forms of carbon.

Spec ref and topic	Content to be covered
5.2.1.1 **Chemical bonds**	• Three types of strong bond: ionic (metal and non-metal; electrons transferred), covalent (non-metals; electron pairs shared) and metallic (delocalised electrons).
5.2.1.2 **Ionic bonding**	• Metal atoms lose electrons to form positive ions; non-metal atoms gain electrons to form negative ions, giving noble gas electronic structures. • Draw dot and cross diagrams for ionic compounds (e.g. NaCl); work out ion charges for Groups 1, 2, 6 and 7 from group number.
5.2.1.3 **Ionic compounds**	• Giant lattice of ions held by strong electrostatic forces acting in all directions (e.g. sodium chloride). • Represent with dot and cross, ball and stick, 2D and 3D diagrams; state the limitations of each model. • Work out the empirical formula of an ionic compound from a model or diagram.
5.2.1.4 **Covalent bonding**	• Atoms share pairs of electrons, forming strong bonds in small molecules, large molecules (polymers) and giant covalent structures (e.g. diamond, silicon dioxide). • Dot and cross diagrams for H_2, Cl_2, O_2, N_2, HCl, H_2O, NH_3 and CH_4; deduce molecular formulae from models and diagrams.
5.2.1.5 **Metallic bonding**	• Giant structure of positive ions in a sea of delocalised electrons; strong metallic bonds.
5.2.2.1–5.2.2.2 **States of matter and state symbols**	• Particle model of solids, liquids and gases; melting/freezing and boiling/condensing points depend on the strength of forces between particles. • Predict states of substances from melting and boiling point data; use state symbols (s), (l), (g) and (aq). • **HT only** Limitations of the simple particle model (no forces shown, particles drawn as solid spheres).
5.2.2.3–5.2.2.6 **Properties of ionic compounds, small molecules, polymers and giant covalent structures**	• Ionic compounds: high melting and boiling points; conduct electricity only when molten or dissolved (ions free to move). • Small molecules: low melting and boiling points because only weak intermolecular forces are overcome; these forces increase with molecular size; do not conduct. • Polymers: very large molecules with strong covalent bonds; relatively strong intermolecular forces make them solids at room temperature. • Giant covalent structures (diamond, graphite, silicon dioxide): very high melting points because many strong covalent bonds must be broken.
5.2.2.7–5.2.2.8 **Metals, alloys and conduction**	• Metals have high melting points; layers of atoms can slide, so pure metals are malleable but often too soft. • Alloys are harder because atoms of different sizes distort the layers. • Metals conduct electricity and heat because delocalised electrons carry charge and thermal energy.

Spec ref and topic	Content to be covered
5.2.3.1–5.2.3.3 **Structure and bonding of carbon**	• Diamond: each carbon forms four covalent bonds; very hard, very high melting point, does not conduct. • Graphite: layers of hexagonal rings, each carbon bonded to three others; weak forces between layers make it soft and slippery; one delocalised electron per atom so it conducts. • Graphene: a single layer of graphite, useful in electronics and composites. • Fullerenes: hollow carbon molecules (e.g. buckminsterfullerene C_{60}); carbon nanotubes have high length-to-diameter ratios, high tensile strength and conductivity. Uses include nanotechnology, electronics, materials, lubricants, catalysts and drug delivery.

C13 Earth's Atmosphere

Specification: 5.9 **Exam:** Chemistry Paper 2 **Required practicals:** None

Moves Year 10 into chemistry in the natural world: how the atmosphere formed and changed, the greenhouse effect and climate change, and pollutants from burning fuels (which links into C9).

Spec ref and topic	Content to be covered
5.9.1.1 **Proportions of gases in the atmosphere**	• For about 200 million years: about four-fifths (~80%) nitrogen, one-fifth (~20%) oxygen, with small proportions of carbon dioxide, water vapour and noble gases.
5.9.1.2 **The Earth's early atmosphere**	• One theory: intense volcanic activity in the first billion years released gases; water vapour condensed to form the oceans. • The early atmosphere was mainly carbon dioxide with little or no oxygen (similar to Mars and Venus today); volcanoes also produced nitrogen; possibly small amounts of methane and ammonia. • Carbon dioxide dissolved in the oceans and carbonates precipitated to form sediments. • Interpret evidence and evaluate theories, recognising why evidence is limited.
5.9.1.3 **How oxygen increased**	• Algae (from about 2.7 billion years ago) and later plants produced oxygen by photosynthesis: $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$. • Oxygen built up over the next billion years to levels that allowed animals to evolve.
5.9.1.4 **How carbon dioxide decreased**	• Removed by photosynthesis and locked up in sedimentary rocks (e.g. limestone from shells and skeletons) and fossil fuels (coal, crude oil, natural gas). • Describe the main changes in the atmosphere over time and their likely causes.
5.9.2.1 **Greenhouse gases**	• Water vapour, carbon dioxide and methane keep the Earth warm enough for life. • The greenhouse effect: short-wavelength radiation from the Sun passes through; longer-wavelength (infrared) radiation emitted by the Earth is absorbed by greenhouse gases.
5.9.2.2 **Human activities and greenhouse gases**	• Carbon dioxide from burning fossil fuels and deforestation; methane from animal farming and decomposition of rubbish in landfill. • Evaluate the quality of evidence on climate change, including peer review, uncertainty and simplified models, and why media reporting may be oversimplified or speculative.
5.9.2.3 **Global climate change**	• Increase in average global temperature is a major cause of climate change. • Describe four potential effects, e.g. rising sea levels (flooding and coastal erosion), more frequent or severe storms, changes in rainfall, temperature and water stress, changes to food production and wildlife distribution.
5.9.2.4 **The carbon footprint and its reduction**	• Carbon footprint: total greenhouse gases emitted over the full life cycle of a product, service or event. • Ways to reduce emissions (e.g. renewable energy, energy conservation, carbon capture and storage, carbon taxes, offsetting) and the problems in doing so.

Spec ref and topic	Content to be covered
5.9.3.1 **Atmospheric pollutants from fuels**	• Burning fuels releases carbon dioxide, water vapour, carbon monoxide, sulfur dioxide and oxides of nitrogen, plus solid particles (soot) and unburned hydrocarbons. • Sulfur dioxide comes from sulfur impurities in fuels; oxides of nitrogen form from nitrogen and oxygen at high temperatures in engines. Predict combustion products from fuel composition.
5.9.3.2 **Properties and effects of pollutants**	• Carbon monoxide is toxic, colourless and odourless. • Sulfur dioxide and oxides of nitrogen cause respiratory problems and acid rain; particulates cause global dimming and health problems.

C9 Hydrocarbons and Fuels

Specification: 5.7.1 **Exam:** Chemistry Paper 2 **Required practicals:** None

Crude oil as a finite resource: separating it into useful fractions, how hydrocarbon properties depend on chain length, combustion, and cracking to make alkenes.

Spec ref and topic	Content to be covered
5.7.1.1 **Crude oil, hydrocarbons and alkanes**	• Crude oil is a finite resource formed from ancient biomass (mainly plankton); it is a mixture of many compounds, mostly hydrocarbons. • Most are alkanes, general formula C_nH_{2n+2}; recognise and draw methane, ethane, propane and butane from formulae.
5.7.1.2 **Fractional distillation and petrochemicals**	• Crude oil is separated into fractions with similar boiling points by fractional distillation. • Fractions are processed into fuels (petrol, diesel oil, kerosene, heavy fuel oil, LPG) and feedstock for the petrochemical industry (solvents, lubricants, polymers, detergents). • Carbon's ability to form families of similar compounds explains the vast range of natural and synthetic carbon compounds.
5.7.1.3 **Properties of hydrocarbons**	• Boiling point, viscosity and flammability depend on molecular size, which affects how fractions are used as fuels. • Complete combustion oxidises carbon and hydrogen to carbon dioxide and water, releasing energy; write balanced equations.
5.7.1.4 **Cracking and alkenes**	• Cracking breaks large hydrocarbons into smaller, more useful molecules by catalytic cracking or steam cracking. • Products include alkanes and alkenes; alkenes are more reactive and turn bromine water from orange to colourless. • Cracking meets demand for fuels with small molecules; alkenes are used to make polymers and as starting materials. Write balanced equations for cracking.

C14 Earth's Resources

Specification: 5.10 **Exam:** Chemistry Paper 2 **Required practicals:** RP13 Water purification

Sustainable use of resources: finite and renewable resources, making water safe to drink, treating waste water, biological extraction of metals, life cycle assessment and recycling. The metal-extraction content leads into C5 Part 1.

Spec ref and topic	Content to be covered
5.10.1.1 **Using the Earth's resources and sustainable development**	• Resources from the Earth, sea and atmosphere provide warmth, shelter, food and transport; natural products are supplemented or replaced by agricultural and synthetic ones. • Sustainable development meets present needs without compromising future generations; distinguish finite from renewable resources. • Extract and interpret information about resources; use orders of magnitude.

Spec ref and topic	Content to be covered
5.10.1.2 **Potable water**	• Potable water is safe to drink (low levels of dissolved salts and microbes) but not chemically pure. • UK fresh water: choose a source, pass through filter beds, then sterilise (chlorine, ozone or UV light). • Where fresh water is scarce: desalination by distillation or reverse osmosis, both needing large amounts of energy. • **Required practical 13:** analysis and purification of water samples (pH, dissolved solids, distillation).
5.10.1.3 **Waste water treatment**	• Sewage, agricultural and industrial waste water must be treated to remove organic matter, harmful microbes and harmful chemicals. • Stages: screening and grit removal, sedimentation, anaerobic digestion of sludge, aerobic biological treatment of effluent. • Compare the ease of obtaining potable water from waste water, ground water and salt water.
5.10.1.4 **Alternative methods of extracting metals** HT only	• Copper-rich ores are limited; low-grade ores can be processed by phytomining (plants absorb metal compounds and are burned to ash) and bioleaching (bacteria produce leachate solutions). • Copper is then obtained by displacement with scrap iron or by electrolysis. Evaluate these biological methods.
5.10.2.1 **Life cycle assessment**	• LCA assesses environmental impact at each stage: raw materials, manufacture and packaging, use, and disposal. • Some impacts are easily quantified (water, energy, waste); pollutant effects involve value judgements; selective LCAs can be misused. • Carry out simple comparative LCAs, e.g. plastic versus paper shopping bags.
5.10.2.2 **Ways of reducing the use of resources**	• Reducing use, reusing and recycling conserve limited resources and energy and cut waste and environmental impact. • Glass bottles reused or melted down; metals recycled by melting and recasting (e.g. scrap steel added to iron from the blast furnace). • Evaluate ways of reducing the use of limited resources, given information.

C5 Part 1 Chemical Changes (Reactivity of Metals)

Specification: 5.4.1 **Exam:** Chemistry Paper 1 **Required practicals:** None

Reactivity of metals and how it determines the method of extraction from ores; introduces oxidation and reduction. Leads directly into C6 for metals more reactive than carbon.

Spec ref and topic	Content to be covered
5.4.1.1 **Metal oxides**	• Metals react with oxygen to form metal oxides (oxidation = gain of oxygen; reduction = loss of oxygen). • Identify which substances are oxidised or reduced in terms of gain or loss of oxygen.
5.4.1.2 **The reactivity series**	• Reactivity relates to a metal's tendency to form positive ions. Order K, Na, Li, Ca, Mg, Zn, Fe, Cu using reactions with water and dilute acids; hydrogen and carbon are often included. • A more reactive metal displaces a less reactive metal from its compounds. • Deduce an order of reactivity from experimental results.
5.4.1.3 **Extraction of metals and reduction**	• Unreactive metals such as gold are found native; most metals are found as compounds and must be extracted chemically. • Metals less reactive than carbon are extracted from their oxides by reduction with carbon. • Interpret and evaluate given extraction processes; identify substances oxidised and reduced.

Spec ref and topic	Content to be covered
5.4.1.4 **Oxidation and reduction in terms of electrons** HT only	• Oxidation is loss of electrons; reduction is gain of electrons. • Write ionic equations for displacement reactions and identify which species are oxidised and reduced.

C6 Electrolysis

Specification: 5.4.3 **Exam:** Chemistry Paper 1 **Required practicals:** RP9 Electrolysis

Decomposing ionic compounds using electricity, both molten and in solution, and its industrial use to extract aluminium. Builds on ionic bonding (C3) and reactivity (C5 Part 1).

Spec ref and topic	Content to be covered
5.4.3.1 **The process of electrolysis**	• Molten or dissolved ionic compounds conduct because their ions are free to move (electrolytes). • Positive ions move to the cathode (negative electrode) and negative ions to the anode (positive electrode), where they are discharged as elements. • HT only Write half equations for the reactions at the electrodes.
5.4.3.2 **Electrolysis of molten ionic compounds**	• Molten binary compounds break down into elements: metal at the cathode, non-metal at the anode (e.g. lead bromide gives lead and bromine). • Predict products of electrolysing binary ionic compounds in the molten state.
5.4.3.3 **Using electrolysis to extract metals**	• Metals more reactive than carbon are extracted by electrolysis of molten compounds, which uses large amounts of energy. • Aluminium: electrolysis of aluminium oxide dissolved in molten cryolite (lowers the melting point); the carbon anodes burn away as carbon dioxide and must be replaced.
5.4.3.4 **Electrolysis of aqueous solutions**	• At the cathode, hydrogen is produced if the metal is more reactive than hydrogen; at the anode, oxygen is produced unless halide ions are present (then the halogen forms). • This is because water provides H^+ and OH^- ions. Predict the products of electrolysing aqueous solutions. • **Required practical 9:** investigate electrolysis of aqueous solutions with inert electrodes (developing a hypothesis).
5.4.3.5 **Half equations at the electrodes** HT only	• Reduction (gain of electrons) occurs at the cathode; oxidation (loss of electrons) at the anode. • Write and balance half equations, e.g. $2H^+ + 2e^- \rightarrow H_2$ and $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$.

C12 Chemical Analysis

Specification: 5.8 **Exam:** Chemistry Paper 2 **Required practicals:** RP12 Chromatography

How chemists establish what a substance is and whether it is pure: formulations, chromatography and tests for common gases.

Spec ref and topic	Content to be covered
5.8.1.1 **Pure substances**	• In chemistry, a pure substance is a single element or compound; in everyday language ‘pure’ can mean nothing added. • Pure substances melt and boil at specific temperatures; use melting and boiling point data to distinguish pure from impure substances.
5.8.1.2 **Formulations**	• A formulation is a mixture designed as a useful product, with components in measured quantities (e.g. fuels, cleaning agents, paints, medicines, alloys, fertilisers, foods). • Identify formulations given appropriate information.

Spec ref and topic	Content to be covered
5.8.1.3 **Chromatography**	• Separation depends on the distribution of substances between a stationary phase and a mobile phase. • R_f = distance moved by substance ÷ distance moved by solvent; a pure compound gives a single spot in all solvents. • Interpret chromatograms and calculate R_f values. • **Required practical 12:** use paper chromatography to separate and distinguish coloured substances and calculate R_f values.
5.8.2.1–5.8.2.4 **Identification of common gases**	• Hydrogen: burning splint gives a squeaky pop. Oxygen: relights a glowing splint. • Carbon dioxide: turns limewater milky (cloudy). Chlorine: bleaches damp litmus paper white.

Revision → Year 10 mock exam. Revision and mock exam on the units taught so far.

C4 Chemical Calculations

Specification: 5.3 **Exam:** Chemistry Paper 1 **Required practicals:** None

Quantitative chemistry and the mathematical skills used throughout the course: conservation of mass, relative formula mass, moles (Higher tier), concentration and uncertainty.

Placement: on the road map C4 follows the Year 10 mock and sits directly before the Year 11 marker, so it bridges the end of Year 10 and the start of Year 11.

Spec ref and topic	Content to be covered
5.3.1.1 **Conservation of mass and balanced equations**	• No atoms are lost or made in a reaction, so the mass of products equals the mass of reactants. • Balanced symbol equations have the same number of atoms of each element on both sides.
5.3.1.2 **Relative formula mass**	• M_r is the sum of the relative atomic masses of the atoms in the formula; in a balanced equation the total M_r of reactants equals that of products. • Calculate the percentage by mass of an element in a compound.
5.3.1.3 **Mass changes when a reactant or product is a gas**	• Apparent mass changes in open systems are due to gases, e.g. a metal gains mass when it forms an oxide; a carbonate loses mass when CO_2 escapes on thermal decomposition. • Explain observed mass changes using the particle model.
5.3.1.4 **Chemical measurements**	• Every measurement has some uncertainty; use the range of results about the mean as a measure of uncertainty.
5.3.2.1–5.3.2.4 **Moles, reacting masses and limiting reactants** **HT only**	• The mole: mass of one mole in grams equals M_r; one mole contains 6.02×10^{23} particles (Avogadro constant). Convert between mass and moles. • Calculate masses of reactants and products from balanced equations; use masses converted to moles to find balancing numbers. • Limiting reactants: the reactant used up limits the amount of product; one reactant is often used in excess.
5.3.2.5 **Concentration of solutions**	• Concentration as mass of solute per volume of solution (g/dm^3); calculate the mass of solute in a given volume. • **HT only** Explain how mass of solute and volume of solution relate to concentration.

Year 11

Year 11 applies the quantitative skills from C4 to energy changes and to acids and salts, followed by revision and the Year 11 mock. Pupils then study rates of reaction and equilibrium before revision for the GCSE exams.

C7 Energy Changes

Specification: 5.5 **Exam:** Chemistry Paper 1 **Required practicals:** RP10 Temperature changes

Energy transfers in reactions and reaction profiles. Higher tier adds bond energy calculations.

Spec ref and topic	Content to be covered
5.5.1.1 **Exothermic and endothermic reactions**	• Energy is conserved in reactions. Exothermic reactions transfer energy to the surroundings (temperature rises), e.g. combustion, many oxidations, neutralisation; uses include self-heating cans and hand warmers. • Endothermic reactions take in energy (temperature falls), e.g. thermal decomposition, citric acid with sodium hydrogencarbonate; uses include some sports injury packs. • Evaluate uses and applications of exothermic and endothermic reactions. • **Required practical 10:** investigate variables that affect temperature changes in reacting solutions (e.g. acid with metals or carbonates, neutralisation, displacement).
5.5.1.2 **Reaction profiles**	• Reactions occur only when particles collide with enough energy; the minimum energy needed is the activation energy. • Draw and interpret reaction profiles showing reactant and product energies, activation energy and overall energy change.
5.5.1.3 **The energy change of reactions** HT only	• Breaking bonds takes in energy; making bonds releases energy. The overall change is the difference between the two. • Calculate energy transferred using given bond energies; explain exothermic and endothermic in terms of bonds.

C5 Part 2 Acids and Salts

Specification: 5.4.2 **Exam:** Chemistry Paper 1 **Required practicals:** RP8 Making salts

Reactions of acids, making salts, the pH scale and neutralisation, building on the calculation skills from C4. Higher tier adds strong and weak acids.

Spec ref and topic	Content to be covered
5.4.2.1 **Reactions of acids with metals**	• Acids react with some metals to produce salts and hydrogen (limited to magnesium, zinc and iron with hydrochloric and sulfuric acids). • HT only Explain these as redox reactions in terms of electron transfer and identify what is oxidised and reduced.
5.4.2.2 **Neutralisation of acids and salt production**	• Acids are neutralised by alkalis (soluble metal hydroxides) and bases (insoluble hydroxides and metal oxides) to give a salt and water, and by metal carbonates to give a salt, water and carbon dioxide. • Hydrochloric acid gives chlorides, nitric acid nitrates, sulfuric acid sulfates. Predict products and deduce formulae of salts from ion charges.
5.4.2.3 **Soluble salts**	• Make soluble salts by adding an insoluble solid (metal, oxide, hydroxide or carbonate) to an acid until in excess, filtering, then crystallising. • Describe how to make a pure, dry sample of a named soluble salt. • **Required practical 8:** prepare a pure, dry sample of a soluble salt from an insoluble oxide or carbonate.
5.4.2.4 **The pH scale and neutralisation**	• Acids produce H^+ ions in solution; alkalis produce OH^- ions. pH 0–14 measured with universal indicator or a pH probe; pH 7 is neutral. • In neutralisation, $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$.

Spec ref and topic	Content to be covered
5.4.2.5 **Strong and weak acids** HT only	• Strong acids (hydrochloric, nitric, sulfuric) are completely ionised in solution; weak acids (ethanoic, citric, carbonic) are only partially ionised. • At the same concentration, a stronger acid has a lower pH; a decrease of one pH unit means the H^+ concentration increases by a factor of 10. • Use and distinguish the terms dilute/concentrated (amount of substance) and weak/strong (degree of ionisation).

Revision → Year 11 mock exam. Revision and mock exam on the units taught so far.

C8 Rates and Equilibria

Specification: 5.6 **Exam:** Chemistry Paper 2 **Required practicals:** RP11 Rates of reaction

How fast reactions go and how far they go: measuring and calculating rates, collision theory, catalysts, reversible reactions and dynamic equilibrium. Higher tier applies Le Chatelier's principle.

Spec ref and topic	Content to be covered
5.6.1.1 **Calculating rates of reactions**	• Mean rate = quantity of reactant used (or product formed) ÷ time taken, in g/s or cm^3/s. • Draw and interpret graphs of quantity against time; use the slope of a tangent as a measure of rate. • HT only Rates in mol/s; calculate the gradient of a tangent to find the rate at a specific time.
5.6.1.2 **Factors affecting rate**	• Concentration of solutions, pressure of gases, surface area of solids, temperature and catalysts. • **Required practical 11:** investigate how concentration affects rate by measuring the volume of gas produced and by a change in colour or turbidity.
5.6.1.3 **Collision theory and activation energy**	• Reactions need collisions with at least the activation energy. • Higher concentration, pressure and surface area increase collision frequency; higher temperature increases both frequency and energy of collisions. • Predict and explain effects of changing conditions, including surface area to volume ratio and simple proportionality.
5.6.1.4 **Catalysts**	• Catalysts speed up reactions without being used up, by providing a pathway with lower activation energy; enzymes are biological catalysts. • Draw reaction profiles for catalysed reactions; identify catalysts from their effect on rate and absence from the equation.
5.6.2.1–5.6.2.3 **Reversible reactions and equilibrium**	• Some reactions are reversible (A + B ⇌ C + D); direction can be changed by changing conditions, e.g. heating ammonium chloride. • If one direction is exothermic the other is endothermic by the same amount, e.g. hydrated/anhydrous copper(II) sulfate. • In a closed system equilibrium is reached when forward and reverse reactions occur at exactly the same rate.
5.6.2.4–5.6.2.7 **Changing conditions at equilibrium (Le Chatelier)** HT only	• If conditions change, the system responds to counteract the change (Le Chatelier's principle); make qualitative predictions. • Concentration: adding reactant or removing product shifts equilibrium towards products. • Temperature: increasing it favours the endothermic direction; decreasing it favours the exothermic direction. • Pressure (gases): increasing it shifts equilibrium towards the side with fewer molecules.

Revision → GCSE exams. Chemistry Paper 1 covers C1–C7 (specification 5.1–5.5); Chemistry Paper 2 covers C8, C9, C12, C13 and C14 (specification 5.6–5.10).

Required practical activities

Combined Science: Trilogy has 21 required practicals; RP8–RP13 are chemistry. Questions drawing on them make up at least 15% of the exam marks.

RP	Activity	Spec ref	Taught in
8	Preparation of a pure, dry sample of a soluble salt from an insoluble oxide or carbonate	5.4.2.3	C5 Part 2 Acids and Salts (Y11)
9	Investigating electrolysis of aqueous solutions using inert electrodes	5.4.3.4	C6 Electrolysis (Y10)
10	Investigating variables that affect temperature changes in reacting solutions	5.5.1.1	C7 Energy Changes (Y11)
11	Investigating how concentration affects rate (volume of gas produced; change in colour or turbidity)	5.6.1.2	C8 Rates and Equilibria (Y11)
12	Using paper chromatography to separate coloured substances and calculate R_f values	5.8.1.3	C12 Chemical Analysis (Y10)
13	Analysis and purification of water samples (pH, dissolved solids, distillation)	5.10.1.2	C14 Earth's Resources (Y10)

Specification coverage map

Every chemistry subsection of the AQA 8464 subject content, with the unit that teaches it. All 21 subsections are covered.

Spec	Title	Unit	Year	Paper
5.1.1	A simple model of the atom, symbols, relative atomic mass, electronic charge and isotopes	C1	9	1
5.1.2	The periodic table	C2	9	1
5.2.1	Chemical bonds, ionic, covalent and metallic	C3	10	1
5.2.2	How bonding and structure are related to the properties of substances	C3	10	1
5.2.3	Structure and bonding of carbon	C3	10	1
5.3.1	Chemical measurements, conservation of mass and the quantitative interpretation of chemical equations	C4	10–11	1
5.3.2	Use of amount of substance in relation to masses of pure substances	C4	10–11	1
5.4.1	Reactivity of metals	C5 Part 1	10	1
5.4.2	Reactions of acids	C5 Part 2	11	1
5.4.3	Electrolysis	C6	10	1
5.5.1	Exothermic and endothermic reactions	C7	11	1
5.6.1	Rate of reaction	C8	11	2
5.6.2	Reversible reactions and dynamic equilibrium	C8	11	2
5.7.1	Carbon compounds as fuels and feedstock	C9	10	2
5.8.1	Purity, formulations and chromatography	C12	10	2
5.8.2	Identification of common gases	C12	10	2
5.9.1	The composition and evolution of the Earth’s atmosphere	C13	10	2
5.9.2	Carbon dioxide and methane as greenhouse gases	C13	10	2
5.9.3	Common atmospheric pollutants and their sources	C13	10	2
5.10.1	Using the Earth’s resources and obtaining potable water	C14	10	2
5.10.2	Life cycle assessment and recycling	C14	10	2

Textbook, revision guides and websites

Textbooks used in class

- AQA GCSE Biology Student Book (AQA GCSE Science 3rd Edition), Ann Fullick; series editor Lawrie Ryan (Oxford University Press)
- AQA GCSE Chemistry Student Book (AQA GCSE Science 3rd Edition), Lawrie Ryan (Oxford University Press)
- AQA GCSE Physics Student Book (AQA GCSE Science 3rd Edition), Jim Breithaupt; series editor Lawrie Ryan (Oxford University Press)

Recommended revision guide

- Collins AQA GCSE 9-1 Combined Science Trilogy Complete Revision & Practice (Higher or Foundation)

Useful websites

- **Physics & Maths Tutor:** GCSE Science past paper questions with mark schemes; some have video solutions. https://www.physicsandmathstutor.com/
- **AQA:** all past papers and mark schemes. https://www.aqa.org.uk/subjects/science/gcse/science-8464/assessment-resources
- **Revisely:** Free Science Lessons revision videos. https://www.revisely.com/gcse/combined-science/aqa

Source: specification content summarised from AQA GCSE Combined Science: Trilogy (8464) specification (aqa.org.uk/8464). Check the AQA website for any later updates.