KS5 Curriculum Map – Chemistry:

Topic	Knowledge *Substantive knowledge*: This is the specific, factual content for the topic, which should be connected into a careful sequence of learning.	Skills *Disciplinary knowledge*: This is the action taken within a particular topic in order to gain substantive knowledge.	Assessment Opportunities What assessments will be used to measure student progress?
Module 1: Development of practical skills in chemistry	• Planning, implementing, analysing and evaluating practical work. • Use of apparatus and techniques; safe working. • Making observations and measurements-recording and processing qualitative and quantitative data.	• Risk assessments, variables, identification of anomalies in data, precision and accuracy, uncertainty. • Following written instructions. • Setting up and use of appropriate equipment, evaluation of methods- limitations and suggestion of improvements. Presentation of data in a scientific way. • Using appropriate software and tools to process data, carry out research and present findings.	• PAG write-ups, • practical endorsement evidence, • exam-style practical questions • data-handling tasks • and evaluation questions.
Module 2: Foundations in chemistry **<u>Atomic Structure and isotopes</u>**	• Isotopes as atoms of the same element with different numbers of neutrons and different masses • Atomic structure in terms of the numbers of protons, neutrons and electrons for atoms and ions	• Recall relative charges and masses of protons, neutrons and electrons. • Calculate numbers of protons, neutrons and electrons in atoms and ions, given atomic number and mass number.	• Practical endorsement evidence • PAG 1.1, 1.2, 1.3 write-ups • Exam-style practical questions • Data-handling tasks and evaluation questions. • Homework • End-of topic tests • PPE

Relative mass	• Explanation of the terms relative isotopic mass and relative atomic mass. • The use of mass spectrometry in the determination of relative isotopic masses • The use of mass spectrometry in the determination of the relative atomic mass of an element from the relative abundances of its isotopes. • Use of the terms relative molecular mass, Mr, and relative formula mass and their calculation from relative atomic masses.	• Define relative atomic mass and relative isotopic mass. • Calculations using data from mass spectrometry to determine relative abundance, isotopic mass, relative atomic mass. • Calculation of relative molecular mass from relative atomic masses.	
Formulae and equations	• Determine ionic formula. • Construction of balanced chemical equations (including ionic equations), including state symbols, for reactions studied and for unfamiliar reactions given appropriate information.	• Determine ionic formula. • Construction of balanced chemical equations (including ionic equations), including state symbols, for reactions studied and for unfamiliar reactions given appropriate information.	
Amount of substance	• *The mole:* explanation and use of the terms: • Amount of substance • Mole (symbol 'mol'), as the unit for amount of substance • The Avogadro constant, NA • Molar gas volume • *Determination of formulae :*use of the terms: • empirical formula • molecular formula • *anhydrous*, *hydrated* and *water of crystallization.* • Relate atom economy to economic and environmental concerns.	• Calculate numbers of particles using Avogadro's Constant. • Calculations of empirical and molecular formulae, from composition by mass or percentage compositions by mass and relative molecular mass and elemental analysis. • *C*alculation of reacting masses, gas volumes and mole concentrations • Calculations, using amount of substance in mol, involving: (i) mass (ii) gas volume (iii) solution volume and concentration. • Use of the Ideal Gas Law in calculations. • Calculations involving reacting quantities, limiting reactants, percentage yield, limiting reactants and atom economy. • Calculation of reacting masses, gas volumes and mole concentrations • Relate atom economy to economic and environmental concerns.	

Acids	• *Acids, bases, alkalis, and neutralization*	• Know the formulae of the common acids (HCl, H_2SO_4, HNO_3 and CH_3COOH) and the common alkalis (NaOH, KOH and NH_3). • Explain that acids release H^+ ions in aqueous solution and alkalis release OH^- ions in aqueous solution. • Give qualitative explanation of strong and weak acids in terms of relative dissociations. • Explain neutralisation as the reaction of: (i) H^+ and OH^- to form H_2O (ii) acids with bases, including carbonates, metal oxides and alkalis (water-soluble bases), to form salts. • Write balanced equations for neutralization of acids with bases (metal oxides, carbonates and hydroxides) to form salts and additional products. • Write ionic equations for neutralization reactions. • Write ionic equations for the redox reaction of metals with acids. • Define standard solution. • The techniques and procedures used when preparing a standard solution of required concentration and carrying out acid–base titrations.	
	• *Acid–base titrations.*	• Complete structured and non-structured titration calculations, based on experimental results of familiar and non-familiar acids and bases.	• PAG 2.1, 2.2, 2.3 write-ups

Redox	• Know the rules for assigning and calculating oxidation number for atoms in elements, compounds and ions. • Using Roman numerals to indicate the magnitude of the oxidation number when an element may have compounds/ions with different oxidation numbers	• Calculating oxidation number of an element, writing formulae using oxidation numbers.	
	• *Redox reactions* • Identify oxidation and reduction in terms of: • electron transfer • changes in oxidation number	• The interpretation of redox equations of metals and acids to form salts and hydrogen gas and unfamiliar redox reactions, to make predictions in terms of oxidation numbers and electron loss/ gain.	
Electrons, Bonding and structure			
Electron structure	• Electron shells, orbitals and sub-shells, • spdf electron configuration of orbitals, • filling of orbitals • Long- and shorthand electron configuration • Electron configuration of ions	• Writing electron configuration using sub-shell notation- both long- and shorthand electron configuration.	
Ionic bonding	• Ionic bonding as the electrostatic attraction between positive and negative ions. • Structure of ionic compounds and associated properties.	• Construction of 'dot-and-cross' diagrams for ionic compounds. • Explanation of melting point and solubility of ionic compounds.	
Covalent bonding	• The covalent bond as the strong electrostatic attraction between a shared pair of electrons and the nuclei of the bonded atoms. • Use of the term average bond enthalpy as a measurement of covalent bond strength.	• Construction of 'dot-and-cross' diagrams of molecules and ions to describe: (i) single covalent bonding (ii) multiple covalent bonding (iii) dative covalent (coordinate) bonding.	

The shapes of simple molecules and ions	• The shapes of molecules with 2,3,4,5and 6 bonding pairs around a central atom. • Distortion of shape due to lone pairs on central atom	• Using electron pair repulsion theory to predict molecular shapes. Drawing shapes of simple molecules.	
Electronegativity and bond polarity	• Explain (i) a polar bond and permanent dipole within molecules containing covalently-bonded atoms with different electronegativities • Explain that apolar molecule has an overall dipole in terms of permanent dipole(s) and molecular shape.	• Define electronegativity. • Using ideas about electronegativity to predict chemical bond type. • Determine whether bonds are polar or non-polar and whether molecules are polar or non-polar.	
Intermolecular forces	• Intermolecular forces based on permanent dipole–dipole interactions and induced dipole– dipole interactions (London forces). • Hydrogen bonding as intermolecular bonding between molecules containing N, O or F and the H atom of –NH, –OH or HF	• Explanation of the solid structures of simple molecular lattices, as covalently bonded molecules attracted by intermolecular forces, e.g. I_2, ice. • Explain the effect of structure and bonding on the physical properties of covalent compounds with simple molecular lattice structures including melting and boiling points, solubility and electrical conductivity. • Explain the anomalous properties of H_2O resulting from hydrogen bonding, e.g.: (i) the density of ice compared with water (ii) its relatively high melting and boiling points.	

Module 3: Periodic Table and Energy **<u>The Periodic Table</u>**	• *<u>Periodicity</u>* • The structure of the periodic table - the periodic table as the arrangement of elements: (i) by increasing atomic (proton) number (ii) in periods showing repeating trends in physical and chemical properties (periodicity) (iii) in groups having similar chemical properties. • Periodic trend in electron configuration and ionisation energy across Periods 2 and 3 • The classification of elements into s-, p- and d-blocks. • Explaining Periodic trends in structure and melting point. • Metallic bonding and giant metallic lattices. • The physical properties of giant metallic and covalent lattices including melting point, boiling point, solubility and electrical conductivity.	• Writing equation for ionization energies including state symbols. • Predicting, either written or graphically, and explaining changes in ionisation energy across a Period (including dips) and down a Group, including dips. • Explaining Periodic trend in structure and melting point. • Explanation of: (i) metallic bonding as strong electrostatic attraction between cations (positive ions) and delocalised electrons (ii) a giant metallic lattice structure, e.g. all metals. • Explanation of the solid giant covalent lattices of carbon (diamond, graphite and graphene) and silicon as networks of atoms bonded by strong covalent bonds.	• Practical endorsement evidence • PAG 3.1, 3.2, 3.3 write-ups • Exam-style practical questions • Data-handling tasks and evaluation questions. • Homework • End-of topic tests • PPE's.
<u>Group 2</u>	• Periodic trend in melting point across periods 2 and 3 • Redox reactions :- the outer shell 2s electron configuration and the loss of these electrons in redox reactions to form 2+ ions	• Explained by the change in structure from giant metallic to giant covalent to simple molecular lattice. • Explain the relative reactivities of the Group 2 elements Mg → Ba shown by their redox reactions with: (i) oxygen (ii) water (iii) dilute acids • The trend in reactivity in terms of the first and second ionisation energies of Group 2 elements down the group.	

	• Reactivity of Group 2 metals :-The action of water on Group 2 oxides and the approximate pH of any resulting solutions, including the trend of increasing alkalinity. • Uses of some Group 2 compounds as bases, including equations, for example (but not limited to): (i) $Ca(OH)_2$ in agriculture to neutralise acid soils (ii) $Mg(OH)_2$ and $CaCO_3$ as 'antacids' in treating indigestion.	• Write balanced symbol equations for the reaction of the Group 2 metals and metal oxides with water and their oxides and hydroxides as bases.	
The Halogens	• Characteristic physical properties. • Redox reactions and reactivity of halogens and their compounds • Disproportionation • Characteristic reactions of halide ions	• Explain the existence of halogens as diatomic molecules and explanation of the trend in the boiling points of Cl_2, Br_2 and I_2, in terms of induced dipole–dipole interactions (London forces) • Explain reactivity in terms of the outer shell s^2p^5 electron configuration and the gaining of one electron in many redox reactions to form 1– ions. • Explain the trend in reactivity of the halogens Cl_2, Br_2 and I_2, illustrated by reaction with other halide ions. • Explanation of the trend in reactivity from the decreasing ease of forming 1– ions, in terms of attraction, atomic radius and electron shielding. • Explanation of the term disproportionation as oxidation and reduction of the same element, illustrated by: (i) the reaction of chlorine with water as used in water treatment (ii) the reaction of chlorine with cold, dilute aqueous sodium hydroxide, as used to form bleach. • Observations and explanations of the precipitation reactions, including ionic equations, of the aqueous anions Cl^-, Br^- and I^- with aqueous silver ions, followed by aqueous ammonia, and their use as a test for different halide ions.	

Qualitative analysis **Physical Chemistry** **Enthalpy Changes**	• Qualitative tests for anions (a) qualitative analysis of ions on a test-tube scale; • To know that some chemical reactions are accompanied by enthalpy changes that are exothermic (ΔH, negative) or endothermic (ΔH, positive). • Explanation and use of the following terms; standard conditions and standard states, enthalpy change of reaction, enthalpy change of formation, enthalpy change of combustion, enthalpy change of neutralization. • Average bond enthalpy and exothermic and endothermic reactions in terms of enthalpy changes associated with the breaking and making of chemical bonds. • Hess' law for construction of enthalpy cycles.	• Writing balanced formula and ionic equations (including state symbols) for precipitation reactions. • Observations and writing of full balanced equations and ionic equations (including state symbols) for these reactions. • Processes and techniques needed to identify the following ions in an unknown compound: (i) anions: • CO_3^{2-}, by reaction with $H^+(aq)$ forming $CO_2(g)$ • SO_4^{2-}, by precipitation with $Ba^{2+}(aq)$ • Cl^-, Br^-, I^- (ii) cations: NH_4^+, by reaction with warm NaOH(aq) forming NH_3. • Construction of enthalpy profile diagrams to show the difference in the enthalpy of reactants compared with products. • Determination of enthalpy changes directly from appropriate experimental results, including use of the relationship: $q = mc\Delta T$. • The use of average bond enthalpies to calculate enthalpy changes and related quantities. • Calculations to determine indirectly: (i) an enthalpy change of reaction from enthalpy changes of combustion (ii) an enthalpy change of reaction from enthalpy changes of formation (iii) enthalpy changes from provided unfamiliar enthalpy cycles. • The techniques and procedures used to determine enthalpy changes directly and indirectly. • Calculation of reaction rate from the gradients	• PAG 4.1, 4.2, 4.3 write-ups

Reaction Rates	• Collision theory and the factors that affect the frequency of collisions.	of graphs measuring how a physical (concentration, mass, gas volume) quantity changes with time. The techniques and procedures used to investigate reaction rates including the measurement of mass, gas volumes and time.	
	• Catalysts- explanation of the role of a catalyst, explanation of the terms homogeneous and heterogeneous catalysts, • Explaining how catalysts have economic importance and benefits for increased sustainability.	• Drawing an uncatalyzed and catalysed reaction profile.	
Chemical Equilibrium	• The Boltzmann distribution; a qualitative explanation of the Boltzmann distribution and its relationship with activation energy- and how it changes with temperature and in the presence of a catalyst	• Drawing and use of Boltzmann distribution model diagrams to explain effects on reaction rates.	
	• Dynamic equilibrium and le Chatelier's principle. • The importance to the chemical industry of a compromise between chemical equilibrium and reaction rate in deciding the operational conditions • The effect of catalysts on the position of equilibrium	• The application of Le Chatelier's principle to homogeneous equilibria to deduce qualitatively the effect of a change in temperature, pressure or concentration on the position of equilibrium.	
	• The equilibrium constant, Kc; • Qualitative estimation of the position of equilibrium from the magnitude of Kc.	• The techniques and procedures used to investigate changes to the position of equilibrium for changes in concentration and temperature.	

Module 4: Core organic chemistry **<u>Basic concepts and hydrocarbons</u>**	• *Naming and representing the formulae of organic compounds.* • Interpretation and use of the terms: (i) general formula (ii) structural formula (iii) (iv) skeletal formula. • Use of the terms homologous series, functional group, aliphatic, alicyclic, aromatic, saturated and unsaturated. • *Structural isomerism* as exhibited in compounds with the same molecular formula but different structural formulae. • Use of the terms reaction mechanism, homolytic fission, heterolytic fission, curly arrows and reaction mechanism.	• Use of IUPAC rules for naming organic compounds. • Representing organic compounds using the various types of formula. • Writing equations using the different types of organic formula • Identify the functional group in organic compounds. • Use of the general formula of a homologous series to predict the formula of any member of the series. • Identifying and drawing structures of isomers. • Determination of possible structural formulae of an organic molecule, given its molecular formula. • Drawing mechanisms using the appropriate curly arrow to show the moment of electrons and the breaking/ making of bonds.	• Practical endorsement evidence • 5.1, 5.2, 5.3, 5.4 write-ups • Exam-style practical questions • Data-handling tasks and evaluation questions. • Homework • End-of topic tests • PPE's.
	• *Alkanes*- molecular shape, physical properties. Reactions of alkanes- combustion and free-radical substitution. Limitations of free-radical substitution.	• Naming and drawing of alkanes. • Explaining the trend in boiling points in terms of increasing magnitude of intermolecular forces. • Explaining how branching effects boiling point. • Writing balanced combustions equations for the alkanes. Writing balanced-free radical substitution reactions of alkanes. Drawing mechanisms for the free-radical substitution reactions of alkanes.	

	• *Alkenes*- bonding in alkenes, properties of alkenes. Stereoisomerism in alkenes. • Addition reactions of alkenes with electrophiles. Using Markownikoff's rule to predict major and minor products from addition reactions of H-X to unsymmetrical alkenes. • Addition polymerization of alkenes including the benefits for sustainability of processing waste polymers and the benefits to the environment of development of biodegradable and photodegradable polymers.	• Naming and drawing alkene structures. Identifying σ- and π bonds in organic molecules. • Explain how E/Z and cis/ trans isomers arise • Identify and draw E/Z and cis/ trans isomers of alkenes. • Writing balanced equations for addition reactions of alkenes. • Drawing mechanism for the electrophilic addition reactions of alkenes. • Identifying monomers from polymer chains	
Alcohols	• Structure and properties of alcohols • Combustion reactions of alcohols. • Oxidation reactions of alcohols. • Elimination reactions of alcohols. • Substitution reactions of alcohols.	• Naming alcohols. • Explaining physical properties of alcohols. Classifying alcohols as primary, secondary or tertiary. • Writing balanced equations for the combustion of alcohols. • Writing equations for the oxidation of alcohols using the different types of organic formula. • Identifying and drawing the dehydration products of alcohols using the different types of organic formula. • Writing equations and drawing structures for the substitution reactions of alcohol with halides.	

Haloalkanes	• Reactivity of haloalkanes. • Nucleophilic substitution reactions of haloalkanes. • Hydrolysis of haloalkanes by aqueous alkali and by water. • The trend in reactivity of haloalkanes. • Environmental concerns from use of organohalogen compounds.	• Naming haloalkanes. • Classifying haloalkanes as primary, secondary or tertiary. • Explaining the reactivity of haloalkanes and the trend in reactivity of haloalkanes. • Writing balanced equations for the hydrolysis of haloalkanes. • Drawing structures and mechanisms for hydrolysis reactions of haloalkanes and other nucleophilic displacement reactions. • Writing equating and drawing mechanisms for ozone depletion in the atmosphere.	
Organic synthesis	• Synthetic routes	• Learners will be expected to be able to devise two stage synthetic routes by applying transformations between all functional groups. • Identify and draw intermediates in synthetic routes. • Identify reagents used in synthetic routes.	
	• Practical techniques	• The techniques and procedures used to for reflux, distillation, drying and the purification of organic liquids.	
Analytical techniques	• Mass spectrometry	• Analysis of mass spectrometry and using fragmentation patterns to identify the structure of organic molecules. Using M and M+1 peaks to identify the number of carbon atoms in an organic molecule.	

	• *Infrared spectroscopy* • How infrared spectroscopy works • The absorption of IR radiation by atmospheric gases. • The use of analytical techniques to provide evidence for law courts, e.g. drink driving and to monitor gases causing air pollution • Combined techniques	• Explain how the absorption of IR radiation by chemical bonds results in bonding and stretching of the covalent bonds in an organic molecule • Identifying -OH, C=O, CO_2H functional groups from IR spectroscopy • The deduction of the structures of organic compounds from different analytical data including (i) elemental analysis (ii) mass spectra (iii) IR spectra.	
Module 5: Physical Chemistry and Transition Elements **How fast?** **How far?**	• Orders, rate equations and rate constants. • Rate graphs and orders. • The rate-determining step • The Effect of temperature on rate constants- the Arrhenius equation • Equilibrium • The qualitative effect on equilibrium constants of changing temperature for exothermic and endothermic reactions. • The constancy of equilibrium constants with changes in concentration, pressure or in the presence of a catalyst	• Determining orders, rate equations and rate constants from initial rates data. • Identifying orders from the shape of concentration-time graphs and rate-concentration graphs. • Clock reactions • Using first order graphs to determine half-life • Graphical analysis and use of the Arrhenius equation to determine (i) the activation energy of a reaction (ii) A, the pre-exponential factor and other associated calculations. • Using the rate equation, balanced equation and mechanistic data to determine the rate determining step. • Identifying intermediates in mechanisms from rate data. • Calculating equilibrium concentrations • Calculating equilibrium partial pressures • Calculating Kc and Kp • Qualitative explanation of Le Chatelier's principle in terms of Kc and Kp	• Practical endorsement evidence • Exam-style practical questions • Data-handling tasks and evaluation questions. • Homework • End-of topic tests • PPE • PAG 9.1, 9.2, 9.3 write-ups • PAGS 10.1, 10.2, 10.3 write-ups

Acids bases and buffers	• Brønsted–Lowry acids and bases • pH, strong acids and [H+(aq)] • Weak acids, Ka and pKa • Kw and strong bases • Buffers: action, uses and calculations • Neutralisation	• Identifying conjugate bases of acids • Writing balanced symbol and ionic equations for the reactions of acids with metals, metal carbonates, metal oxides and alkalis (metal hydroxides) • Calculating pH from $[H^+]$ for strong acids • Compare weak acid strength from Ka and pKa data. • Calculating pH given the Ka and acid concentration for a weak acid • State the approximations made when calculating the pH of a weak acid using Ka • Calculate the pH of a strong base using Kw or pKb. • Define what a buffer is. • Explain how a buffer controls small changes in the pH of a solution • Calculate the pH of a buffer solution and associated calculations. • Use of the Henderson- Hasselbach equation to determine the [A-]/[HA ratio in a buffer system • Explain how the bicarbonate /carbonic acid buffer in the body controls blood pH. • Use the shape of a pH curve to determine the acid and base type employed in the titration. • Use pH curves to determine reacting volumes, endpoint and equivalence point in a titration. • Use the pH range of an indicator to determine its suitability for a given titration.	• PAG 11.1,11.2, 11.3 write-ups
Enthalpy and entropy	• Lattice enthalpy • Enthalpy of solution and hydration	• Calculate standard entropy changes • Draw or label Born-Haber cycles given appropriate information. • Use Born-Haber cycles to calculate lattice enthalpy or associated energy values in the energy cycle. • Calculate enthalpy change of solution, or hydration given appropriate data.	

		• Explain how melting point, lattice enthalpy and enthalpy of hydration is affected by charge and size of ion	
	• *Entropy* - a measure of the dispersal of energy in a system which is greater, the more disordered a system. • *Free energy*- Feasibility and spontaneous processes depends upon the entropy change and temperature in the system, TΔS, and the enthalpy change of the system, ΔH. The limitations of predictions made by ΔG about feasibility	• Explanation of the difference in magnitude of the entropy of a system: (i) of solids, liquids and gases (ii) for a reaction in which there is a change in the number of gaseous molecules. • Predict entropy changes associated with change of physical state • Entropy change calculations. • Explanation, and related calculations, of the free energy change. • Determination of a temperature at which a reaction becomes feasible • Limitations of predictions of feasibility	
Redox and electrode potentials	• Oxidising and reducing agents. • Redox equations. • Redox titrations. • *Electrode potentials and* prediction of the feasibility of a reaction using standard cell potentials and the limitations of such predictions in terms of kinetics and concentration. • *Storage and fuel cells:* application of principles of electrode potentials to modern storage cells. Benefits of electrochemical cells counteracted by risks from toxicity and fire from Li-based cells.	• Construction of redox equations using half-equations and oxidation numbers. • Interpretation and prediction of reactions involving electron transfer. • Techniques and procedures used when carrying out redox titrations. • Structured and non-structured titration calculations • Drawing electrochemical cells. • The techniques and procedures used for the measurement of cell potentials. • Calculation of a standard cell potential by combining two standard electrode potentials. • Explanation that a fuel cell uses the energy	• PAG 12.1, 12.2, 12.3 write-ups • PAG 8.1, 8.2, 8.3 write-ups

Transition elements	• Properties, • electron configuration, • benefits from acting as catalysts. • Ligands and complex ions, stereoisomerism in complex ions. • Use of cis-platin as an anti-cancer drug and its action by binding to DNA preventing cell division ligand. • Ligand substitution reactions of complex ions • precipitation reactions of complex ions. • Redox chemistry and qualitative analysis	from the reaction of a fuel with oxygen to create a voltage and the changes that take place at each electrode. • Use sub-shell notation. • Learners should be able to draw 3-D diagrams to illustrate stereoisomerism. • Observation and writing equations to reflect ligand substitution reactions and the accompanying colour changes. • Observations and writing full and ionic equations for the reactions of aqueous Cu^{2+}, Fe^{2+}, Fe^{3+}, Mn^{2+} and Cr^{3+} with aqueous sodium hydroxide and aqueous ammonia. • Construct and interpret redox equations using relevant half-equations and oxidation numbers. • Processes and techniques needed to identify the following ions in an unknown compound: (i) anions: CO_3^{2-}, Cl^-, Br^-, I–, SO_4^{2-} (ii) cations: NH_4^+; Cu^{2+}, Fe^{2+}, Fe^{3+}, Mn^{2+}, Cr^{3+}	
Module 6: Organic Chemistry and Analysis **Aromatic compounds** **Carbonyl compounds**	• Kekule and delocalized model of bonding in benzene. • Electrophilic substitution reactions of benzene and phenol. • Directing effects of substituents. • Reactions of carbonyl compounds • oxidation of aldehydes, • nucleophilic addition reactions with $NaBH_4$	• Naming of aromatic compounds. • Writing balanced equations for aromatic electrophilic substitution reactions of benzene and derivatives. • Drawing mechanisms for electrophilic substitution reactions of benzene. • Naming aldehydes and ketones. • Writing balanced equations for the reactions of aldehydes and ketone.	• Practical endorsement evidence • PAG 6.1, 6.2, 6.3 write-ups • Exam-style practical questions • Data-handling tasks and evaluation questions. • Homework • End-of topic tests • PPE.

	and HCN • chemical test for aldehyde and ketone functionality.	• Drawing mechanisms for the nucleophilic addition reactions with $NaBH_4$ and HCN. • Writing observations for chemical tests with Tollens Reagent and Brady's reagent (2,4-DNPH).	
Carboxylic acids and esters	• Properties of carboxylic acids- water solubility and high boiling points. • Reactions in aqueous conditions of carboxylic acids with metals and bases. • Ester formation and hydrolysis.	• Writing balanced equations for reactions of acids and bases. • Writing balanced equations for esterification between a carboxylic acid and an alcohol and the hydrolysis of esters either using aqueous acid or alkali	
Acid chlorides	• Formation from carboxylic acids and thionyl chloride, $SOCl_2$. • The use of acyl chlorides in synthesis in formation of esters, carboxylic acids and primary and secondary amides	• Writing balanced equations for reactions between carboxylic acids and $SOCl_2$, and between acid chlorides and alcohols, phenols, amines and water.	
Amines	• The basicity of amines and the formation of salts with acids. • The preparation of aliphatic and aromatic amines. • The structure of amino acids. Reactions of the carboxyl group and amino group.	• The classification of amines as primary, secondary, tertiary and quaternary. • Naming amines. • Writing balanced equations for the formation of primary amines from ammonia and secondary/tertiary amines from amine from reactions with haloalkanes. • Writing balanced equations for the reduction of nitrobenzenes to give aromatic amines.	
Amino acids	• The general structure of amino acids	• Writing equations using appropriate organic formula for the reactions of the amino group and carboxylic acid group in amino acids. • Drawing zwitterions.	
Amides	• Structures of primary and secondary amides	• The classification of amides as primary, secondary and tertiary. • Naming of amides. • Writing balanced equations for the acid and alkaline hydrolysis of amides.	

Chirality	• Optical isomerism as an example of stereoisomerism, in terms of non-superimposable mirror images about a chiral centre.	• Identifying chiral centres in molecules. • Using flying wedge notation when drawing optical isomers.	
Condensation polymers	• The use of condensation polymerisation to form polyesters and amides	• Identification and drawing structures of the products of the condensation of monomers each containing two functional groups or one monomer with two functional groups. • The identification and drawing structures of the products acid and base hydrolysis of polyesters and polyamides. • Prediction, and writing equations and drawing structures from addition and condensation polymerisation of: (i) the repeat unit from a given monomer(s) (ii) the monomer(s) required for a given section of a polymer molecule (iii) the type of polymerisation.	
Organic synthesis	• Carbon–carbon bond formation. • The use of C–C bond formation in synthesis using the nitrile (CN group) to increase the length of a carbon chain. • The reactions of nitriles	• Write balanced equations and draw mechanisms for the nucleophilic substitution of haloalkanes with CN^- ions. • Writing balanced equations for the reduction of nitriles to form amines and hydrolysis to give carboxylic acids.	
	• Friedel crafts reactions.	• The writing of balanced equations to give alkylated and acylated aromatic compounds.	
	• Further practical techniques	• The techniques and procedures used for the preparation and purification of organic solids. • Melting point determination of an organic solid.	
	• Further Synthetic routes • Designing synthetic routes for an organic molecule containing several functional groups: (i) identification of individual functional groups (ii) prediction of properties and reactions. •	• Devising multistage synthetic routes by applying transformations between all functional groups studied throughout the specification.	

<u>Chromatography and spectroscopy</u>	• Multi-stage synthetic routes for preparing organic compounds. • <u>Chromatography and qualitative analysis</u> • Thin -layer chromatography. • Gas Chromatography • Qualitative tests for organic functional groups on a test-tube scale. • <u>Spectroscopy</u> • NMR Spectroscopy- analysis of proton decoupled ^{13}C NMR • Analysis of proton NMR. • Combined techniques for the deduction of the structures of organic compounds from different analytical data	• Interpretation of gas chromatograms- including creation and use of external calibration curves to confirm concentrations of components. • Processes and techniques needed to identify functional groups in an unknown compound. • Analysis of a high-resolution NMR spectra of an organic molecule to make suggest structures of molecules. Predicting NMR spectra from a given structure of an organic compound. Determining the presence of exchangeable protons using deuterated solvents. • Analysis of data including elemental analysis, mass spectra, IR spectra and NMR spectra.	