**Module 1 – Development of practical skills - Personal Learning Checklist**

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| **Learning Objectives:** | **Confidence** | | |
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| ***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**

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| **Learning Objectives:** | **Confidence** | | |
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| ***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 12C 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*, Mr, (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: NO3–, CO32-, SO42–, OH–, NH4+, Zn2+, 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 *Mr*. |  |  |  |
| Calculate the formula of a hydrated salt from data giving composition by mass or % by mass. |  |  |  |
| ***Carry out calculations using:*** | | | |
| * the Avagadro constant |  |  |  |
| * mass of substance, *Mr* 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: HC*l*, H2SO4, HNO3 and CH3COOH. |  |  |  |
| Recall the names and formulae of the following common alkalis: NaOH, KOH and NH3. |  |  |  |
| 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 H2O, 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. |  |  |  |

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| ***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 -*** | | | |
| Describe what ionic bonding 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. |  |  |  |

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| ***Shapes of Molecules and Intermolecular Forces*** |  |  |  |
| ***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. CH4, NH3 and H2O. |  |  |  |
| ***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. H2O) but others do not (e.g. CO2) . |  |  |  |
| 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 H2O, 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. I2, 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**

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| **Learning Objectives:** | **Confidence** | | |
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| ***Periodicity (Chapter 7) -*** |  |  |  |
| ***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) -*** |  |  |  |
| ***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 CaCO3 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:   * CO32- * SO42- * Cl- , Br-, I- * NH4+ |  |  |  |

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| ***Enthalpy Changes (Chapter 9) -*** |  |  |  |
| 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 (ΔrH)* |  |  |  |
| * *enthalpy change of formation (ΔfH)* |  |  |  |
| * *enthalpy change of combustion (ΔcH)* |  |  |  |
| * *enthalpy change of neutralisation (ΔneutH)* |  |  |  |
| * *average bond enthalpy* |  |  |  |
| Determine the enthalpy changes directly from experimental results, including the use of **q = mcΔT**. |  |  |  |
| 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 molecuels 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, *Kc*, for homogeneous reactions. |  |  |  |
| Calculate the equilibrium constant, *Kc*, from provided equilibrium concentrations. (Do not need to determine units for *Kc*.) |  |  |  |
| Estimate the position of equilibrium from the magnitude of *Kc*. (qualitative only). |  |  |  |

**Module 4 – Core Organic Chemistry - Personal Learning Checklist**

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| **Learning Objectives:** | **Confidence** | | |
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| ***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. |  |  |  |

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| ***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. H3PO4. |  |  |  |
| 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. Cr2O72-/H+ (i.e. K2Cr2O7/H2SO4) * elimination of H2O by heating with an acid catalyst (H3PO4 or H2SO4) * substitution with halide ions in the presence of acid (e.g. NaBr/ H2SO4) |  |  |  |
| 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. |  |  |  |

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| ***Haloalkanes (Chapter 15) -*** | | | |
| Describe how haloalkanes are hydrolysed in a substitution reaction by:   * aqueous alkali * water in the presence of AgNO3 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 AgNO3 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**·** and other radicals e.g. **·**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. MgSO4, CaCl2) * 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. |  |  |  |

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| ***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 (CO2, H2O and CH4), 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. |  |  |  |