Module 1 – Development of practical skills - Personal Learning Checklist

Learning Objectives:	Confidence		
Planning			
Design experiments to solve problems set in a practical context.			
Identify variables that must be controlled			
Evaluate experimental methods to ensure they are appropriate to meet the expected outcomes.			
Implementing			
Use a wide range of practical apparatus and techniques correctly			
Use appropriate units for measurements			
Present observations and data in an appropriate format			
Analysis			
Process, analyse and interpret qualitative and quantitative experimental results.			
Use appropriate significant figures			
Plot and interpret suitable graphs from experimental results, including; - Selection and labelling of axes with appropriate scales, quantities and units. - Measurement of gradients and intercepts.			
Evaluation			
Evaluate results and draw conclusions			
Identify anomalies in experimental measurements			
Understand the limitations in experimental procedures			
Understand precision and accuracy of measurements and data, including margins of error, percentage errors and uncertainties in apparatus.			
Refine experimental design by suggesting improvements to the procedures and apparatus.			
Independent Thinking			
Apply investigative approaches and methods to practical work.			
Use and Application of Scientific Methods and Practices			
Safely and correctly use a range of practical equipment and materials.			
Safely and correctly use a range of practical equipment and materials.			
Follow written instructions.			
Keep appropriate records of experimental activities.			
Present information and data in a scientific way.			
Use appropriate software and tools to process data, carry out research and report findings.			

Research and Referencing			
Use online and offline research skills including websites, textbooks and other printed scientific sources of information.			
Correctly cite sources of information.			
Instruments and Equipment			
Use a wide range of experimental and practical instruments, equipment and techniques appropriate to the knowledge and understanding included in the specification.			
Use of Apparatus and Techniques			
Use appropriate apparatus to record a range of measurements to include mass, time, volume of liquids and gases and temperature.			
Use of a water bath, electric heater or sand bath for heating			
Measure pH using pH charts, a pH meter or pH probe on a data logger.			
Use laboratory apparatus for a variety of experimental techniques including; - Titration, using burette and pipette - Distillation and heating under reflux, including setting up glassware using retort stand and clamps - Qualitative tests for ions and organic functional groups - Filtration under reduced pressure			
Use of a volumetric flask, including accurate techniques for making up a standard solution.			
Use of acid-base indicators in titrations of weak/strong acids with weak/strong alkalis			
Purification of a - Solid product by recrystallization - Liquid product, including use of a separating funnel			
Use of melting point apparatus.			
Use of thin layer or paper chromatography.			
Set up electrochemical cells and measure voltage.			
Safety and carefully handle solids and liquids, including corrosive, irritant, flammable and toxic substances.			
Measure rates of reaction by at least two different methods, including - An initial rate method such as a clock reaction - A continuous monitoring method			

Module 2 - Foundations in Chemistry - Personal Learning Checklist

	Confidence		
Learning Objectives:			
Atomic Structure and Isotopes (Chapter 2)			
Define the term *isotope*.			
Describe atomic structure in terms of the numbers of protons, neutrons and electrons for atoms and ions, given the atomic number, mass number and any ionic charge.			
Explain the terms *relative isotopic mass* and *relative atomic mass*, based on the mass of a ^{12}C atom.			
Describe how to determine relative isotopic masses and relative abundances of the isotope using mass spectrometry.			
Calculate the relative atomic mass of an element from the relative abundances of its isotopes.			
Use of the terms *relative molecular mass*, M_r, (for simple molecules) and relative formula mass (for giant structures) and calculate their values from relative atomic masses.			
Write formulae of ionic compounds from ionic charges.			
Predict ionic charge from the position of an element in the periodic table.			
Recall the names and formulae for the following ions: NO_3^-, CO_3^{2-}, SO_4^{2-}, OH^-, NH_4^+, Zn^{2+}, and Ag^+.			
Construct balanced chemical equations (including ionic equations), including state symbols, for reactions studied and for unfamiliar reactions given appropriate information.			
Amount of Substance (Chapter 3) -			
Predict ionic charges from the position of an element on the periodic table.			
Recall the names and formulae for the following ions: nitrate, carbonate, sulfate, hydroxide, ammonium, zinc and silver.			
Write formulae for ionic compounds from ionic charges.			
Write balanced chemical equations (full and ionic), including state symbols for familiar reactions and for unfamiliar reactions when given information.			
Understand and use the following terms correctly:			
• amount of substance			
• mole			
• Avagadro constant			
• molar mass			
• molar gas volume			
• empirical formula			
• molecular formula			
• anhydrous			
• hydrated			
• water of crystallisation			
Calculate empirical formula from data giving composition by mass or % by mass.			
Calculate molecular formula from the empirical formula and M_r.			
Calculate the formula of a hydrated salt from data giving composition by mass or % by mass.			

Carry out calculations using:			
- the Avogadro constant			
- mass of substance, M_r and amount in moles			
- concentration, volume and amount of substance in a solution			
Use balanced equations to calculate:			
- masses			
- volumes of gases			
- % yields			
- % atom economy			
- concentrations and volumes of solutions			
State the ideal gas equation.			
State the correct SI unit for each variable in the ideal gas equation.			
Convert values into the correct unit for the ideal gas equation.			
Rearrange the ideal gas equation and use it to calculate p, V, n and T.			
Describe how to make up a volumetric solution.			
Describe how to carry out an acid-base titration.			
Discuss the benefits for sustainability of developing chemical processes with a high atom economy.			
Acids (Chapter 4) -			
Recall the names and formulae of the following common acids: HCl, H_2SO_4, HNO_3 and CH_3COOH.			
Recall the names and formulae of the following common alkalis: NaOH, KOH and NH_3.			
Explain that acids release H^+ ions in aqueous solution.			
Explain that alkalis release OH^- ions in aqueous solution.			
Explain the difference between strong and weak acids in terms of relative dissociations.			
Describe neutralisation as the reaction of H^+ and OH^- to form H_2O, including an ionic equation.			
Describe neutralisation reactions of acids with bases, including carbonates, metal oxides and alkalis (water-soluble bases), to form salts, including full equations.			
Describe the techniques and procedures used to prepare a standard solution of required concentration.			
Describe the techniques and procedures used to carry out acid–base titrations.			
Carry out structured and non-structured titration calculations, based on experimental results of familiar and non-familiar acids and bases.			
State and apply the rules for assigning and calculating oxidation number for atoms in elements, compounds and ions (including O in peroxides and H in metal hydrides).			
Write chemical formulae using oxidation numbers.			
Use Roman numerals to indicate the magnitude of the oxidation number when an element has compounds/ions with different oxidation numbers.			
Describe oxidation and reduction in terms of electron transfer.			
Describe oxidation and reduction in terms of changes in oxidation number.			
Describe redox reactions of metals with acids to form salts, including full equations.			
Interpret redox equations and unfamiliar redox reactions, to make predictions in terms of oxidation numbers and electron loss/gain.			

Electrons, Bonding and Structure (Chapter 5) -			
Electronic Configuration -			
State the number of electrons that can fill the first four shells.			
Define the term orbital.			
State the shape of s- and p-orbitals.			
State the number of orbitals that make up s-, p- and d-sub-shells and the number of electrons that they can hold.			
Describe how the orbitals fill.			
Deduce the electron configurations for atoms and ions up to atomic number $Z = 36$ and represent them using the ‘electrons in box’ and sub-shell notations.			
Ionic Bonding -X			
Describe what ionic Xbonding is.			
Draw ‘dot-and-cross’ diagrams for ionic compounds.			
Explain the solid structures of giant ionic lattices, e.g. NaCl.			
Use knowledge of structure and bonding to explain the physical properties of ionic compounds, including: • melting and boiling points; • solubility; • electrical conductivity in solid, liquid and aqueous states.			
Covalent Bonding -			
Describe what a covalent bond is.			
Draw ‘dot-and-cross’ diagrams of molecules and ions to show single, multiple and dative (coordinate) covalent bonding (up to six electron pairs, including lone pairs, surrounding a central atom).			
Relate *average bond enthalpy* to covalent bond strength.			

Shapes of Molecules and Intermolecular Forces (Chapter 6)-			
Shapes of Simple Molecules and Ions -			
Use electron pair repulsion to explain the following shapes of molecules and ions: linear, non-linear, trigonal planar, pyramidal, tetrahedral and octahedral.			
Use electron pair repulsion theory to predict the shapes of, and bond angles in, molecules and ions with up to six electron pairs, including lone pairs, surrounding a central atom.			
Draw 3-D diagrams to show the shapes of molecules and ions.			
Describe the relative repulsive strengths of bonded pairs and lone pairs of electrons.			
Explain how the relative repulsive strengths affect the bond angles in molecules, e.g. CH_4, NH_3 and H_2O.			
Electronegativity and Bond Polarity -			
Define the term *electronegativity.*			
Interpret Pauling electronegativity values.			
Describe the trends in electronegativity across the periodic table.			
Use electronegativity to predict chemical bond type.			
Describe what polar bonds and permanent dipoles are and explain why they are formed.			
Explain why some molecules that contain polar bonds have a permanent dipole (e.g. H_2O) but others do not (e.g. CO_2) .			
Predict whether a molecule will be polar or non-polar			
Intermolecular Forces -			
Describe how permanent dipole-dipole interactions, induced dipole-dipole interactions and hydrogen bonds form (including the role of lone pairs in H-bonding).			
Draw diagrams to represent the hydrogen bonding between molecules.			
Deduce the type of intermolecular forces that would occur between given molecules.			
Explain the anomalous properties of H_2O, e.g. the density of ice compared with water and its relatively high melting and boiling points.			
Explain how intermolecular forces influence the solid structures of simple molecular lattices, e.g. I_2, ice.			
Use knowledge of structure, bonding and intermolecular forces to explain the physical properties of covalent compounds with simple molecular lattice structures, including: • melting and boiling points; • solubility; • electrical conductivity.			

Module 3 – Periodic Table and Energy - Personal Learning Checklist

Learning Objectives:	**Confidence**		
Periodicity (Chapter 7) - INORGANIC			
The structure of the periodic table -			
Describe the periodic table as the arrangements of elements by increasing atomic number.			
Describe the periodic table as the arrangement of elements in periods showing repeating trends in physical and chemical properties (periodicity).			
Describe the periodic table as the arrangement of elements in groups having similar chemical properties.			
Periodic trends in electron configuration and ionisation energy -			
Explain the periodic trend in electron configurations across Periods 2 and 3.			
Classify elements into s-, p- and d-blocks.			
Define *first ionisation energy*.			
Describe first ionisation energy and successive ionisation energies.			
Describe the trend in first ionisation energies across periods 2 and 3.			
Explain the trend in first ionisation energies across periods 2 and 3, in terms of attraction, nuclear charge and atomic radius.			
Describe the trend in first ionisation energies down a group.			
Explain the trend in first ionisation energies down a group, in terms of attraction, nuclear charge and atomic radius.			
Explain the anomalies (small decreases) in first ionisation energies from group 2 to group 3 as a result of s- and p-sub-shell energies (e.g. between Be and B).			
Explain the anomalies (small decreases) in first ionisation energies from group 5 to group 6 as a result of p-orbital repulsion (e.g. between N and O).			
Predict the number of electrons in each shell of an atom and the group of an element from successive ionisation energies.			
Periodic trends in structure and melting point -			
Explain metallic bonding as strong electrostatic attraction between cations (positive ions) and delocalised electrons.			
Explain a giant metallic lattice structure (e.g. all metals).			
Explain solid giant covalent lattices of carbon (diamond, graphite and graphene) and silicon as networks of atoms bonded by strong covalent bonds.			
Describe the physical properties of giant metallic lattices, including melting and boiling points, solubility and electrical conductivity in terms of structure and bonding.			
Describe the physical properties of giant covalent lattices, including melting and boiling points, solubility and electrical conductivity in terms of structure and bonding.			
Explain the variation in melting points across Periods 2 and 3 in terms of structure and bonding.			

Reactivity Trends (Chapter 8) - INORGANIC			
Group 2 -			
Describe what happens to Group 2 atoms during redox reactions, in terms of electrons.			
Describe the relative reactivities of the Group 2 elements from Mg to Ba, based on their redox reactions with oxygen, water and dilute acids. (Reaction with acid limited to those that produce a salt and hydrogen.)			
Describe the trend in reactivity down Group 2 in terms of the first and second ionisation energies. (Definition for second ionisation NOT required, but should be able to write an equation for the change.)			
Describe the action of water on Group 2 oxides and the approximate pH of resulting solutions.			
Describe the trend in alkalinity of Group 2 oxides.			
Describe some uses of Group 2 compounds as bases, including (but not limited to): • $Ca(OH)_2$ in agriculture to neutralise acid soils • $Mg(OH)_2$ and $CaCO_3$ as 'antacids' in treating indigestion.			
The Halogens -			
Describe halogens a diatomic molecules.			
Explain the trend in boiling points down Group 7 in terms of intermolecular forces.			
Describe what happens to Halogen atoms during redox reactions, in terms of electrons.			
Describe and explain the trend in reactivity of the halogens in terms of: • attraction • atomic radius • electron shielding.			
Explain the term *disproporionation* as illustrated by: • the reaction of chlorine with water as used in water treatment • the reaction of chlorine with cold, dilute aqueous sodium hydroxide, as used to form bleach • similar reactions to those above.			
Compare the benefits of using chlorine in water treatment to the risks.			
Describe and write equations for the precipitation reactions of aqueous halide ions with aqueous silver ions, followed by ammonia, and the use of this as a test for halide ions.			
Qualitative Analysis -			
Describe the processes and techniques needed to identify the following ions in unknown compounds: • CO_3^{2-} • SO_4^{2-} • Cl^-, Br^-, I^- • NH_4^+			

Enthalpy Changes (Chapter 9) - INORGANIC			
Explain the enthalpy changes associated with endothermic and exothermic reactions in terms of bond breaking and making.			
Construct enthalpy profile diagrams to show the difference in the enthalpy of reactants compared to products.			
Explain the term *activation energy*, including use of enthalpy profile diagram.			
Explain and use the terms:			
• *standard conditions* and *standard states*			
• *enthalpy change of reaction* ($\Delta_r H$)			
• *enthalpy change of formation* ($\Delta_f H$)			
• *enthalpy change of combution* ($\Delta_c H$)			
• *enthalpy change of neutralisation* ($\Delta_{neut} H$)			
• *average bond enthalpy*			
Determine the enthalpy changes directly from experimental results, including the use of **$q = mc\Delta T$**.			
Calculate enthalpy changes and related quantities from average bond enthalpies.			
Use Hess' Law to construct enthalpy cycles and calculate the following indirectly:			
• an enthalpy change of reaction from enthalpy changes of combustion			
• an enthalpy change of reaction from enthalpy changes of formation			
• enthalpy changes from unfamiliar enthalpy cycles.			
Describe the techniques and procedures used to determine enthalpy changes directly and indirectly.			
Reaction Rates and Equilibrium (Chapter 10) -			
Reaction Rates -			
Describe the effect of concentration, including pressure of gases, on the rate of a reaction in terms of frequency of collisions.			
Calculate the rate of reaction from the gradient of a graph.			
Explain the role of a catalyst in: • increasing reaction rate without being used up by the overall reaction • in allowing a reaction to proceed via a different route.			
Explain the terms: • *homogeneous catalyst* • *heterogeneous catalyst.*			
Explain the economic importance of catalysts and benefits for sustainability.			
Describe the techniques and procedures used to investigate reaction rates including the measurement of mass, gas volumes and time.			
Explain the Boltzmann distribution and its relationship with activation energy.			
Use Boltzmann distributions to explain the qualitative effect on the proportion of molecules exceeding the activation energy and hence the reaction rate for: • temperature changes • catalytic behaviour.			

Equilibrium -			
Explain what *dynamic equilibrium* is.			
Apply le Chatelier's principle in homogeneous equilibria to deduce the effect of change in: • temperature • pressure • concentration on the position of equilibrium.			
Explain why a catalyst does not affect the equilibrium position.			
Describe the techniques and procedures used to investigate changes to the position of equilibrium for changes in concentration and temperature.			
Explain why it is important for the chemical industry to consider a compromise between chemical equilibrium and reaction rate in deciding the operational conditions.			
Give expressions for the equilibrium constant, K_c, for homogeneous reactions.			
Calculate the equilibrium constant, K_c, from provided equilibrium concentrations. (Do not need to determine units for K_c.)			
Estimate the position of equilibrium from the magnitude of K_c. (qualitative only).			

Module 4 – Core Organic Chemistry - Personal Learning Checklist

Learning Objectives:	**Confidence**		
Basic Concepts of Organic Chemistry (Chapter 11) -			
Use IUPAC rules of nomenclature to systematically name organic molecules with up to 10 carbon atoms in the longest chain.			
Interpret and use the terms: • *general formula* • *structural formula* • *displayed formula* • *skeletal formula* • *homologous series* (including definition) • *functional group* • *alkyl group* • *aliphatic* • *alicyclic* • *aromatic* • *saturated* • *unsaturated* • use of R to represent alkyl groups or fragments of organic compounds not involved in a reaction.			
Predict the formula of a molecule using the general formula of the homologous series.			
Define the term *structural isomer*.			
Predict possible structural isomers of an organic molecule from its molecular formula.			
Describe the two types of covalent bond fission: • *homolytic fission* • *heterolytic fission*			
Describe what a *radical* is and represent radicals with ‘dots’ in mechanisms.			
Describe what a ‘*curly arrow’* shows.			
Describe what a *reaction mechanism* is and use reaction mechanisms to explain what happens in organic reactions, including correct use of curly arrows and dipoles.			

Alkanes (Chapter 12) -			
Describe and explain: • what alkanes are • the bonding in alkanes in terms of orbital overlap • the shape and bond angles around each carbon atom.			
Explain how boiling points vary with carbon-chain length and branching, in terms of induced dipole-dipole interactions.			
Explain the low reactivity of alkanes.			
Compare and write equations for complete and incomplete combustion of alkanes.			
Explain the potential dangers from CO.			
Reaction of alkanes with chlorine: • name the mechanism • state what is needed for this reaction to take place • state the type of fission involved • name the three stages of the mechanism • use a series of equations to represent the reaction mechanism, using a single ‘dot’ to represent the unpaired electron.			
Explain why radical substitution is of limited use in organic synthesis.			
Alkenes (Chapter 13) -			
Describe and explain: • what alkenes are • the bonding in alkenes in terms of orbital overlap • the shape and bond angles around each carbon atom in the C=C.			
Explain the terms: • *stereoisomer* • *E/Z isomerism* • *cis-trans isomerism* • *electrophile*			
Identify molecules that will have E/Z or cis-trans stereoisomers from their structural formulae.			
Use Cahn-Ingold-Prelog (CIP) priority rules to identify E and Z stereoisomers.			
Explain the reactivity of alkenes.			

State what an addition reaction is and describe the reactions of alkenes with: - hydrogen in the presence of a suitable catalyst, e.g. Ni - halogens - hydrogen halides - steam in the presence of an acid catalyst, e.g. H_3PO_4.			
State the type of fission that takes place during electrophilic addition.			
Draw the mechanism for electrophilic addition reactions			
Use Markownikoff's rule to predict the major and minor organic products in addition reactions and explain these products in terms of the relative carbocation stability.			
Addition polymerisation: - describe the reaction - draw the repeat unit of an addition polymer from a given monomer - identify the monomer from a section of polymer.			
Describe the benefit for sustainability of processing waste polymers by: - combustion for energy production - use as an organic feedstock - removal of toxic waste products e.g. HCl..			
Discuss the benefits to the environment of the development of biodegradable and photodegradable polymers.			
Alcohols (Chapter 14) -			
Explain the solubility and relatively low volatility of alcohols compared with alkanes.			
Classify alcohols as primary, secondary or tertiary.			
Describe and write equations for the following reactions of alcohols: - combustion - oxidation by an oxidising agent, e.g. $Cr_2O_7^{2-}/H^+$ (i.e. $K_2Cr_2O_7/H_2SO_4$) - elimination of H_2O by heating with an acid catalyst (H_3PO_4 or H_2SO_4) - substitution with halide ions in the presence of acid (e.g. NaBr/ H_2SO_4)			
Explain how to control the oxidation products from primary alcohols using different reaction conditions.			
Compare the oxidation/resistance to oxidation of primary, secondary and tertiary alcohols.			

Haloalkanes (Chapter 15) -			
Describe how haloalkanes are hydrolysed in a substitution reaction by: • aqueous alkali • water in the presence of $AgNO_3$ and ethanol			
Describe how the rate of hydrolysis of different carbon-halogen bonds can be determined experimentally using the reaction with water in the presence of $AgNO_3$ and ethanol.			
Explain the trend in rates of hydrolysis in terms of bond enthalpies of the carbon-halogen bonds (C-F, C-Cl, C-Br and C-I).			
Define the term *nucleophile*.			
Draw the mechanism for the nucleophilic substitution of primary haloalkanes with aqueous alkali.			
Explain why the use of organohalogen compounds is an environmental concern in terms of: • production of halogen radicals from CFCs in the upper atmosphere • catalysed breakdown of ozone by $Cl^{\bullet}$ and other radicals e.g. $^{\bullet}NO$.			
Organic Synthesis (Chapter 16) -			
Describe the techniques and procedures used to purify organic liquids, including: • use of a separating funnel to remove an organic layer from an aqueous layer • drying with an anhydrous salt (e.g. $MgSO_4$, $CaCl_2$) • redistillation.			
Identify individual functional groups in molecules containing several functional groups.			
Predict the properties and reactions of molecules containing several functional groups.			
Devise two-stage synthetic routes for preparing an organic compound. These could involve transformations between all the functional groups studied, plus use of additional information provided.			

Analytical Techniques (Chapter 17) -			
Describe the effect of infrared (IR) radiation on covalent bonds, e.g. the gases containing C=O, O-H and C-H bonds in the atmosphere (CO_2, H_2O and CH_4), and the suspected link to global warming.			
Use IR spectra to identify: • an alcohol from an absorption peak of the O=H bond • an aldehyde or ketone from an absorption peak of the C=O bond • a carboxylic acid from an absorption peak of the C=O bond and a broad an absorption peak of the O-H bond.			
Interpret or predict IR spectra of familiar and unfamiliar substances.			
Describe how IR spectroscopy can be used to monitor gases causing air pollution (e.g. CO and NO from car emissions) and in modern breathalysers to measure ethanol in the breath.			
Identify the molecular ion peak from a mass spectrum and use it to determine molecular mass.			
Appreciate that mass spectra may contain a small M+1 peak due to the small proportion of carbon-13.			
Analyse fragmentation peaks in mass spectra and use them to identify parts of a structure.			
Deduce the structure of organic compounds using analytical data including: • elemental analysis (use to calculate empirical and molecular formulae) • mass spectra • IR spectra.			

Module 5 – Physical Chemistry and Transition Elements
Personal Learning Checklist (PLC)

Learning Objectives	**Dates**		
Rates (Chapters 18)			
Explain and use the terms: • *rate of reaction* • *order* • *overall order* • *rate constant* • *half-life* • *rate-determining step*			
Deduce: • the order of a reaction from experimental data • the rate equation from orders in the form rate = $k[A]^m[B]^n$, where m and n are 0, 1 or 2.			
Calculate the rate constant, k, and related quantities, from a rate equation including determination of units.			
Use a concentration-time graph to: • deduce the order (0 or 1) with respect to a reactant from the shape of the graph • calculate a reaction rate from the gradient.			
Determine the rate constant, k, from the constant half-life, $t_{1/2}$, using the relationship: $k = \ln 2/t_{1/2}$, for a first order reaction.			
Use a rate-concentration graph to: • deduce the order (0, 1 or 2) with respect to a reactant from the shape of the graph • determine the rate constant for a first order reaction from the gradient.			
Describe the techniques and procedures used to investigate reaction rates by the initial rates method and by continuous monitoring, including use of colorimetry.			
For a multi-step reaction, predict: • the rate equation (consistent with the rate determining step) • possible steps in a reaction mechanism from the rate equation and balanced equation for the overall reaction.			
Explain (qualitatively) the effect of temperature change on rate of reaction and hence the rate constant.			
Describe what the Arrhenius equation, $k = Ae^{-Ea/RT}$, shows (equation provided on the data sheet).			
Determine the activation energy (E_A) and pre-exponential factor (A) graphically using $\mathbf{\ln k = - E_A / RT + \ln A}$ (derived from the Arrhenius equation).			

Equilibria (Chapters 19)			
Use the terms: • *mole fraction* • *partial pressure*			
Calculate the quantities present at equilibrium from data.			
Describe the techniques and procedures used to determine quantities present at equilibrium.			
Give the expressions for K_c and K_p for homogeneous and heterogeneous equilibria.			
Calculate K_c and K_p , or related quantities, and determine the units.			
Describe the qualitative effect of changing temperature for exothermic and endothermic reactions on equilibrium constants.			
Describe the constancy of equilibrium constants with changes in concentration, pressure or in the presence of a catalyst.			
Explain how an equilibrium constant controls the position of equilibrium on changing concentration, pressure and temperature.			
Apply the above principles for K_c and K_p to other equilibrium constants.			
Acids, Bases and Buffers (Chapters 20 and 21)			
Define and use the terms: • Brønsted-Lowry acid • Brønsted-Lowry base • *conjugate acid-base pairs* • monobasic, dibasic and tribasic acids.			
Use ionic equations to show the role of H^+ in the reactions of acids with metals and bases (including carbonates, metal oxides and alkalis).			
Use the acid dissociation constant, K_a, as a measure of the extent of acid dissociation.			
Describe the relationship between K_a and pK_a.			
Use the expressions: $\mathbf{pH = -log[H^+]}$ and $\mathbf{[H^+] = 10^{-pH}}$			
Use the expression for the ionic product of water, K_W.			
Calculate pH, or related quantities, for: • strong monobasic acids • strong bases, using K_W.			
Calculate the pH, K_a or related quantities for a weak monobasic acid using approximations.			
Describe the limitations of using approximations to K_a related calculations for 'stronger' weak acids.			

Describe what a buffer solution is and how buffer solutions are formed from: - a weak acid and a salt of the weak acid, e.g. CH_3COOH / CH_3COONa - excess of a weak acid and a strong alkali, e.g. excess CH_3COOH / NaOH.			
Explain the role of the conjugate acid-base pair in an acid buffer solution, e.g. CH_3COOH / CH_3COO^-, in the control of pH.			
Calculate the pH of a buffer solution from the K_a value of a weak acid and the equilibrium concentrations of the conjugate acid-base pair.			
Explain the control of blood pH by the carbonic acid-hydrogencarbonate buffer system.			
pH titration curves for strong and weak acids with strong and weak bases: - sketch and interpret the shapes - explain the choice of an indicator - explain the indicator colour changes in terms of equilibrium shift between HA and A^- forms of the indicator.			
Describe the techniques and procedures used when measuring pH with a pH meter.			
Energy (Chapters 22)			
Explain the term *lattice enthalpy* and use as a measure of the strength of ionic bonding in a giant ionic lattice.			
Use the lattice enthalpy of a simple ionic solid (e.g. NaCl, $MgCl_2$) and relevant energy terms in calculations and to construct Born-Haber cycles.			
Explain and use the following terms: - *enthalpy change of solution* ($\Delta_{sol}H$) - *enthalpy change of hydration* ($\Delta_{hyd}H$).			
Use the enthalpy change of solution for a simple ionic solid (e.g. NaCl, $MgCl_2$) and relevant energy terms (*enthalpy change of hydration* and *lattice enthalpy*) in calculations and to construct enthalpy cycles.			
Explain (qualitatively) the effect of ionic charge and ionic radius on the exothermic value of a lattice enthalpy and enthalpy change of hydration.			
Explain what *entropy* is.			
Explain the difference in magnitude of the entropy of a system: - of solids, liquids and gases - for a reaction in which there is a change in the number of gaseous molecles.			
Calculate the entropy change of a system, ΔS, and related quantities for a reaction given the entropies of the reactants and products.			
Explain how the feasibility of a process depends on the entropy change and the temperature change in the system, $T\Delta S$, and the enthalpy change of the system, ΔH.			
Explain (and carry out related calculations) the free energy change, ΔG, as: $\boldsymbol{\Delta G = \Delta H - T\Delta S}$, (the Gibbs' equation) and that a process is feasible when ΔG has a negative value.			
Explain the limitations of predictions made by ΔG about feasibility, in terms of kinetics.			

Redox and Electrode Potentials (Chapters 23)			
Explain and use the terms: • *oxidising agent* • *reducing agent*			
Construct redox equations using half-equations and oxidation numbers.			
Interpret and predict reactions involving electron transfer.			
Describe the techniques and procedures used when carrying out redox titrations including those involving Fe^{2+}/MnO_4^- and $I_2/S_2O_3^{2-}$.			
Use experimental results of to carry out calculations for redox titrations involving: • Fe^{2+}/MnO_4^- and $I_2/S_2O_3^{2-}$ • non-familiar redox systems.			
Use the term *standard electrode (redox) potential, $E^{\ominus}$,* including its measurement using a hydrogen electrode.			
Describe the techniques and procedures used to measure cell potentials of: • metals or non-metals in contact with their ions in aqueous solution • ions of the same element in different oxidation states in contact with a Pt electrode.			
Calculate the standard cell potential by combining two standard electrode potentials.			
Predict the feasibility of a reaction using standard cell potentials and the limitations of such predictions in terms of kinetics and concentration.			
Describe how the principles of electrode potentials are applied to modern storage cells.			
Explain that a fuel cell uses the energy from the reaction of a fuel with oxygen to create a voltage and the changes that take place at each electrode.			

Transition Elements (Chapters 24)			
Deduce the electron configuration of atoms and ions of the d-block elements of Period 4 (Sc-Zn).			
Describe the elements Ti-Cu as transition elements, i.e. d-block elements tht have an ion with an incomplete d-sub-shell.			
Illustrate, using at least two transition elements, the: • existence of more than one oxidation state for each element in its compounds • formation of coloured ions • catalytic behaviour of the elements and their compounds and their importance in the manufacture of chemicals by industry.			
Explain and use the term *ligand* in terms of coordinate (dative covalent) bonding to a metal ion or metal, including the examples: • H_2O, Cl^- and NH_3 (monodentate) • $NH_2CH_2CH_2NH_2$ (‘en’) (bidentate) Other examples could be given in the exams.			
Use the terms *complex ion* and *coordination number* and examples of complexes with: • six-fold coordination with an octahedral shape • four-fold coordination with either a planar or tetrahedral shape.			
Describe the types of stereoisomerism shown by complexes, including those associated with bidentate and multidentate ligands: • *cis-trans* isomerism e.g. $Pt(NH_3)_2Cl_2$ • optical isomerism e.g.$[Ni(NH_2CH_2CH_2NH_2)_3]^{2+}$			
Describe the use of cis-platin as an anti-cancer drug and its action by binding to DNA preventing cell division.			
Ligand substitution reactions and the associated colour changes in the formation of: • $[Cu(NH_3)_4(H_2O)_2]^{2+}$ and $[CuCl_4]^{2-}$ from $[Cu(H_2O)_6]^{2+}$ • $[Cr(NH_3)_6]^{3+}$ from $[Cr(H_2O)_6]^{3+}$			
Explain the biochemical importance of iron in haemoglobin, including ligand substitution involving O_2 and CO.			
Write ionic equations and state the colour changes of the reactions of aqueous: • Cu^{2+} • Fe^{2+} • Fe^{3+} • Mn^{2+} • Cr^{3+} with aqueous: • sodium hydroxide • ammonia including: • precipitation reactions • complex formation with excess aqueous sodium hydroxide or ammonia.			

Redox reactions and colour changes for: • interconversions between Fe^{2+} and Fe^{3+} • interconversions between Cr^{3+} and $Cr_2O_7^{2-}$ • reduction of Cu^{2+} to Cu^+ and disproportionation of Cu^+ to Cu^{2+} and Cu You do not need to recall equations but may be required to construct and interpret redox equations using relevant half-equations and oxidation numbers.			
Interpret and predict unfamiliar reactions including ligand substitution, precipitation and redox.			
Describe the processes and techniques needed to identify the following ions in an unknown compound: • anions: CO_3^{2-}, Cl^-, Br^-, I^-, SO_4^{2-}, • cations: NH_4^+, Cu^{2+}, Fe^{2+}, Fe^{3+}, Mn^{2+}, Cr^{3+}.			

Module 6 – Organic Chemistry and Analysis - Personal Learning Checklist (PLC)

Learning Objectives	**Dates**		
Aromatic Compounds (Chapter 25)			
Compare the Kekulé and delocalised models of benzene.			
Describe the experimental evidence for a delocalised, rather than Kekulé model, for benzene (bond lengths, enthalpy change of hydrogenation, resistance to reaction).			
Name substituted aromatic compounds using the IUPAC system.			
Give the products and write equations for electrophilic substitution of aromatic compounds with: • concentrated nitric acid in the presence of concentrated sulfuric acid • a halogen in the presence of a halogen carrier (iron, iron halide or aluminium halide) • a haloalkane or acyl chloride in the presence of a halogen carrier (Friedel-Crafts reaction)			
Explain why Fridel-Crafts reactions are important in organic synthesis.			
Show the mechanism of electrophilic substitution in arenes for nitration and halogenation, including equations to show the formation of the electrophiles and re-formation of the catalysts.			
Explain why benzene is relatively resistant to bromination, compared to alkenes, in terms of delocalised electron density of the π-system in benzene compared with the localised electron density of the π-bond in alkenes.			
Interpret unfamiliar electrophilic substitution reactions of aromatic compounds, including predicting mechanisms.			
Describe how phenols are weak acids, in terms of their neutralisation reaction with NaOH but not with carbonates.			
Give the products and write equations for electrophilic substitution of phenols with: • bromine to form 2,4,6-tribromophenol • dilute nitric acid to form a mixture of 2-nitrophenol and 4-nitrophenol.			
Explain why phenols undergo electrophilic substitution reactions more readily than benzene, in terms of electron pair donation to the π-system from an oxygen p-orbital in phenol (e.g. reactions with bromine and with nitric acid). Explain in terms of susceptibility of ring to ‘attack’, **not** stability of the intermediate.			

State the effect of the following groups in electrophilic substitution of aromatic compounds and explain why this is important in organic synthesis: • 2- and 4-directing electron-donating groups (OH, NH_2) • 3-directing electron-withdrawing groups (NO_2).			
Predict substitution products based on groups already on the benzene ring. You should learn the effect of OH, NH_2 and NO_2, but may be given and need to use information on other groups.			
Carbonyl Compounds (Chapter 26)			
Give the products and write equations for the following reactions: • oxidation of aldehydes using $Cr_2O_7^{2-}/H^+$ (i.e. acidified potassium dichromate (VI), $K_2Cr_2O_7/H_2SO_4$) • nucleophilic addition of carbonyl compounds with $NaBH_4$ and HCN (i.e. $NaCN_{(aq)}/H^+_{(aq)}$).			
Show mechanisms for nucleophilic addition reactions of aldehydes and ketones with: • $NaBH_4$ (nucleophile hydride ion, H^- and protonation from H_2O) • HCN (nucleophile cyanide ion, CN^- and protonation from H_2O or H^+)			
Describe how to test for carbonyl compounds using 2,4-dinitrophenylhydrazine and measuring the melting point of the derivative. (Equation for reaction and structure of derivative **not** required)			
Describe how distinguish between aldehydes and ketones using Tollens' reagent (ammoniacal silver nitrate) and explain the observations with an aldehyde in terms of oxidation and reduction.			
Explain why lower carboxylic acids are water soluble.			
Give the products and write equations for the reactions of carboxylic acids with: • metals • bases (including carbonates, metal oxides and alkalis) • alcohols in the presence of an acid catalyst (e.g. H_2SO_4).			
Give the products and write equations for the following reactions: • acid anhydrides with alcohols • acyl chlorides with alcohols • acyl chlorides with water • acyl chlorides with ammonia • acyl chlorides with amines • hydrolysis of esters in hot aqueous acid • hydrolysis of esters in hot aqueous alkali.			
Write equations for the formation of acyl chlorides from carboxylic acids using $SOCl_2$.			

Amines, Amino Acids and Polymers (Chapter 27)			
Describe the basicity of amines in terms of proton acceptance by the nitrogen lone pair.			
State products and write equations for the reactions of amines with dilute inorganic acids, e.g. HCl.			
Describe how to prepare: • aliphatic amines by substitution of haloalkanes with excess ethanolic ammonia and amines • aromatic amines by reduction of nitroarenes using tin and concentrated hydrochloric acid.			
Give the general formula and structure of an amino acid.			
State products and write equations for the reactions of amino acids with alkalis and acids.			
Draw the structures of primary and secondary amides.			
Describe what optical isomerism is, in terms of non-superimposable mirror images about a chiral centre.			
Draw 3D diagrams to show stereoisomers.			
Identify chiral centres in given molecules			
Describe condensation polymerisation to form: • polyesters • polyamides (from carboxylic acids/dicarboxylic acids or respective acyl chlorides and alcohols/diols or amines/diamines).			
Describe the acid and base hydrolysis of: • the ester groups in polyesters • the amide groups in polyamides.			
For addition and condensation polymerisation, predict: • the repeat unit from a given monomer(s) • the monomer required from a given section of the polymer structure • the type of polymerisation.			

Organic Synthesis (Chapter 28 and 16)			
Describe the techniques and procedures used to prepare organic solids, including distillation and reflux.			
Describe the techniques and procedures used to purify organic liquids, including: • use of a separating funnel to remove an organic layer from an aqueous layer • drying with an anhydrous salt (e.g. $MgSO_4$, $CaCl_2$) • redistillation.			
Describe the techniques and procedures used to purify organic solids, including: • filtration under reduced pressure • recrystallisation • measurement of melting points.			
Identify individual functional groups in molecules containing several functional groups.			
Predict the properties and reactions of molecules containing several functional groups.			
Devise multi-stage synthetic routes for preparing an organic compound. These could involve transformations between all the functional groups studied, plus use of additional information provided.			
Describe how to increase the carbon chain length by formation of C-C≡N via the reactions of: • haloalkanes with CN^- and ethanol • carbonyl compounds with HCN (including mechanisms). Subsequent reactions of the nitriles produced: • reduction with H_2/Ni to form amines • acid hydrolysis to form carboxylic acids			
Describe how to form a substituted aromatic C-C by Fridel-Crafts reactions: alkylation (using a haloalkane) or acylation (using an acyl chloride) in the presence of a halogen carrier.			

Spectroscopy (Chapter 29 and 17)			
Describe the effect of infrared (IR) radiation on covalent bonds, e.g. the gases containing C=O, O-H and C-H bonds in the atmosphere (CO_2, H_2O and CH_4), and the suspected link to global warming.			
Use IR spectra to identify: • an alcohol from an absorption peak of the O=H bond • an aldehyde or ketone from an absorption peak of the C=O bond • a carboxylic acid from an absorption peak of the C=O bond and a broad an absorption peak of the O-H bond.			
Interpret or predict IR spectra of familiar and unfamiliar substances.			
Describe how IR spectroscopy can be used to monitor gases causing air pollution (e.g. CO and NO from car emissions) and in modern breathalysers to measure ethanol in the breath.			
Identify the molecular ion peak from a mass spectrum and use it to determine molecular mass.			
Appreciate that mass spectra may contain a small M+1 peak due to the small proportion of carbon-13.			
Analyse fragmentation peaks in mass spectra and use them to identify parts of a structure.			
Analyse carbon-13 NMR spectra to predict: • the number of carbon environments in a molecule • the different types of carbon environment present from chemical shift values • possible structures for the molecule.			
Analyse proton NMR spectra to predict: • the number of proton environments in a molecule • the different types of proton environment present from chemical shift values • the relative numbers of each type of proton present from relative peak areas, using integration traces or ratio numbers • the number of non-equivalent protons adjacent to a given proton from the spin-spin splitting pattern, using the n+1 rule • possible structures for the molecule.			
Predict carbon-13 and proton NMR spectra for a given molecule.			
State that tetramethylsilane, TMS, is used as the standard for chemical shift measurements.			
Explain why deuterated solvents, e.g. $CDCl_3$, are needed when running NMR spectra.			
Identify O-H and N-H protons by proton exchange using D_2O.			
Deduce the structure of organic compounds using analytical data including: • elemental analysis (use to calculate empirical and molecular formulae) • mass spectra • IR spectra • NMR spectra.			

Chromatography and Analytical Techniques (Chapter 29 and 17)			
Interpret one-way TLC chromatograms using R_f values.			
Interpret gas chromatograms using: • retention times • the amount and proportions of the components in a mixture (including making and using external calibration curves to confirm concentrations of components).			
Describe how to test for: • alkenes using bromine water • haloalkanes by reaction with aqueous silver nitrate in ethanol • phenols by weak acidity but no reaction with carbonates • carbonyl compounds by reaction with 2,4-DNP • aldehydes by reaction with Tollens' reagent • primary and secondary alcohols and aldehydes by reaction with acidified dichromate • carboxylic acids by reaction with carbonates.			