Personalised Learning Checklist

A2 Unit 5: Physical Chemistry and Transition Elements

Module 5 – Physical Chemistry and Transition Elements			
5.1.1 How Fast?	☺	😐	☹
Explain and use the terms: rates of reaction, order, overall order, rate constant, half-life, rate-determining step			
Deduction of: - Orders from experimental date - A rate equation from orders in the form:rate = $K[A]^m[B]^n$, where m and n are 0, 1 or 2			
Calculation of the rate constant, k, and related quantities, from a rate equation including determination of units			
From concentration-time graphs: - Deduction of the order(0 or 1) with respect to a reactant from the shape of the graph - Calculation of reaction rates from the measurement of gradients			
From a concentration – time graph of a 1st order reaction, measurement of a constant half-life, $t_{1/2}$			
For a 1st order reaction, determination of the rate constant, k, from the constant half-life using the relationship: $k=\ln 2/t_{1/2}$			
From a rate-concentration graph: - Deduce the order (0, 1 or 2) with respect to a reactant from the shape of the graph - Determination of rate constant for a first order reaction from the gradient			
Techniques and proceedures used to investigate reaction rates by the initial rates method and by continuous monitoring, including the use of colorimetry			
For a multi-step reaction, prediction of: - A rate equation which is consistent with the rate determining step - Possible steps in a reation mechanism from the rate equation and the balanced equation for the overall reaction			
Qualitative explanation of the effect of temperature change on the rate of reaction and hence the rate constant			
The Arrhenius equation: - The exponential relationship between the rate constant, k and temperature, T given by the Arrhenius equation, $k = Ae^{-Ea/RT}$ - Determination of E_a and A graphically using: $\ln k = -E_a/RT + \ln A$ derived from the Arrhenius equation			

5.1.2 How Far?	☺	😐	☹
Use of the terms mole fraction and partial pressure			
Calculation of the quantities present at equilibrium, given appropriate data			
The techniques used to determine quantities present at equilibrium			
Expressions for K_c and K_p for homogeneous and heterogeneous equilibria			
Calculations of K_c and K_p , or related quantities, including determination of units			
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			
Explain how an equilibrium constant controls the position of equilibrium on changing concentration, pressure and temperature			
5.1.3 Acids, bases and buffers	☺	😐	☹
Describe a Bronsted - Lowry acid as a species that can donate a proton and a Bronsted - Lowry base as a species that can accept a proton.			
Use the term conjugate acid – base pairs			
Identify monobasic, dibasic and tribasic acids			
The role of H^+ ions in the reactions of acids with metals and bases (including carbonates, metal oxides and alkalis), using ionic equations			
Use the acid dissociation constant, K_a, to ascertain the extent of acid dissociation			
Understand the relationship between K_a and pK_a			
Use the expression for pH as: - $pH = -\log[H^+]$ - $[H^+] = 10^{-pH}$			
Use the expression for the ionic product of water, K_w			
Calculations of pH, K_a or related quantities, for: - Strong monobasic acids - Strong bases, using K_w			
Calculations of pH, K_a or related quantities, for a weak monobasic acid using approximations			
Limitaions of using approximations to K_a related calculations for “stronger” weak acids			

	☺	😐	☹
Describe a buffer solution as a system that minimises pH changes on addition of small amounts of an acid or a base			
State that a buffer solution can be made from a weak acid and a salt of the weak acid, e.g. CH_3COOH/CH_3COONa.			
State that a buffer solution can be made from an excess of a weak acid and a strong alkali, e.g. $CH_3COOH/NaOH$			
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 role of carbonic acid–hydrogencarbonate as a buffer in the control of blood pH.			
For acid–base titration pH curves for strong and weak acids and bases: (i) interpret, or sketch, their shapes (ii) explain the choice of suitable indicators for acid–base titrations, given the pH range of the indicator. (iii) Explain indicator colour change in terms of equilibrium shift between the HA and A^- forms of the indicator			
Describe the procedures used when measuring pH with a pH meter			

5.2.1 Lattice Enthalpy	☺	😐	☹
Explain and use the term *lattice enthalpy* (ΔH negative, i.e. gaseous ions to the solid lattice) as a measure of ionic bond strength.			
Use the lattice enthalpy of a simple ionic solid (i.e. NaCl, $MgCl_2$) and relevant energy terms to: (i) construct Born–Haber cycles (ii) carry out related calculations.			
Explain and use the terms *enthalpy change of solution* and *enthalpy change of hydration.*			
Use the enthalpy change of solution of a simple ionic solid (i.e. NaCl, $MgCl_2$) and relevant energy terms (*enthalpy change of hydration,* and *lattice enthalpy*), to: (i) construct Born–Haber cycles (ii) carry out related calculations.			
Explain, in qualitative terms, the effect of ionic charge and ionic radius on the exothermic value of a lattice enthalpy and enthalpy change of hydration.			
5.2.2 Enthalpy and Entropy	☺	😐	☹
Explain the difference in magnitude of entropy: (i) of a solid and a gas (ii) for a reaction in which there is a change in the number of gaseous molecules.			
Calculate the entropy change for a reaction given the entropies of the reactants and products.			
Explain that the tendency of a process to take place depends on absolute temperature, T, the entropy change in the system, ΔS, and the enthalpy change, ΔH, with the surroundings.			
Explain that the balance between entropy and enthalpy changes is the *free energy change*, ΔG, which determines the feasibility of a reaction.			
State and use the relationship $\Delta G = \Delta H - T\Delta S$.			
Explain the limitations of predictions made by ΔG about feasibility, in terms of kinetics			

5.2.3 Redox and electrode potentials	☺	😐	☹
Explain, for simple redox reactions, the terms *redox, oxidation number, half-reaction, oxidising agent* and *reducing agent*.			
Construct redox equations using relevant half-equations or oxidation numbers.			
Interpret and make predictions for reactions involving electron transfer.			
The techniques and procedures used when carrying out redox titrations including those involving Fe^{2+}/MnO_4^- and $I_2/S_2O_3^{2-}$			
Structured and non-structured titration calculations, based on experimental results of redox tirations involving: - Fe^{2+}/MnO_4^- and $I_2/S_2O_3^{2-}$ - Non-familiar redox systems			
Define the term *standard electrode (redox) potential*, $E^{\ominus}$.			
Describe how to measure, using a hydrogen electrode, standard electrode potentials of: (i) metals or non-metals in contact with their ions in aqueous solution (ii) ions of the same element in different oxidation states.			
Calculate a standard cell potential by combining two standard electrode potentials.			
Predict, using standard cell potentials, the feasibility of a reaction. Describe the limitations of such predictions in terms of kinetics and concentration			
Apply principles of electrode potentials to modern storage cells.			
Explain that a fuel cell uses the energy from the reaction of a fuel with oxygen to create a voltage.			
Explain the changes that take place at each electrode in a hydrogen–oxygen fuel cell.			

5.3.1 Transition Elements	☺	😐	☹
Deduce the electron configurations of atoms and ions of the d-block elements of Period 4 (Sc–Zn), given the atomic number and charge.			
Describe the elements Ti–Cu as *transition elements*, i.e. d-block elements that have an ion with an incomplete d sub-shell.			
Illustrate, using at least 2 transition metals: (i) the existence of more than one oxidation state for each element in its compounds (ii) the formation of coloured ions (iii) the catalytic behaviour of the elements and/or their compounds;			
Explain the term *ligand* in terms of coordinate bonding.			
State and use the terms *complex ion* and *coordination number*.			
Explain and use the term *bidentate ligand* (e.g. $NH_2CH_2CH_2NH_2$, 'en').			
State and give examples of complexes with: - Six-fold coordination with an octahedral shape. - Four-fold coordination with either planar or tetrahedral shape			
Describe the types of stereoisomerism shown by complexes, including those associated with bidentate and multidentate ligands: (i) *cis-trans* isomerism, e.g. $Pt(NH_3)_2Cl_2$ (ii) optical isomerism, e.g. $[Ni(NH_2CH_2CH_2NH_2)_3]^{2+}$			
Describe the use of *cis*-platin as an anticancer drug and its action by binding to DNA in cancer cells, preventing division.			
Describe the process of ligand substitution and the accompanying colour changes in the formation of: (i) $[Cu(NH_3)_4(H_2O)_2]^{2+}$ and $[CuCl_4]^{2-}$ from $[Cu(H_2O)_6]^{2+}$ (ii) $[Cr(NH_3)_6]^{3+}$ from $[Cr(H_2O)_6]^{3+}$.			
Explain the biochemical importance of iron in haemoglobin, including ligand substitution involving O_2 and CO.			
Describe, including ionic equations, the reactions and the accompanying colour changes of $Cu^{2+}(aq)$, $Fe^{2+}(aq)$, $Fe^{3+}(aq)$, $Mn^{2+}(aq)$ and $Cr^{3+}(aq)$, with aqueous sodium hydroxide and aqueous ammonia, including: - Precipitation reactions - Complex formation with excess aqueous sodium hydroxide and aqueous ammonia			

Describe the redox reactions and accompanying 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			
Interpret and predict unfamiliar reactions including ligand substitution , precipitation and redox			
5.3.2 Qualitative analysis	☺	😐	☹
Describe the qualitative analysis on a test tube scale:process 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+}			

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Module 6: Organic Chemistry and Analysis

6.1.1 Aromatic Compounds	☺	😐	☹
Compare the Kekulé and delocalised models for benzene in terms of p-orbital overlap forming a delocalised π system.			
Review the evidence for a delocalised model of benzene in terms of bond lengths, enthalpy change of hydrogenation and resistance to reaction.			
Use of IUPAC rules of nomenclature for systematically naming substituted aromatic compounds			
Describe the electrophilic substitution of aromatic compounds with: (i) concentrated nitric acid in the presence of concentrated sulfuric acid (ii) a halogen in the presence of a halogen carrier. (iii) a haloalkane or acyl chloride in the presence of a halogen carrier (Friedel-Crafts reaction) and its importance to synthesis by formation of a C-C bond to an aromatic ring			
Outline the mechanism of electrophilic substitution in arenes for nitration and halogenation			
Explain the relative resistance to bromination of benzene, compared with alkenes, in terms of the 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 prediction of mechanisms			
Describe the weak acidity of phenols shown by the neutralisation reaction with NaOH but absence of a reaction with carbonates			
Describe the electrophilic substitution reactions of phenol: - With bromine to form 2,4,6 tribromophenol - With dilute nitric acid to form 2-nitrophenol			
Explain the relative ease of bromination of phenol compared with benzene, in terms of electron-pair donation to the π system from an oxygen *p* orbital in phenol.			
Recall the 2- and 4-directing effect of electron donating groups (OH, NH_2) and the 3-directing effect of electron-withdrawing groups (NO_2) in electrophilic substitution of aromatic compounds			
The prediction of substitution products of aromatic compounds by directing effects and the importance to organic synthesis			

6.1.2 Carbonyl Compounds	☺	😐	☹
Describe the oxidation of aldehydes using $Cr_2O_7^{2-}/H^+$ to form carboxylic acids.			
Nucleophilic addition reactions of carbonyl compounds with: - $NaBH_4$ to form alcohols - HCN [i.e. $NaCN_{(aq)}/H^+_{(aq)}$], to form hydroxynitriles			
Outline the mechanism for nucleophilic addition reactions of aldehydes and ketones with $NaBH_4$ and HCN			
Describe the use of 2,4-dinitrophenylhydrazine to: (i) detect the presence of a carbonyl group in an organic compound (ii) identify a carbonyl compound from the melting point of the derivative.			
Describe the use of Tollens' reagent (ammoniacal silver nitrate) to: (i) detect the presence of an aldehyde group (ii) distinguish between aldehydes and ketones, explained in terms of the oxidation of aldehydes to carboxylic acids with reduction of silver ions to silver.			
6.1.3 Carboxylic Acids and Esters	☺	😐	☹
Explain the water solubility of carboxylic acids in terms of hydrogen bonding			
Describe the reactions of carboxylic acids in aqueous conditions with metals and bases (including carbonates, metal oxides and alkalis)			
Describe esterification of: - carboxylic acids with alcohols, in the presence of an acid catalyst (e.g. concentrated H_2SO_4) - of acid anhydrides with alcohols.			
Describe the hydrolysis of esters: (i) in hot aqueous acid to form carboxylic acids and alcohols (ii) in hot aqueous alkali to form carboxylate salts and alcohols.			
Describe the formation of acyl chlorides from carboxylic acids using $SOCl_2$			
Recall the use of acyl chlorides in synthesis in formation of esters, carboxylic acids and primary and secondary amides			

6.2.1 Amines	☺	😐	☹
Explain the basicity of amines in terms of proton acceptance by the nitrogen lone pair.			
Describe the reactions of amines with dilute acids, e.g. $HCl_{(aq)}$, to form salts.			
Describe the preparation of: (i) aliphatic amines by substitution of haloalkanes with excess ethanolic ammonia (ii) aromatic amines by reduction of nitroarenes using tin and concentrated hydrochloric acid.			
6.2.2 Amino acids, amides and chirality	☺	😐	☹
State the general formula for an α-amino acid as $RCH(NH_2)COOH$.			
Recall the following reactions of an α-amino acid: - Reaction of the carboxylic acid group with alkalis and in the formation of esters - Reaction of the amine group with acids			
Amides	☺	😐	☹
Recall the structures of primary and secondary amides			
Describe *optical isomers* as non-superimposable mirror images about an organic chiral centre: four different groups attached to a carbon atom.			
Identify chiral centres in a molecule of given structural formula.			
Polyesters and polyamides	☺	😐	☹
Describe *condensation polymerisation* to form: (i) polyesters, (ii) polyamides			
Describe the acid and bas hydrolysis of: - The ester groups in polyesters - The amide groups in polyamides			
Prediction from addition and condensation polymerisation of: - The repeat unit from a given monomer - The monomer required for a given section of a polymer molecule - The type of polymerisation			

6.2.4 Carbon-carbon bond formation	☺	😐	☹
The use of C–C bond formation in synthesis to increase the length of a carbon chain			
formation of C–C≡N by reaction of: (i) haloalkanes with CN^- and ethanol, including nucleophilic substitution mechanism (ii) carbonyl compounds with HCN, including nucleophilic addition mechanism			
Reaction of nitriles from **(from above):** (i) by reduction (e.g. with H_2/Ni) to form amines (ii) by acid hydrolysis to form carboxylic acids			
formation of a substituted aromatic C–C by alkylation (using a haloalkane) and acylation (using an acyl chloride) in the presence of a halogen carrier (Friedel–Crafts reaction)			
6.2.5 Organic Synthesis	☺	😐	☹
The techniques and procedures used for the preparation and purification of organic solids involving use of a range of techniques including: (i) organic preparation: - use of Quickfit apparatus - distillation and heating under reflux (ii) purification of an organic solid: - filtration under reduced pressure - recrystallisation - measurement of melting points			
Synthetic routes for an organic molecule containing several functional groups: (i) identification of individual functional groups (ii) prediction of properties and reactions			
Multi-stage synthetic routes for preparing organic compounds			

6.3 Analysis			
6.3.1 Chromatography and Qualitative Analysis	☺	😐	☹
Interpretation of one-way TLC chromatograms in terms of R_f values			
Interpretation of gas chromatograms in terms of: (i) retention times (ii) the amounts and proportions of the components in a mixture			
Qualitative analysis of organic functional groups on a test-tube scale; processes and techniques needed to identify the following functional groups in an unknown compound: (i) alkenes by reaction with bromine (ii) haloalkanes by reaction with aqueous silver nitrate in ethanol (iii) phenols by weak acidity but no reaction with CO_3^{2-} (iv) carbonyl compounds by reaction with 2,4-DNP (v) aldehydes by reaction with Tollens' reagent (vi) primary and secondary alcohols and aldehydes by reaction with acidified dichromate (vii) carboxylic acids by reaction with CO_3^{2-}			
6.3.2 Spectroscopy	☺	😐	☹
Analysis of a carbon-13 NMR spectrum of an organic molecule to make predictions about: (i) the number of carbon environments in the molecule (ii) the different types of carbon environment present, from chemical shift values (iii) possible structures for the molecule			
Analysis of a high resolution proton NMR spectrum of an organic molecule to make predictions about: (i) the number of proton environments in the molecule (ii) the different types of proton environment present, from chemical shift values (iii) the relative numbers of each type of proton present from relative peak areas, using integration traces or ratio numbers, when required (iv) the number of non-equivalent protons adjacent to a given proton from the spin–spin splitting pattern, using the $n + 1$ rule (v) possible structures for the molecule			
Prediction of a carbon-13 or proton NMR spectrum for a given molecule			
The use of tetramethylsilane, TMS, as the standard for chemical shift measurements (ii) the need for deuterated solvents, e.g. $CDCl_3$, when running an NMR spectrum (iii) the identification of O–H and N–H protons by proton exchange using D_2O			
Deduction of the structures of organic compounds from different analytical data including: (i) elemental analysis (ii) mass spectra (iii) IR spectra (iv) NMR spectra			

A Level in Chemistry A (H432)

Module 2- Foundations in Chemistry			
2.1 Atoms and Reactions			
Describe protons, neutrons and electrons in terms of relative charge and relative mass.			
Describe the distribution of mass and charge within an atom.			
Describe the contribution of protons and neutrons to the nucleus of an atom, in terms of atomic (proton) number and mass (nucleon) number.			
Deduce the numbers of protons, neutrons and electrons in: (i) an atom given its atomic and mass number (ii) an ion given its atomic number, mass number and ionic charge.			
Explain the term *isotopes* as atoms of an element with different numbers of neutrons and different masses.			
State that ^{12}C is used as the standard measurement of relative masses.			
Define the terms *relative isotopic mass* and *relative atomic mass*, based on the ^{12}C scale.			
Describe the use of mass spectrometry to determine relative isotopic masses and relative abundances of the isotope			
Calculate the relative atomic mass of an element given the relative abundances of its isotopes.			
Use the terms *relative molecular mass* and *relative formula mass* and calculate values from relative atomic masses.			
Explain the terms: (i) *amount of substance* (ii) *mole* (symbol 'mol') as the unit for amount of substance (iii) the *Avogadro constant*, N_A, as the number of entities per mole (6.02×10^{23} mol^{-1}).			
Define and use the term *molar mass* (units g mol^{-1}) as the mass per mole of a substance.			
Explain the terms: (i) *empirical formula* as the simplest whole number ratio of atoms of each element present in a compound (ii) *molecular formula* as the actual number of atoms of each element in a molecule.			

Calculate empirical and molecular formulae using composition by mass and percentage compositions.			
Construct balanced chemical equations for reactions studied and for unfamiliar reactions given reactants and products.			
Carry out calculations, using amount of substance in mol, involving: (i) mass (ii) gas volume (iii) solution volume and concentration.			
Deduce stoichiometric relationships from calculations.			
Use the terms *concentrated* and *dilute* as qualitative descriptions for the concentration of a solution.			
Calculate percentage yield for chemical reactions			
Calculate the atom economy of a reaction			
Explain that an acid releases H^+ ions in aqueous solution.			
State the formulae of the common acids: hydrochloric, sulfuric and nitric acids.			
State that common bases are metal oxides, metal hydroxides and ammonia.			
State that an alkali is a soluble base that releases OH^- ions in aqueous solution.			
State the formulae of the common alkalis: sodium hydroxide, potassium hydroxide and aqueous ammonia.			
Explain that a salt is produced when the H^+ ion of an acid is replaced by a metal ion or NH_4^+.			
Describe the reactions of an acid with carbonates, bases and alkalis, to form a salt.			
Explain that a base readily accepts H^+ ions from an acid: e.g. OH^- forming H_2O; NH_3 forming NH_4^+ .			
Explain the terms *anhydrous*, *hydrated* and *water of crystallisation.*			
Calculate the formula of a hydrated salt from given percentage composition, mass composition or experimental data.			
Perform acid–base titrations and carry out structured titrations.			
Apply rules for assigning oxidation number to atoms in elements, compounds and ions.			

Describe the terms *oxidation* and *reduction* in terms of: (i) electron transfer (ii) changes in oxidation number.			
Use a Roman numeral to indicate the magnitude of the oxidation state of an element when a name may be ambiguous, e.g. nitrate(III) and nitrate(V).			
Write formulae using oxidation numbers.			
Explain that: (i) metals generally form ions by losing electrons with an increase in oxidation number to form positive ions. (ii) non-metals generally react by gaining electrons with a decrease in oxidation number to form negative ions.			
Describe the redox reactions of metals with dilute hydrochloric and dilute sulfuric acids.			
Interpret and make predictions from redox equations in terms of oxidation numbers and electron loss/gain.			

2.2 - Electrons, bonding and structure			
Electron structure			**Done ✓**
Define the terms *first ionisation energy* and *successive ionisation energy.*			
Explain that ionisation energies are influenced by nuclear charge, electron shielding and the distance of the outermost electron from the nucleus.			
Predict from successive ionisation energies of an element: (i) the number of electrons in each principal quantum shell of an atom (ii) the group of the element.			
State the number of electrons that can fill the first four principal quantum energy levels (shells).			
Describe an orbital as a region that can hold up to two electrons, with opposite spins.			
Describe the shapes of s and p-orbitals.			
State the number of: (i) orbitals making up s, p and d sub-shells (ii) electrons that occupy s, p and d sub-shells.			
Describe the relative energies of s, p and d-orbitals for the principal quantum numbers (shells) 1, 2, 3 and the 4s and 4p-orbitals.			
Deduce the electronic structures of: (i) atoms, given the atomic number, up to $Z = 36$ (ii) ions, given the atomic number and ionic charge, limited to s and p blocks up to $Z = 36$.			
Classify the elements into s, p and d blocks.			
Bonding and structure			**Done ✓**
Describe the term *ionic bonding* as electrostatic attraction between oppositely-charged ions.			
Construct 'dot-and-cross' diagrams to describe ionic bonding.			
Predict ionic charge from the position of an element in the Periodic Table.			
State the formulae for the following ions: NO_3^-, CO_3^{2-}, SO_4^{2-} and NH_4^+			
Describe the term *covalent bond* as a shared pair of electrons.			

Construct 'dot-and-cross' diagrams to describe: (i) single covalent bonding, e.g. as in H_2, Cl_2, HCl, H_2O, NH_3, CH_4, BF_3 and SF_6 (ii) multiple covalent bonding, e.g. as in O_2, N_2 and CO_2 (iii) dative covalent (coordinate) bonding, e.g. as in CO and NH_4^+ (iv) molecules and ions analogous to those specified in (i), (ii) and (iii).				
Explain that the shape of a simple molecule is determined by repulsion between electron pairs surrounding a central atom.				
State that lone pairs of electrons repel more than bonded pairs.				
Explain the shapes of, and bond angles in, molecules and ions for molecules and ions with up to six electron pairs (including lone pairs) surrounding a central atom, e.g. as in: (i) BF_3 (trigonal planar) (ii) CH_4 and NH_4^+ (tetrahedral) (iii) SF_6 (octahedral) (iv) NH_3 (pyramidal) (v) H_2O (non-linear) (vi) CO_2 (linear)				
Predict the shapes of, and bond angles in, molecules and ions analogous to those above.				
Describe the term *electronegativity* as the ability of an atom to attract the bonding electrons in a covalent bond.				
Explain that a permanent dipole polarity may arise when covalently-bonded atoms have different electronegativities, resulting in a polar bond.				
Describe intermolecular forces based on permanent dipoles, as in hydrogen chloride, and instantaneous dipoles (van der Waals' forces), as in the noble gases.				
Describe *hydrogen bonding*, including the role of a lone pair, between molecules containing –OH and –NH groups, i.e. as in H_2O, NH_3 and analogous molecules.				
Describe and explain the anomalous properties of water resulting from hydrogen bonding, e.g.: (i) the density of ice compared with water (ii) its relatively high freezing point and boiling point.				
Describe *metallic bonding* as the attraction of positive ions to delocalised electrons.				
Describe structures as: (i) giant ionic lattices with strong ionic bonding, i.e. as in NaCl (ii) giant covalent lattices, i.e. as in diamond and graphite (iii) giant metallic lattices (iv) simple molecular lattices, i.e. as in I_2 and ice.				

Describe, interpret and/or predict physical properties, including melting and boiling points, electrical conductivity and solubility in terms of: (i) different structures of particles (atoms, molecules, ions and electrons) and the forces between them (ii) different types of bonding (ionic bonding, covalent bonding, metallic bonding, hydrogen bonding, other intermolecular interactions)				
Deduce the type of structure and bonding present from given information.				

Content Checklist

A Level in Chemistry A (H432)

Module 3 - The Periodic Table and Energy			
The Periodic Table			
Describe key events in the development of the modern Periodic Table			
Describe the Periodic Table in terms of the arrangement of elements: (i) by increasing atomic (proton) number (ii) in periods showing repeating trends in physical and chemical properties (iii) in groups having similar physical and chemical properties.			
Describe *periodicity* in terms of a repeating pattern across different periods.			
Explain that atoms of elements in a group have similar outer shell electron structures, resulting in similar properties.			
Describe and explain the variation of the first ionisation energies of elements shown by: (i) a general increase across a period in terms of increasing nuclear charge (ii) a decrease down a group in terms of increasing atomic radius and increasing electron shielding outweighing increasing nuclear charge.			
For the elements of Periods 2 and 3: (i) Describe the variation in electron structures, atomic radii, melting points and boiling points. (ii) Explain variations in melting and boiling points in terms of structure and bonding.			
Interpret data on electronic structures, atomic radii, first ionisation energies, melting points and boiling points to demonstrate periodicity.			
Group 2			
Describe the redox reactions of the Group 2 elements Mg → Ba: (i) with oxygen (ii) with water (iii) with dilute acids			
Explain the trend in reactivity of Group 2 elements down the group due to the increasing ease of forming cations, in terms of atomic size, shielding and nuclear attraction.			
Describe the action of water on oxides of elements in Group 2 and state the approximate pH of any resulting solution.			
Describe the thermal decomposition of the carbonates of elements in Group 2 and the trend in their ease of decomposition.			

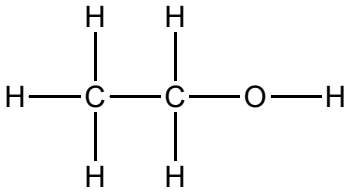
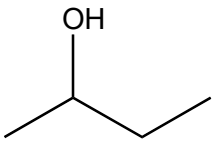
Interpret and make predictions from the chemical and physical properties of Group 2 elements and compounds.			
Explain the use of $Ca(OH)_2$ in agriculture to neutralise acid soils; the use of $Mg(OH)_2$ in some indigestion tablets as an antacid.			
Group 7			
Explain in terms of van der Waals' forces the trend in the boiling points of Cl_2, Br_2 and I_2.			
Describe the redox reactions, including ionic equations, of the Group 7 elements Cl_2, Br_2 and I_2 with other halide ions, in the presence of an organic solvent, to illustrate the relative reactivity of Group 7 elements.			
Explain the trend in reactivity of Group 7 elements down the group from the decreasing ease of forming negative ions, in terms of atomic size, shielding and nuclear attraction.			
Describe the term *disproportionation* as a reaction in which a substance is simultaneously oxidised and reduced, illustrated by: (i) the reaction of chlorine with water as used in water purification (ii) the reaction of chlorine with cold, dilute aqueous sodium hydroxide, as used to form bleach (iii) reactions analogous to those specified in (i) and (ii).			
Interpret and make predictions from the chemical and physical properties of the Group 7 elements and their compounds.			
Contrast the benefits of chlorine use in water treatment (killing bacteria) with associated risks (hazards of toxic chlorine gas and possible risks from formation of chlorinated hydrocarbons).			
Describe the precipitation reactions, including ionic equations, of the aqueous anions Cl^-, Br^- and I^- with aqueous silver ions, followed by aqueous ammonia.			
Describe the use of the precipitation reactions specified above as a test for different halide ions.			
Describe processes and techniques needed to identify the following anions - Carbonate ions - Sulfate ions - Cl^-, Br^-, I^-			
Describe processes and techniques needed to identify the following cations NH_4^+			

Physical Chemistry	Done ✓

Enthalpy changes			
Explain that some chemical reactions are accompanied by enthalpy changes which can be exothermic (ΔH, negative) or endothermic (ΔH, positive).			
Describe the importance of oxidation as an exothermic process in the combustion of fuels and the oxidation of carbohydrates such as glucose in respiration.			
Describe that endothermic processes require an input of heat energy, i.e. the thermal decomposition of calcium carbonate.			
Construct a simple enthalpy profile diagram for a reaction to show the difference in the enthalpy of the reactants compared with that of the products.			
Explain qualitatively, using enthalpy profile diagrams, the term *activation energy.*			
Define and use the terms: (i) *standard conditions* (ii) *enthalpy change of reaction* (iii) *enthalpy change of formation* (iv) *enthalpy change of combustion* (v) enthalpy change of neutralisation			
Calculate enthalpy changes from appropriate experimental results directly, including use of the relationship: energy change = $mc\Delta T$			
Explain exothermic and endothermic reactions in terms of enthalpy changes associated with the breaking and making of chemical bonds.			
Define and use the term *average bond enthalpy* (ΔH positive; bond breaking of one mole of bonds).			
Calculate an enthalpy change of reaction from average bond enthalpies.			
Use Hess' law to construct enthalpy cycles and carry out calculations to determine: (i) an enthalpy change of reaction from enthalpy changes of combustion (ii) an enthalpy change of reaction from enthalpy changes of formation (iii) an enthalpy change of reaction from an unfamiliar enthalpy cycle.			
Describe the techniques and procedures used to determine enthalpy changes directly and indirectly			
Rates and equilibrium			
Describe qualitatively, in terms of collision theory, the effect of concentration changes on the rate of a reaction.			
Explain why an increase in the pressure of a gas, increasing its concentration, may increase the rate of a reaction involving gases.			

	Calculation of reaction rate from the gradients of graphs measuring how a physical quantity changes with time			
	State that a catalyst speeds up a reaction without being consumed by the overall reaction.			
	Explain that catalysts: (i) affect the conditions that are needed, often requiring lower temperatures and reducing energy demand and CO_2 emissions from burning of fossil fuels (ii) enable different reactions to be used, with better atom economy and with reduced waste (iii) are often enzymes, generating very specific products, and operating effectively close to room temperatures and pressures (iv) have great economic importance, i.e. iron in ammonia production; Ziegler–Natta catalyst in poly(ethene) production; platinum/palladium/rhodium in catalytic converters.			
	Explanation of the terms homogeneous and heterogeneous catalysts			
	Explain, using enthalpy profile diagrams, how the presence of a catalyst allows a reaction to proceed *via* a different route with a lower activation energy, giving rise to an increased reaction rate.			
	Explain qualitatively the Boltzmann distribution and its relationship with activation energy.			
	Describe qualitatively, using the Boltzmann distribution, the effect of temperature changes on the proportion of molecules exceeding the activation energy and hence the reaction rate.			
	Interpret the catalytic behaviour above (how the presence of a catalyst allows a reaction to proceed *via* a different route with a lower activation energy, giving rise to an increased reaction rate), in terms of the Boltzmann distribution			
	Explain that a dynamic equilibrium exists in a closed system when the rate of the forward reaction is equal to the rate of the reverse reaction and the concentrations of reactants and products do not change.			
	State le Chatelier's principle.			
	Apply le Chatelier's principle to deduce qualitatively (from appropriate information) the effect of a change in temperature, concentration or pressure, on a homogeneous system in equilibrium.			
	Explanation that a catalyst increases the rate of both forward and reverse reactions in an equilibrium by the same amount resulting in an unchanged position of equilibrium			
	The techniques and procedures used to investigate changes to the position of equilibrium for changes in concentration and temperature			
	Explain from given data the importance in the chemical industry of a compromise between chemical equilibrium and reaction rate.			
	Write expressions for the equilibrium constant, K_c, and calculate values for the equilibrium constant, K_c, from provided equilibrium concentrations			
	Estimate the position of equilibrium from the magnitude of K_c			

Content Checklist

Module 4, Part 1 - Basic concepts and hydrocarbons			
Basic concepts			
Interpret and use the terms: (i) *empirical formula* as the simplest whole number ratio of atoms of each element present in a compound (ii) *molecular formula* as the actual number of atoms of each element in a molecule (iii) *general formula* as the simplest algebraic formula of a member of a homologous series, *i.e.* for an alkane: C_nH_{2n+2} (iv) *structural formula* as the minimal detail that shows the arrangement of atoms in a molecule e.g. for butane: $CH_3CH_2CH_2CH_3$ or $CH_3(CH_2)_2CH_3$ (v) *displayed formula* as the relative positioning of atoms and the bonds between them, *i.e.* for ethanol:  (vi) skeletal formula as the simplified organic formula, shown by removing hydrogen atoms from alkyl chains, leaving just a carbon skeleton and associated functional groups, i.e. for butan-2-ol: 			
Interpret and use the terms: (i) *homologous series* as a series of organic compounds having the same functional group but with each successive member differing by CH_2 (ii) *functional group* as a group of atoms responsible for the characteristic reactions of a compound.			
Use the general formula of a homologous series to predict the formula of any member of the series.			
State the names of the first ten members of the alkanes homologous series.			
Use IUPAC rules of nomenclature for systematically naming organic compounds.			

Describe and explain the terms: (i) *structural isomers* as compounds with the same molecular formula but different structural formulae (ii) *stereoisomers* as compounds with the same structural formula but with a different arrangement in space (iii) *E*/*Z* isomerism as an example of stereoisomerism, in terms of restricted rotation about a double bond and the requirement for two different groups to be attached to each carbon atom of the C=C group (iv) *cis-trans isomerism* as a special case of *EIZ* isomerism in which there are two non-hydrogen groups and two hydrogens around the double bond.			
Determine the possible structural formulae and/or stereoisomers of an organic molecule, given its molecular formula.			
Describe the different types of covalent bond fission: (i) homolytic fission forming two radicals (ii) heterolytic fission forming a cation and an anion.			
Explain the term radical (a species with an unpaired electron) and use “dots” to represent species that are radicals in mechanisms			
Describe a ‘curly arrow’ as the movement of an electron pair, showing either breaking or formation of a covalent bond.			
Outline reaction mechanisms, using diagrams, to show clearly the movement of an electron pair with ‘curly arrows’. And relevant dipoles.			
Alkanes			
Explain that a *hydrocarbon* is a compound of hydrogen and carbon only.			
Explain the use of crude oil as a source of hydrocarbons, separated as fractions with different boiling points by fractional distillation, which can be used as fuels or for processing into petrochemicals.			
State that alkanes and cycloalkanes are saturated hydrocarbons containing single C-C and C-H sigma bonds (overlap of orbitals directly between the bonding atoms); free rotation of the sigma bond.			
State and explain the tetrahedral shape around each carbon atom in alkanes in terms of electron pair repulsion.			
Explain the variations in the boiling points of alkanes with different carbon-chain length and branching in terms of induced dipole- dipole attractions (London Forces).			
The low reactivity of alkanes with many reagents in terms of high bond enthalpy and the very low bond polarity of the sigma bonds present.			

Complete combustion of alkanes as used as fuels			
Explain using equations the incomplete combustion of alkanes in a limited supply of oxygen and outline the potential dangers arising from production of CO in the home and from car use.			
Describe the reaction of alkanes with Cl_2 and by Br_2 by radical substitution using ultraviolet radiation, to form halogenoalkanes			
Define the term *radical* as a species with an unpaired electron.			
Describe how homolytic fission leads to the mechanism of radical substitution in alkanes in terms of initiation, propagation and termination reactions.			
Explain the limitations of radical substitution in synthesis, arising from further substitution with and reactions at different positions in a carbon chain. Leading to the formation of a mixture of products.			
Alkenes			
State that alkenes and cycloalkenes are unsaturated hydrocarbons containing a C=C bond comprising a pi bond (sideways overlap of adjacent p-orbitals above and below the bonding C atoms) and a sigma bond (overlap of orbitals directly between the bonding atoms); restricted rotation of the pi bond.			
State and explain the trigonal planar shape and bond angle around each carbon in the C=C of alkenes in terms of electron pair repulsion.			
Explanation of isomerism in terms of; - Stereoisomers (compounds with the same structural formula but with a different arrangement in space - E/Z isomerism (an example of stereoisomerism, in terms of restricted rotation around a double bond and the requirement for 2 different groups to be attached to each carbon atom of the C=C group) - Cis-trans isomerism (a special case of E/Z isomerism in which 2 of the substituent groups attached to each carbon of the same C=C group are the same)			
Use of Cahn-Ingold-Prelog (CIP) priority rules to identify the E and Z stereoisomers			
Determination of the possible E/Z or cis-trans stereoisomers of an organic molecule, given its structural formula			
The reactivity of alkenes in terms of the relatively low bond enthalpy of the pi bond.			

Describe addition reactions of alkenes, *i.e*. by ethene and propene with: (i) hydrogen in the presence of a suitable catalyst, i.e. Ni, to form alkanes (ii) halogens to form dihalogenoalkanes, including the use of bromine to detect the presence of a double C=C bond as a test for unsaturation (iii) hydrogen halides to form halogenoalkanes (iv) steam in the presence of an acid catalyst, e.g. H_3PO_4, to form alcohols.			
Define an *electrophile* as an electron pair acceptor.			
Describe how heterolytic fission leads to the mechanism of electrophilic addition in alkenes.			
Use of Markownikoff's rule to predict formation of major organic product in addition reactions of H-X to unsymmetrical alkanes, e.g. H-Br to propene, in terms of the relative stabilities of carbocation intermediates in the mechanism.			
Describe the addition polymerisation of alkenes.			
Deduce the repeat unit of an addition polymer obtained from a given monomer.			
Identify the monomer that would produce a given section of an addition polymer.			
Outline the benefits for sustainability of processing waste polymers by: - Combustion for energy production - Use as an organic feedstock for the production of plastics and other organic chemicals - Removal of toxic waste products, e.g. removal of HCl formed during disposal by combustion of halogenated plastics (e.g. PVC)			
Benefits to the environment of development of biodegradable and photodegradable polymers			

Module 4, Part 2 - Alcohols, halogenoalkanes and analysis			
Alcohols			
Explain in terms of hydrogen bonding the water solubility and the relatively low volatility of alcohols.			
Describe the industrial production of ethanol by: (i) fermentation from sugars, *i.e.* from glucose (ii) the reaction of ethene with steam in the presence of an acid catalyst.			
Outline for alcohols: (i) the use of ethanol in alcoholic drinks and as a solvent in the form of methylated spirits (ii) the use of methanol as a petrol additive to improve combustion and its increasing importance as a feedstock in the production of organic chemicals.			
Classify alcohols into primary, secondary and tertiary alcohols.			
Describe the combustion of alcohols.			
Describe the oxidation of alcohols using $Cr_2O_7^{2-}/H^+$ (*i.e.* $K_2Cr_2O_7/H_2SO_4$), including: (i) the oxidation of primary alcohols to form aldehydes and carboxylic acids; the control of the oxidation product using different reaction conditions (ii) the oxidation of secondary alcohols to form ketones (iii) the resistance to oxidation of tertiary alcohols.			
Describe elimination of H_2O from alcohols in the presence of an acid catalyst (e.g. H_3PO_4 or H_2SO_4) and heat to form alkenes.			
Substitution with halide ions in the presence of acid (e.g. $NaBr/H_2SO_4$) to form haloalkanes			
Haloalkanes			
Describe the hydrolysis of haloalkanes as a substitution reaction: - By aqueous alkali - By water in the presence of $AgNO_3$ and ethanol to compare experimentally the rates of hydrolysis of different carbon-halogen compounds.			
Define the term *nucleophile* as an electron pair donor.			
Describe the mechanism of nucleophilic substitution in the hydrolysis of primary haloalkanes with aqueous alkali.			

Explain the trend in the rates of hydrolysis of primary halogenoalkanes in terms of the relative bond enthalpies of carbon–halogen bonds (C–F, C–Cl, C–Br and C–I).			
Production of halogen radicals by the action of UV radiation on CFCs in the upper atmosphere and the resulting catalysed breakdown of the Earth's protective ozone layer, including equations to represent: - The production of halogen radicals - The catalysed breakdown of ozone by Cl. And other radicals e.g. .NO			
Organic Synthesis : Practical Skills			
The techniques and procedures for: - Use of quickfit apparatus including distillation and heating under reflux - Preparation and purification an organic liquid including the use of a separating funnel in order to remove an organic layer from an aqueous layer, drying with an anhydrous salt (e.g. $MgSO_4$, $CaCl_2$) and redistillation			
Synthetic routes for an organic molecule containing several functional groups: - Identification of individual functional groups - Prediction of properties and reactions			
2 stage synthetic routes for preparing organic compounds			
Modern analytical techniques			
State that absorption of infrared radiation causes covalent bonds to vibrate.			
Absorption of IR radiation by atmospheric gases containing C=C, O-H and C-H bonds (e.g. H_2O, CO_2, and CH_4), the suspected link to global warming and resulting changes to energy usage.			
Identify, using an infrared spectrum of an organic compound: (i) an alcohol from an absorption peak of the O–H bond (ii) an aldehyde or ketone from an absorption peak of the C=O bond (iii) a carboxylic acid from an absorption peak of the C=O bond and a broad absorption peak of the O–H bond.			
Interpretations and predictions of an infrared spectrum of familiar or unfamiliar substances using supplied data.			
Use of IR spectroscopy to monitor gases causing air pollution (e.g. NO and CO from car emissions) and in modern breathalysers to measure ethanol in breath.			

Outline the use of mass spectrometry: (i) in the determination of relative isotopic masses (ii) as a method for identifying elements, *i.e.* for use in the Mars space probe and in monitoring levels of environmental pollution such as lead.				
Interpret mass spectra of elements in terms of isotopic abundances.				
Use the molecular ion peak in a mass spectrum of an organic molecule to determine its molecular mass.				
Suggest the identity of the major fragment ions, *i.e.* m/z = 29 as $CH_3CH_2^+$, in a given mass spectrum (limited to alkanes, alkenes and alcohols).				
Use molecular ion peaks and fragmentation peaks to identify structures (limited to unipositive ions).				
Deduction of the structures of organic compounds from different analytical data including: - Elemental analysis - Mass Spectra - IR Spectra				