AQA Chemistry	Working towards expected outcomes Your child is not yet making the expected progress within this course.	Working at expected outcomes Your child is achieving the expected progress for this point within the course.	Working beyond expected outcomes Your child is exceeding the expected progress.
Year 12 Autumn 1	Students working **towards** expected outcomes in Year 12 can:	Students working **at** expected in Year 12 can:	Students working **beyond** expected in Year 12 can:
Atomic Structure & Periodicity	*Students are beginning to grasp foundational ideas but require support to apply them confidently.* • Recall relative charges and masses of subatomic particles. • Describe atomic structure in terms of nucleus and electrons. • Define mass number, atomic number, and isotopes. • Identify numbers of subatomic particles in atoms and simple ions. • Describe the purpose of mass spectrometry. • Write electron configurations up to calcium (Z = 20). • Define and write equations for first ionisation energy.	*Students can apply their knowledge and skills to solve standard problems and interpret familiar data.* • Use atomic and mass numbers to determine subatomic particles in atoms and ions. • Explain isotope existence and calculate relative atomic mass from abundance data. • Interpret simple mass spectra to identify isotopes and relative abundances. • Explain how TOF mass spectrometry identifies elements and molecular masses. • Write full electron configurations for atoms and ions up to Z = 36. • Write equations for successive ionisation energies. • Use ionisation energy trends in Period 3 and Group 2 to explain electron shell and sub-shell structure.	*Students demonstrate a deeper understanding and can apply their knowledge in unfamiliar or more complex contexts.* • Evaluate isotope effects on chemical/physical properties and atomic mass. • Explain how TOF mass spectrometry identifies elements and molecular masses. • Analyse complex mass spectra, including fragmentation and isotopic patterns. • Explain ionisation energy trends in terms of nuclear charge, electron shielding, and sub-shell structure. • Predict electron configuration anomalies using orbital energy levels (e.g., 3d vs 4s). • Critically assess how experimental evidence led to current atomic models.
Bonding	• Define ionic, covalent, dative covalent, and metallic bonding. • Construct formulas for ionic compounds, including those with compound ions.	• Predict charges and write correct formulas for ionic compounds, including those with polyatomic ions • Represent covalent and dative bonds clearly using appropriate symbols.	• Explain physical properties of unfamiliar substances by analysing their bonding and structure. • Analyse molecules with lone and bonding pairs to predict bond angles using electron-pair repulsion principles.

	• Recognise covalent and coordinate bonds using lines and arrows. • Identify basic properties of ionic, covalent, and metallic substances. • Recall basic shapes of molecules with two or more electron pairs. • Define electronegativity and polarity in simple terms. • Name basic intermolecular forces (e.g. van der Waals, hydrogen bonding).	• Describe metallic bonding in terms of delocalised electrons and lattice structure. • Link bonding and structure type to physical properties (melting point, conductivity, solubility). • Distinguish between crystal types: ionic, molecular, metallic, and macromolecular, using examples (e.g., NaCl, iodine, graphite). • Use VSEPR theory to explain shapes and bond angles of molecules and ions with up to six electron pairs. • Use partial charges (δ+ and δ–) to indicate bond polarity and identify when a molecule has a permanent dipole. • Explain how intermolecular forces influence boiling and melting points, including anomalies like ice and hydrogen bonding.	• Predict and explain cases where polar bonds do not result in a permanent dipole (due to molecular symmetry). • Compare strengths of intermolecular forces and link them to observed trends in boiling and melting points across groups or homologous series. • Interpret data or experimental observations to deduce bonding types or structure (e.g., solubility, conductivity, melting point tests). • Justify anomalous properties such as the low density of ice or high boiling point of water using hydrogen bonding theory.
Autumn 2 Amount of Substance	• Define relative atomic mass (Ar) and relative molecular mass (Mr), using the ^{12}C scale. • Define the mole and the Avogadro constant. • Use mol = mass / Mr for simple substances. • Recognise the term "relative formula mass" for ionic compounds. • State what empirical and molecular formulas represent. • Recognise percentage yield and atom economy from formulae. • Use correct units in the ideal gas equation when prompted.	• Calculate moles from mass, Mr, concentration, or gas volume. • Use the Avogadro constant in calculations involving particles and moles. • Rearrange and use the ideal gas equation $pV=nRT$ with SI units. • Determine empirical formula from composition by mass or percentage. • Determine molecular formula using empirical formula and Mr. • Write and balance full and ionic chemical equations. • Calculate mass, volume, concentration, percentage yield, and atom economy from balanced equations.	• Interpret mole-based problems involving unfamiliar contexts or multi-step calculations. • Justify use of different mole relationships depending on chemical species • Evaluate assumptions behind the ideal gas equation and its limitations. • Critically assess ethical, economic, and environmental implications of low vs high atom economy processes. • Construct and manipulate complex equations for quantitative analysis in practical contexts

		• Report calculated values using correct significant figures and account for measurement accuracy.	
Alkanes	• Define alkanes as saturated hydrocarbons. • Describe crude oil as a mixture mainly composed of alkanes. • State that fractional distillation separates hydrocarbons by boiling point. • Identify products of complete and incomplete combustion. • Recall that burning hydrocarbons produces pollutants such as CO_2, CO, NO_x, carbon, and unburned hydrocarbons. • State that alkanes can react with halogens (e.g. chlorine).	• Describe how fractional distillation works and why it separates alkanes based on chain length and boiling point. • Explain why alkanes are cracked to meet demand and improve economic value. • Compare conditions for thermal and catalytic cracking and their main products. • Write balanced equations for complete and incomplete combustion of alkanes. • Identify pollutants from combustion and describe how catalytic converters reduce them. • Explain how sulfur dioxide forms from sulfur-containing fuels and how it can be removed using calcium oxide or carbonate. • Describe the free radical substitution of methane with chlorine using initiation, propagation, and termination steps.	• Evaluate the environmental impact of alkane use as fuels, including greenhouse gas emissions and acid rain formation. • Analyse and compare the efficiency and product distribution of thermal vs catalytic cracking. • Discuss the design and function of catalytic converters in reducing multiple pollutants. • Justify how desulfurisation of flue gases helps prevent acid rain and relate this to industrial processes. • Explain limitations and side reactions in free radical substitution (e.g. multiple substitutions or formation of isomers). • Apply knowledge of bond breaking and formation to assess the energy profile of combustion and cracking processes
Halogenoalkanes	• Identify halogenoalkanes as molecules with polar carbon–halogen bonds. • Recognise OH^-, CN^- and NH_3 as nucleophiles. • State that halogenoalkanes can undergo substitution and elimination reactions. • Describe ozone as beneficial in absorbing UV radiation.	• Outline the nucleophilic substitution mechanisms of halogenoalkanes with OH^-, CN^- and NH_3. • Explain how carbon–halogen bond enthalpy influences reaction rate in hydrolysis. • Outline the elimination mechanism of halogenoalkanes with hot, ethanolic OH^-. • Explain how OH^- acts as both a nucleophile (substitution) and a base (elimination). • Compare conditions favouring substitution versus elimination.	• Analyse hydrolysis experiments to compare reactivity of different halogenoalkanes based on halogen type and structure. • Predict major and minor products of substitution and elimination in asymmetric halogenoalkanes. • Explain competing substitution and elimination pathways using reaction conditions and mechanism diagrams.

	• State that CFCs can break down in the upper atmosphere to produce chlorine radicals.	• Use equations to show how chlorine radicals catalyse ozone decomposition. • Explain how UV radiation causes C–Cl bond homolysis in CFCs. • Describe the impact of CFCs on ozone and the resulting environmental consequences.	• Evaluate the significance of bond enthalpy vs polarity in determining halogenoalkane reactivity. • Use radical equations to describe the catalytic cycle of ozone depletion and show the regeneration of chlorine radicals.
Spring 1 Energetics	• Recognise that reactions can be exothermic or endothermic. • Identify enthalpy change (ΔH) as heat change under constant pressure. • Recognise standard conditions as 100 kPa and a stated temperature. • Use $q = mc\Delta T$ to calculate energy transferred in a reaction. • Use enthalpy change equations with values provided. • Define mean bond enthalpy.	• Define standard enthalpy of combustion ($\Delta cH°$) and formation ($\Delta fH°$). • Use $q = mc\Delta T$ to calculate molar enthalpy change, using appropriate units and significant figures. • Perform related calculations involving energy change, mass, temperature change, or moles. • Calculate enthalpy changes using Hess's Law with combustion or formation data. • Calculate approximate ΔH values using mean bond enthalpies. • Explain why ΔH values from bond enthalpy differ from those using Hess's Law. • Represent energy changes using energy profile diagrams or enthalpy cycles.	• Evaluate the accuracy of experimental ΔH values against data book values, considering heat loss or incomplete combustion. • Construct and interpret Hess cycles using multiple routes with combustion or formation enthalpies. • Justify differences between mean bond enthalpy values and those derived experimentally. • Discuss assumptions and limitations in bond enthalpy calculations.
Alkenes	• Recognise that alkenes are unsaturated hydrocarbons with a double bond. • Identify the C=C bond as a region of high electron density.	• Outline electrophilic addition mechanisms with HBr, H_2SO_4, and Br_2 using curly arrows. • Explain the formation of major and minor products based on carbocation stability. • Draw the repeating unit from a given alkene monomer.	• Analyse carbocation stability (primary, secondary, tertiary) to predict major products in addition reactions of unsymmetrical alkenes. • Justify the physical properties and uses of polymers like PVC based on their structure.

	• Describe the general reactions of alkenes with bromine, hydrogen halides, and sulfuric acid. • Use bromine water to test for unsaturation and describe observations. • Recognise that polymers are formed from alkenes.	• Draw the repeating unit from a polymer segment and deduce the original monomer. • Apply IUPAC rules to name simple addition polymers. • Explain why addition polymers are chemically unreactive. • Describe intermolecular forces between polymer chains and how they influence polymer properties.	• Explain the effect of plasticisers on the flexibility and use of poly(chloroethene). • Compare properties of addition polymers and relate them to their molecular structure and bonding. • Discuss how polymer chemistry and applications have evolved over time with scientific advances.
Alcohols	• Define the term biofuel. • Recognise that ethanol can be produced by fermentation and fractional distillation. • Describe the hydration of alkenes using steam and an acid catalyst. • Identify primary, secondary, and tertiary alcohols. • Describe the action of acidified potassium dichromate(VI) as an oxidising agent. • Recognise that alcohols can be dehydrated to form alkenes.	• Justify the conditions used in the fermentation of glucose to produce ethanol. • Write balanced equations for the production of ethanol by hydration and fermentation. • Outline the mechanism for hydration of alkenes to form alcohols using electrophilic addition. • Outline the mechanism for the acid-catalysed elimination of water from alcohols to form alkenes. • Write oxidation equations for primary alcohols to aldehydes and carboxylic acids, and for secondary alcohols to ketones. • Explain how the method of oxidation affects whether an aldehyde or carboxylic acid is produced. • Use Fehling's solution and Tollens' reagent to distinguish between aldehydes and ketones. • Describe fractional distillation as a method for purifying ethanol. • Explain that tertiary alcohols are not easily oxidised.	• Evaluate whether ethanol produced by fermentation is truly carbon neutral using chemical equations and reasoning. • Discuss ethical and environmental issues around the use of biofuels, including land use, carbon emissions, and food supply. • Compare the sustainability of polymer production using alcohol-derived alkenes versus crude oil. • Interpret experimental results from oxidation or elimination reactions involving alcohols. • Analyse competing pathways for alcohol reactions and justify conditions and reagents for specific outcomes (e.g. aldehyde vs acid formation).

Spring 2 Kinetics	• Define activation energy. • Explain why most particle collisions do not cause a reaction. • Describe the Maxwell–Boltzmann distribution of molecular energies. • Explain the meaning of rate of reaction. • Recognise that increasing temperature increases reaction rate. • Describe how concentration and pressure changes affect collision frequency. • Define a catalyst and its basic role in a chemical reaction.	• Draw and interpret Maxwell–Boltzmann distribution curves at different temperatures. • Use the Maxwell–Boltzmann distribution to explain why a small increase in temperature causes a large increase in reaction rate. • Explain how changes in concentration or pressure influence reaction rate by affecting collision frequency. • Describe how catalysts provide an alternative reaction pathway with lower activation energy. • Use Maxwell–Boltzmann distribution to explain how catalysts increase reaction rates for gases. • Carry out and analyse experiments investigating the effect of temperature & concentration on reaction rate (e.g. sodium thiosulfate and hydrochloric acid).	• Investigate and explain how understanding factors affecting reaction rate has influenced food storage and cooking methods. • Analyse Maxwell–Boltzmann distribution changes quantitatively and predict effects on rate constants. • Evaluate the practical applications of catalysts in industry and everyday life. • Design experiments to investigate the effect of pressure on reaction rates and interpret data quantitatively.
Chemical Equilibria	• Recognise that many reactions are reversible. • Describe the features of a system at equilibrium (equal forward and reverse rates, constant concentrations). • State that a catalyst does not change the position of equilibrium. • Identify when a reaction is in a homogeneous system. • Write basic Kc expressions from balanced equations.	• Use Le Chatelier's principle to predict how changes in temperature, pressure, and concentration affect the position of equilibrium in homogeneous reactions. • Explain why a compromise temperature and pressure may be used in industrial reversible reactions. • Construct correct Kc expressions from balanced equations for homogeneous systems. • Calculate Kc values from equilibrium concentrations. • Perform calculations involving changes in concentration to determine Kc or unknown concentrations at equilibrium.	• Analyse the extent of equilibrium shift in response to changing conditions using particle and energy-level reasoning. • Explore the industrial optimisation of reactions (e.g. Haber or Contact process) with detailed reasoning behind selected pressure and temperature conditions. • Justify which experimental results would cause the largest uncertainty in calculated Kc values.

		• Predict qualitatively how temperature changes affect the value of Kc based on whether a reaction is exothermic or endothermic. • Understand that Kc is not affected by changes in concentration or the presence of a catalyst.	
Organic Analysis	• Recall that different organic compounds give different results in test-tube reactions. • Recognise visible outcomes of common tests. • Understand that mass spectrometry gives molecular mass from the molecular ion peak. • Recognise that infrared radiation is absorbed by specific bonds in molecules. • Identify simple functional groups using data sheets and IR spectra with guidance. • Be aware that gases like CO_2, CH_4, and H_2O absorb infrared and contribute to global warming.	• Use test-tube reactions to distinguish between alcohols, aldehydes, alkenes, and carboxylic acids. • Use mass spectra to determine the molecular formula using accurate relative atomic masses. • Identify molecular ion peaks and simple fragments. • Interpret IR spectra using a data sheet to: Identify key bonds (O–H, C=O, C–H) Detect impurities Understand the use of the fingerprint region • Explain the role of greenhouse gases in global warming due to IR absorption.	• Explain the chemical basis of test-tube reactions (e.g. redox for aldehydes). • Design a logical sequence of tests to identify multiple functional groups. • Interpret fragmentation patterns in mass spectra to deduce parts of molecular structure. • Combine MS and IR evidence to propose possible structures. • Analyse complex IR spectra, distinguishing overlapping peaks. • Evaluate limitations of analytical techniques and suggest confirmatory methods.
Redox	• Identify oxidation as loss of electrons, reduction as gain. • Recognise oxidising and reducing agents. • Assign basic oxidation states (e.g. O = –2, H = +1)	• Assign oxidation states to elements in compounds/ions. • Identify species oxidised/reduced in redox reactions. • Write and combine half-equations, balancing atoms and charges.	• Balance complex redox equations in acidic/basic conditions. • Analyse unfamiliar redox processes using oxidation states and electron transfer. • Explain redox trends and disproportionation reactions.

Summer 1 Group 2&7	• Identify Group 2 and Group 7 elements and recognise key trends. • Recall some basic reactions. • Group 2 with water. • Halogens with halide ions • Recognise precipitates formed in halide tests. • Know that chlorine is used in water treatment.	• Explain trends down Group 2 (atomic radius, ionisation energy, melting point) and Group 7 (electronegativity, boiling point, oxidising ability). • Describe and explain: Reactions of Mg–Ba with water and steam, Solubility trends of Group 2 hydroxides and sulfates, Uses of Group 2 compounds (e.g. $BaSO_4$ in medicine, $Ca(OH)_2$ in agriculture). • Clearly show their knowledge and understanding regarding displacement reactions between halogens and halides • Reactions of halide salts with H_2SO_4	• Justify trends using structure, bonding, and charge density (e.g. explain why $BaSO_4$ is insoluble). • Explain redox changes in Group 2 and Group 7 reactions using half-equations. • Evaluate industrial and medical uses (e.g. magnesium in titanium extraction, risks vs benefits of chlorination). • Interpret complex multi-step test-tube results, linking observations to chemical principles.
Aldehydes and Ketones	• Recognise aldehydes and ketones from structural/formula representations. • Recall that aldehydes are oxidised to carboxylic acids and ketones are not easily oxidised • Identify Fehling's solution and Tollens' reagent as tests for aldehydes. • Demonstrate they know that $NaBH_4$ reduces aldehydes/ketones to alcohols.	• Write balanced equations for oxidation of aldehydes, • reduction using $NaBH_4$ and KCN (use [H] and HCN as appropriate) • Outline the nucleophilic addition mechanisms for: • Reduction with $NaBH_4$ (showing H^- as nucleophile) • Reaction with KCN followed by dilute acid • Explain that reduction of aldehydes gives primary alcohols, ketones give secondary alcohols, KCN + aldehyde/unsymmetrical ketone forms a racemic mixture.	• Justify enantiomer formation via attack of CN^- on planar carbonyl group from either side. • Compare chemical vs spectroscopic methods for identifying carbonyls. • Predict outcomes and sketch mechanisms for unfamiliar nucleophilic addition scenarios. • Evaluate practical and safety considerations of using reagents like KCN in lab/industry.

The Rate Equation	• Recognise that rate depends on concentration. • Understand basic terms: rate, order, and rate constant. • Identify zero-, first-, and second-order relationships from simple rate–concentration data.	• Use the rate equation to calculate rates or rate constants (including units). • Use concentration–time or rate–concentration graphs to determine order. • Explain qualitatively how temperature affects rate constant (k). • Rearrange and use the Arrhenius equation: • $\ln k = -Ea/RT + \ln A$ to plot graphs and calculate activation energy.	• Deduce rate equations and orders from experimental data. • Use rate data to propose likely rate-determining steps in mechanisms. • Interpret complex graphs to calculate gradients and initial rates. • Solve problems involving logarithmic forms of the Arrhenius equation and extrapolate activation energy from graphical data.