Student Learning Reflection & Personalised Learning Checklist

Subject/Course:	Chemistry
Student Name:	

Year 12 Curriculum

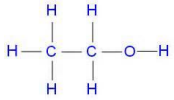
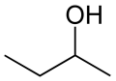
Topic	Key knowledge/skills	Self Assessment		
		Red	Amber	Green
2.1.1 - Atomic Structure & Isotopes	(a) isotopes as atoms of the same element with different numbers of neutrons and different masses			
	(b) atomic structure in terms of the numbers of protons, neutrons and electrons for atoms and ions, given the atomic number, mass number and any ionic charge			
	(c) explanation of the terms relative isotopic mass (mass compared with 1/12th mass of carbon-12) and relative atomic mass (weighted mean mass compared with 1/12th mass of carbon-12), based on the mass of a ^{12}C atom, the standard for atomic masses			
	(d) use of mass spectrometry in: **(i)** the determination of relative isotopic masses and relative abundances of the isotope, **(ii)** calculation of the relative atomic mass of an element from the relative abundances of its isotopes			
	(e) use of the terms relative molecular mass, Mr, and relative formula mass and their calculation from relative atomic masses.			
2.1.2 - Compounds, Formulae & Equations	(a) the writing of formulae of ionic compounds from ionic charges, including: **(i)** prediction of ionic charge from the position of an element in the periodic table **(ii)** recall of the names and formulae for the following ions: NO^{3-}, CO_3^{2-}, SO_4^{2-}, OH^-, NH_4^+, Zn^{2+} and Ag^+			
	(b) construction of balanced chemical equations (including ionic equations), including state symbols, for reactions studied and for unfamiliar reactions given appropriate information.			
2.1.3 - Amount of Substance	(a) explanation and use of the terms: **(i)** *amount of substance* **(ii)** *mole* (symbol 'mol'), as the unit for amount of substance **(iii)** the *Avogadro constant*, *N*A (the number of particles per mole, 6.02×10^{23} mol^{-1}) **(iv)** *molar mass* (mass per mole, units g mol^{-1}), **(v)** *molar gas volume* (gas volume per mole, units dm^3 mol^{-1})			
	(b) use of the terms: **(i)** *empirical formula* (the simplest whole number ratio of atoms of each element present in a compound) **(ii)** *molecular formula* (the number and type of atoms of each element in a molecule)			
	(c) calculations of empirical and molecular formulae, from composition by mass or percentage compositions by mass and relative molecular mass			
	(d) the terms *anhydrous*, *hydrated* and *water of crystallisation* and calculation of the formula of a hydrated salt from given percentage composition, mass composition or based on experimental results			
	(e) calculations, using amount of substance in mol, involving: **(i)** mass **(ii)** gas volume **(iii)** solution volume and concentration			
	(f) the ideal gas equation: $pV = nRT$			
	(g) use of stoichiometric relationships in calculations			

	(h) calculations to determine: **(i)** the percentage yield of a reaction or related quantities **(ii)** the atom economy of a reaction			
	(i) the techniques and procedures required during experiments requiring the measurement of mass, volumes of solutions and gas volumes			
	(j) the benefits for sustainability of developing chemical processes with a high atom economy.			
2.1.4 - Acids	(a) the formulae of the common acids (HC/, H_2SO_4, HNO_3 and CH_3COOH) and the common alkalis (NaOH, KOH and NH_3) and explanation that acids release H^+ ions in aqueous solution and alkalis release OH^- ions in aqueous solution			
	(b) qualitative explanation of strong and weak acids in terms of relative dissociations			
	(c) neutralisation as the reaction of: **(i)** H^+ and OH^- to form H_2O **(ii)** acids with bases, including carbonates, metal oxides and alkalis (water-soluble bases), to form salts, including full equations			
	(d) the techniques and procedures used when preparing a standard solution of required concentration and carrying out acid–base titrations			
	(e) structured and non-structured titration calculations, based on experimental results of familiar and non-familiar acids and bases.			
	(f) describe the redox reactions of metals with dilute hydrochloric and dilute sulfuric acids			
	(g) interpret and make predictions from redox equations in terms of oxidation numbers and electron loss/gain			
2.1.5 - Redox	(a) rules for assigning and calculating oxidation number for atoms in elements, compounds and ions			
	(b) writing formulae using oxidation numbers			
	(c) use of a Roman numeral to indicate the magnitude of the oxidation number when an element may have compounds/ions with different oxidation numbers			
	(d) oxidation and reduction in terms of: **(i)** electron transfer **(ii)** changes in oxidation number			
	(e) redox reactions of metals with acids to form salts, including full equations			
	(f) interpretation of redox equations in **(e)**, and unfamiliar redox reactions, to make predictions in terms of oxidation numbers and electron loss/ gain.			
	(g) interpret and make predictions from redox equations in terms of oxidation numbers and electron loss/gain			
2.2.1 - Electron Structure	(a) the number of electrons that can fill the first four shells			
	(b) atomic orbitals, including: **(i)** as a region around the nucleus that can hold up to two electrons, with opposite spins **(ii)** the shapes of s- and p-orbitals **(iii)** the number of orbitals making up s-, p- and d-sub-shells, and the number of electrons that can fill s-, p- and d-sub-shells			
	(c) filling of orbitals: **(i)** for the first three shells and the 4s and 4p orbitals in order of increasing energy **(ii)** for orbitals with the same energy, occupation singly before pairing			
	(d) deduction of the electron configurations 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 $Z = 36$.			
2.2.2 - Bonding and Structure	(a) ionic bonding as electrostatic attraction between positive and negative ions, and the construction of '*dot-and-cross*' diagrams			
	(b) explanation of the solid structures of giant ionic lattices, resulting from oppositely charged ions strongly attracted in all directions e.g. NaC/			
	(c) explanation of the effect of structure and bonding on the physical properties of ionic compounds, including melting and boiling points, solubility and electrical conductivity in solid, liquid and aqueous states			

	(d) covalent bond as the strong electrostatic attraction between a shared pair of electrons and the nuclei of the bonded atoms			
	(e) construction of '*dot-and-cross*' diagrams of molecules and ions to describe: **(i)** single covalent bonding **(ii)** multiple covalent bonding **(iii)** dative covalent (coordinate) bonding			
	(f) use of the term *average bond enthalpy* as a measurement of covalent bond strength			
	(g) the shapes of, and bond angles in, molecules and ions with up to six electron pairs (including lone pairs) surrounding the central atom as predicted by electron pair repulsion, including the relative repulsive strengths of bonded pairs and lone pairs of electrons			
	(h) electron pair repulsion to explain the following shapes of molecules and ions: linear, non-linear, trigonal planar, pyramidal, tetrahedral and octahedral			
	(i) electronegativity as the ability of an atom to attract the bonding electrons in a covalent bond; interpretation of Pauling electronegativity values			
	(j) explanation of: a polar bond and permanent dipole within molecules containing covalently-bonded atoms with different electronegativities **(ii)** a polar molecule and overall dipole in terms of permanent dipole(s) and molecular shape			
	(k) intermolecular forces based on permanent dipole–dipole interactions and induced dipole– dipole interactions			
	(l) hydrogen bonding as intermolecular bonding between molecules containing N, O or F and the H atom of –NH, –OH or HF			
	(m) explanation of anomalous properties of H_2O resulting from hydrogen bonding, e.g.: **(i)** the density of ice compared with water **(ii)** its relatively high melting and boiling points			
	(n) explanation of the solid structures of simple molecular lattices, as covalently bonded molecules attracted by intermolecular forces, e.g. I_2, ice			
	(o) explanation of the effect of structure and bonding on the physical properties of covalent compounds with simple molecular lattice structures including melting and boiling points, solubility and electrical conductivity			
3.1.1 - Periodicity	(a) the periodic table as the arrangement of elements: **(i)** by increasing atomic (proton) number **(ii)** in periods showing repeating trends in physical and chemical properties (periodicity) **(iii)** in groups having similar chemical properties			
	(b) **(i)** the periodic trend in electron configurations across Periods 2 and 3 **(ii)** classification of elements into s-, p- and d-blocks			
	(c) first ionisation energy (removal of 1 mol of electrons from 1 mol of gaseous atoms) and successive ionisation energy, and: **(i)** explanation of the trend in first ionisation energies across Periods 2 and 3, and down a group, in terms of attraction, nuclear charge and atomic radius **(ii)** prediction from successive ionisation energies of the number of electrons in each shell of an atom and the group of an element			
	(d) explanation of: **(i)** metallic bonding as strong electrostatic attraction between cations (positive ions) and delocalised electrons **(ii)** a giant metallic lattice structure, e.g. all metals			
	(e) explanation of the solid giant covalent lattices of carbon (diamond, graphite and graphene) and silicon as networks of atoms bonded by strong covalent bonds			
	(f) explanation of physical properties of giant metallic and giant covalent lattices, including melting and boiling points, solubility and electrical conductivity in terms of structure and bonding			
	(g) explanation of the variation in melting points across Periods 2 and 3 in terms of structure and bonding.			

3.1.2 - Group 2	(a) the outer shell s2 electron configuration and the loss of these electrons in redox reactions to form 2^+ ions			
	(b) the relative reactivities of the Group 2 elements Mg → Ba shown by their redox reactions with: **(i)** oxygen **(ii)** water **(iii)** dilute acids			
	(c) the trend in reactivity in terms of the first and second ionisation energies of Group 2 elements down the group			
	(d) the action of water on Group 2 oxides and the approximate pH of any resulting solutions, including the trend of increasing alkalinity			
	(e) uses of some Group 2 compounds as bases, including equations, for example (but not limited to): (i) $Ca(OH)_2$ in agriculture to neutralise acid soils (ii) $Mg(OH)_2$ and $CaCO_3$ as 'antacids' in treating indigestion.			
3.1.3 - Group 7	(a) existence of halogens as diatomic molecules and explanation of the trend in the boiling points of $C\ell_2$, Br_2 and I_2, in terms of induced dipole–dipole interactions (London forces)			
	(b) the outer shell s2p5 electron configuration and the gaining of one electron in many redox reactions to form 1^- ions			
	(c) the trend in reactivity of the halogens $C\ell_2$, Br_2 and I_2, illustrated by reaction with other halide ions			
	(d)) explanation of the trend in reactivity shown in **(c)**, from the decreasing ease of forming 1– ions, in terms of attraction, atomic radius and electron shielding			
	(e) explanation of the term *disproportionation* as oxidation and reduction of the same element, illustrated by: **(i)** the reaction of chlorine with water as used in water 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)**			
	(f) the benefits of chlorine use in water treatment (killing bacteria) contrasted with associated risks (e.g. hazards of toxic chlorine gas and possible risks from formation of chlorinated hydrocarbons)			
	(g) the precipitation reactions, including ionic equations, of the aqueous anions $C\ell^-$, Br^- and I^- with aqueous silver ions, followed by aqueous ammonia, and their use as a test for different halide ions.			
3.1.4 - Qualitative Analysis	(a) qualitative analysis of ions on a test-tube scale; processes and techniques needed to identify the following ions in an unknown compound: **(i)** anions: • $CO3^{2-}$, by reaction with $H^+(aq)$ forming $CO_2(g)$ • $SO4^{2-}$, by precipitation with $Ba^{2+}(aq)$ • $C\ell^-$, Br^-, I^- **(ii)** cations: NH^{4+}, by reaction with warm NaOH(aq) forming NH_3.			
3.2.1 - Enthalpy Changes	(a) explanation that some chemical reactions are accompanied by enthalpy changes that are exothermic (ΔH, negative) or endothermic (ΔH, positive)			
	(b) construction of enthalpy profile diagrams to show the difference in the enthalpy of reactants compared with products			
	(c) qualitative explanation of the term *activation energy*, including use of enthalpy profile diagrams			
	(d) explanation and use of the terms: **(i)** *standard conditions* and *standard states* (physical states under standard conditions) **(ii)** *enthalpy change of reaction* (enthalpy change associated with a stated equation, ΔrH) **(iii)** *enthalpy change of formation* (formation of 1 mol of a compound from its elements, ΔfH) **(iv)** *enthalpy change of combustion* (complete combustion of 1 mol of a substance, ΔcH) **(v)** *enthalpy change of neutralisation* (formation of 1 mol of water from neutralisation, $\Delta neutH$)			
	(e) determination of enthalpy changes directly from appropriate experimental results, including use of the relationship: $q = mc\Delta T$			
	(f) **(i)** explanation of the term *average bond enthalpy* (breaking of 1 mol of bonds in gaseous molecules)			

	(ii) explanation of exothermic and endothermic reactions in terms of enthalpy changes associated with the breaking and making of chemical bonds **(iii)** use of average bond enthalpies to calculate enthalpy changes and related quantities			
	(g) Hess' law for construction of enthalpy cycles and calculations to determine indirectly: **(i)** an enthalpy change of reaction from enthalpy changes of combustion **(ii)** an enthalpy change of reaction from enthalpy changes of formation **(iii)** enthalpy changes from unfamiliar enthalpy cycles			
	(h) the techniques and procedures used to determine enthalpy changes directly and indirectly.			
3.2.2 - Reaction Rates	(a) the effect of concentration, including the pressure of gases, on the rate of a reaction, in terms of frequency of collisions			
	(b) calculation of reaction rate from the gradients of graphs measuring how a physical quantity changes with time			
	(c) explanation of the role of a catalyst: **(i)** in increasing reaction rate without being used up by the overall reaction **(ii)** in allowing a reaction to proceed via a different route with lower activation energy, as shown by enthalpy profile diagrams			
	(d) **(i)** explanation of the terms *homogeneous* and *heterogeneous* catalysts **(ii)** explanation that catalysts have great economic importance and benefits for increased sustainability by lowering temperatures and reducing energy demand from combustion of fossil fuels with resulting reduction in CO_2 emissions			
	(e) the techniques and procedures used to investigate reaction rates including the measurement of mass, gas volumes and time			
	(f) qualitative explanation of the Boltzmann distribution and its relationship with activation energy			
	(g) explanation, using Boltzmann distributions, of the qualitative effect on the proportion of molecules exceeding the activation energy and hence the reaction rate, for: **(i)** temperature changes **(ii)** catalytic behaviour.			
3.2.3 - Chemical Equilibrium	(a) explanation 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			
	(b) le Chatelier's principle and its application for homogeneous equilibria to deduce qualitatively the effect of a change in temperature, pressure or concentration on the position of equilibrium			
	(c) 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			
	(d) the techniques and procedures used to investigate changes to the position of equilibrium for changes in concentration and temperature			
	(e) explanation of the importance to the chemical industry of a compromise between chemical equilibrium and reaction rate in deciding the operational conditions			
	(f) expressions for the equilibrium constant, K_c , for homogeneous reactions and calculations of the equilibrium constant, K_c , from provided equilibrium concentrations			
	(g) estimation of the position of equilibrium from the magnitude of K_c.			
4.1.1 - Basic Organic Chemistry	(a) application of IUPAC rules of nomenclature for systematically naming organic compounds			
	(b) interpretation and use of the terms: **(i)** *general formula* (the simplest algebraic formula of a member of a homologous series) e.g. for an alkane: C_nH_{2n+2}			

	(ii) *structural formula* (the minimal detail that shows the arrangement of atoms in a molecule) e.g. for butane: $CH_3CH_2CH_2CH_3$ or $CH_3(CH_2)_2CH_3$ **(iii)** *displayed formula* (the relative positioning of atoms and the bonds between them) e.g. for ethanol:  **(iv)** *skeletal formula* (the simplified organic formula, shown by removing hydrogen atoms from alkyl chains, leaving just a carbon skeleton and associated functional groups) e.g. for butan-2-ol: 			
	(c) interpretation and use of the terms: **(i)** *homologous series* (a series of organic compounds having the same functional group but with each successive member differing by CH_2) **(ii)** *functional group* (a group of atoms responsible for the characteristic reactions of a compound) **(iii)** *alkyl group* (of formula C_nH2_{n+1}) **(iv)** *aliphatic* (a compound containing carbon and hydrogen joined together in straight chains, branched chains or non-aromatic rings) **(v)** *alicyclic* (an aliphatic compound arranged in non-aromatic rings with or without side chains) *aromatic* (a compound containing a benzene ring) **(vii)** *saturated* (single carbon–carbon bonds only) and *unsaturated* (the presence of multiple carbon–carbon bonds, including C=C, CC /and aromatic rings)			
	(d) use of the general formula of a homologous series to predict the formula of any member of the series			
	(e) explanation of the term *structural isomers* (compounds with the same molecular formula but different structural formulae) and determination of possible structural formulae of an organic molecule, given its molecular formula			
	(f) the different types of covalent bond fission: **(i)** homolytic fission (in terms of each bonding atom receiving one electron from the bonded pair, forming two radicals) **(ii)** heterolytic fission (in terms of one bonding atom receiving both electrons from the bonded pair)			
	(g) the term *radical* (a species with an unpaired electron) and use of 'dots' to represent species that are radicals in mechanisms			
	a 'curly arrow' described as the movement of an electron pair, showing either heterolytic fission or formation of a covalent bond			
	(i) reaction mechanisms, using diagrams, to show clearly the movement of an electron pair with 'curly arrows' and relevant dipoles.			
4.1.2 - Alkanes	(a) alkanes as saturated hydrocarbons containing single C–C and C–H bonds as σ-bonds (overlap of orbitals directly between the bonding atoms); free rotation of the σ-bond			
	(b) explanation of the tetrahedral shape and bond angle around each carbon atom in alkanes in terms of electron pair repulsion			
	(c) explanation of the variations in boiling points of alkanes with different carbon-chain length and branching, in terms of induced dipole–dipole interactions (London forces)			
	(d) the low reactivity of alkanes with many reagents in terms of the high bond enthalpy and very low polarity of the σ-bonds present			
	(e) complete combustion of alkanes, as used in fuels, and the incomplete combustion of alkane fuels in a limited supply of oxygen with the resulting potential dangers from CO			
	(f) the reaction of alkanes with chlorine and bromine by radical substitution using ultraviolet radiation, including a mechanism involving homolytic fission and radical reactions in terms of initiation, propagation and termination			

	(g) the limitations of radical substitution in synthesis by the formation of a mixture of organic products, in terms of further substitution and reactions at different positions in a carbon chain.			
4.1.3 - Alkenes	(a) alkenes as unsaturated hydrocarbons containing a C=C bond comprising a π-bond (sideways overlap of adjacent p-orbitals above and below the bonding C atoms) and a σ-bond (overlap of orbitals directly between the bonding atoms); restricted rotation of the π-bond			
	(b) explanation of the trigonal planar shape and bond angle around each carbon in the C=C of alkenes in terms of electron pair repulsion			
	(c) **(i)** explanation of the terms: • *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 about a double bond and the requirement for two 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 two of the substituent groups attached to each carbon atom of the C=C group are the same) **(ii)** use of Cahn–Ingold–Prelog (CIP) priority rules to identify the *E* and *Z* stereoisomers			
	(d) determination of possible *E*/*Z* or *cis–trans* stereoisomers of an organic molecule, given its structural formula			
	(e) the reactivity of alkenes in terms of the relatively low bond enthalpy of the π-bond			
	(f) addition reactions of alkenes with: **(i)** hydrogen in the presence of a suitable catalyst, e.g. Ni, to form alkanes **(ii)** halogens to form dihaloalkanes, including the use of bromine to detect the presence of a double C=C bond as a test for unsaturation in a carbon chain **(iii)** hydrogen halides to form haloalkanes **(iv)** steam in the presence of an acid catalyst, e.g. H_3PO_4, to form alcohols			
	(g) definition and use of the term *electrophile* (an electron pair acceptor)			
	(h) the mechanism of electrophilic addition in alkenes by heterolytic fission			
	(i) use of Markownikoff's rule to predict formation of a major organic product in addition reactions of H–X to unsymmetrical alkenes, e.g. H–Br to propene, in terms of the relative stabilities of carbocation intermediates in the mechanism			
	(j) addition polymerisation of alkenes and substituted alkenes, including: **(i)** the repeat unit of an addition polymer deduced from a given monomer **(ii)** identification of the monomer that would produce a given section of an addition polymer			
	(k) the benefits for sustainability of processing waste polymers by: **(i)** combustion for energy production **(ii)** use as an organic feedstock for the production of plastics and other organic chemicals **(iii)** removal of toxic waste products, e.g. removal of HC/ formed during disposal by combustion of halogenated plastics (e.g. PVC)			
4.2.1 - Alcohols	(a) **(i)** the polarity of alcohols and an explanation, in terms of hydrogen bonding, of the water solubility and the relatively low volatility of alcohols compared with alkanes **(ii)** classification of alcohols into primary, secondary and tertiary alcohols			
	(b) combustion of alcohols			
	(c) oxidation of alcohols by an oxidising agent, e.g. $Cr_2O_7^{2-}$/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			
	(d) elimination of H2O from alcohols in the presence of an acid catalyst (e.g. H_3PO_4 or H2SO4) and heat to form alkenes			
	(e) substitution with halide ions in the presence of acid (e.g. $NaBr/H_2SO_4$) to form haloalkanes.			
4.2.2 - Haloalkanes	(a) hydrolysis of haloalkanes in a substitution reaction: **(i)** by aqueous alkali **(ii)** by water in the presence of $AgNO_3$ and ethanol to compare experimentally the rates of hydrolysis of different carbon– halogen bonds			
	(b) definition and use of the term *nucleophile* (an electron pair donor)			
	(c) the mechanism of nucleophilic substitution in the hydrolysis of primary haloalkanes with aqueous alkali			
	(d) explanation of the trend in the rates of hydrolysis of primary haloalkanes in terms of the bond enthalpies of carbon–halogen bonds (C–F, C–C/, C–Br and C–I)			
4.2.3 - Organic Synthesis	(a) the techniques and procedures for: **(i)** use of Quickfit apparatus including for distillation and heating under reflux **(ii)** preparation and purification of an organic liquid including: • use of a separating funnel to remove an organic layer from an aqueous layer • drying with an anhydrous salt (e.g. $MgSO_4$, $CaC/_2$) • redistillation			
	(b) for an organic molecule containing several functional groups: **(i)** identification of individual functional groups **(ii)** prediction of properties and reactions			
	(c) two-stage synthetic routes for preparing organic compounds.			
4.2.4 - Analytical Techniques	(a) infrared (IR) radiation causes covalent bonds to vibrate more and absorb energy			
	(b) absorption of infrared radiation by atmospheric gases containing C=O, 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			
	(c) use of an infrared spectrum of an organic compound 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 **(iii)** a carboxylic acid from an absorption peak of the C=O bond and a broad absorption peak of the O–H bond			
	(d) interpretations and predictions of an infrared spectrum of familiar or unfamiliar substances using supplied data			
	(e) use of infrared spectroscopy to monitor gases causing air pollution (e.g. CO and NO from car emissions) and in modern breathalysers to measure ethanol in the breath			
	(f) use of a mass spectrum of an organic compound to identify the molecular ion peak and hence to determine molecular mass			
	(g) analysis of fragmentation peaks in a mass spectrum to identify parts of structures.			
	(h) deduction of the structures of organic compounds from different analytical data including: **(i)** elemental analysis **(ii)** mass spectra **(iii)** IR spectra.			

Year 13 Curriculum

Topic	Key knowledge/skills	Self Assessment Red	Amber	Green
5.1.1 - How Fast?	(a) explanation and use of the terms: *rate of reaction, order, overall order, rate constant, half-life, rate-determining step*			
	(b) deduction of: (i) orders from experimental data (ii) a rate equation from orders of the form: rate = $k[A]^m[B]^n$, where m and n are 0, 1 or 2			
	(c) calculation of the rate constant, k, and related quantities, from a rate equation including determination of units			
	(d) from a concentration–time graph: (i) deduction of the order (0 or 1) with respect to a reactant from the shape of the graph (ii) calculation of reaction rates from the measurement of gradients			
	(e) from a concentration–time graph of a first order reaction, measurement of constant half-life, $t_{1/2}$			
	(f) for a first order reaction, determination of the rate constant, k, from the constant half-life, $t_{1/2}$, using the relationship: $k = \ln 2/t_{1/2}$			
	(g) from a rate–concentration graph: (i) deduction of the order (0, 1 or 2) with respect to a reactant from the shape of the graph (ii) determination of rate constant for a first order reaction from the gradient			
	(h) the techniques and procedures used to investigate reaction rates by the initial rates method and by continuous monitoring, including use of colorimetry			
	(i) for a multi-step reaction, prediction of, (i) a rate equation that is consistent with the rate-determining step (ii) possible steps in a reaction mechanism from the rate equation and the balanced equation for the overall reaction			
	(j) a qualitative explanation of the effect of temperature change on the rate of a reaction and hence the rate constant			
	(k) the Arrhenius equation: (i) the exponential relationship between the rate constant, k and temperature, T given by the Arrhenius equation, $k = Ae^{-Ea/RT}$ ii) determination of E_a and A graphically using: $\boldsymbol{\ln k = -E_a/RT + \ln A}$ derived from the Arrhenius equation.			
5.1.2 - How Far?	(a) use of the terms *mole fraction* and *partial pressure*			
	(b) calculation of quantities present at equilibrium, given appropriate data			
	(c) the techniques and procedures used to determine quantities present at equilibrium			
	(d) expressions for Kc and Kp for homogeneous and heterogeneous equilibria			
	(e) calculations of Kc and Kp, or related quantities, including determination of units			
	(f) (i) the qualitative effect on equilibrium constants of changing temperature for exothermic and endothermic reactions (ii) the constancy of equilibrium constants with changes in concentration, pressure or in the presence of a catalyst			
	(g) explanation of how an equilibrium constant controls the position of equilibrium on changing concentration, pressure and temperature			
	(h) application of the above principles in 5.1.2 How far? for Kc , Kp to other equilibrium constants, where appropriate			
5.1.3 - Acids, Bases and Buffers	(a) (i) a Brønsted–Lowry acid as a species that donates a proton and a Brønsted–Lowry base as a species that accepts a proton (ii) use of the term conjugate acid–base pairs (iii) monobasic, dibasic and tribasic acids			
	(b) the role of H+ in the reactions of acids with metals and bases (including carbonates, metal oxides and alkalis), using ionic equations			

	(c) (i) the acid dissociation constant, Ka, for the extent of acid dissociation (ii) the relationship between Ka and pKa			
	(d) use of the expression for pH as: pH = –log[H+] [H+] = 10^{-pH}			
	(e) use of the expression for ionic product of water, Kw			
	(f) calculations of pH, or related quantities, for: (i) strong monobasic acids (ii) strong bases, using Kw			
	(g) calculations of pH, Ka or related quantities, for a weak monobasic acid using approximations			
	(h) limitations of using approximations to Ka related calculations for 'stronger' weak acids			
	(i) a buffer solution as a system that minimises pH changes on addition of small amounts of an acid or a base			
	(j) formation of a buffer solution from: (i) a weak acid and a salt of the weak acid, e.g. CH_3COOH/CH_3COONa (ii) excess of a weak acid and a strong alkali, e.g. excess $CH_3COOH/NaOH$			
	(k) explanation of the role of the conjugate acid–base pair in an acid buffer solution, e.g. CH_3COOH/CH_3COO^-, in the control of pH			
	(l) calculation of the pH of a buffer solution, from the Ka value of a weak acid and the equilibrium concentrations of the conjugate acid–base pair; calculations of related quantities			
	(m) explanation of the control of blood pH by the carbonic acid–hydrogencarbonate buffer system			
	(n) pH titration curves for combinations of strong and weak acids with strong and weak bases, including: (i) sketch and interpretation of their shapes (ii) explanation of the choice of suitable indicators, given the pH range of the indicator (iii) explanation of indicator colour changes in terms of equilibrium shift between the HA and A– forms of the indicator			
	(o) the techniques and procedures used when measuring pH with a pH meter.			
5.2.1 - Lattice Enthalpy	(a) explanation of the term *lattice enthalpy* (formation of 1 mol of ionic lattice from gaseous ions, $\Delta_{LE}H$) and use as a measure of the strength of ionic bonding in a giant ionic lattice			
	(b) use of the lattice enthalpy of a simple ionic solid (e.g. NaCl, $MgCl_2$) and relevant energy terms for: (i) the construction of Born–Haber cycles (ii) related calculations			
	(c) explanation and use of the terms: (i) *enthalpy change of solution* (dissolving of 1 mol of solute, $\Delta_{sol}H$) (ii) *enthalpy change of hydration* (dissolving of 1 mol of gaseous ions in water, $\Delta_{hyd}H$)			
	(d) use of the enthalpy change of solution of a simple ionic solid (e.g. NaCl, $MgCl_2$) and relevant energy terms (*enthalpy change of hydration* and *lattice enthalpy*) for: (i) the construction of enthalpy cycles (ii) related calculations			
	(e) qualitative explanation of 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	(a) explanation that entropy is a measure of the dispersal of energy in a system which is greater, the more disordered a system			
	(b) explanation of the difference in magnitude of the entropy of a system: (i) of solids, liquids and gases (ii) for a reaction in which there is a change in the number of gaseous molecules			
	(c) calculation of the entropy change of a system, ΔS, and related quantities for a reaction given the entropies of the reactants and products			

	(d) explanation that the feasibility of a process depends upon the entropy change and temperature in the system, TΔS, and the enthalpy change of the system, ΔH			
	(e) explanation, and related calculations, of the free energy change, ΔG, as: ΔG = ΔH – TΔS (the Gibbs' equation) and that a process is feasible when ΔG has a negative value			
	(f) the limitations of predictions made by ΔG about feasibility, in terms of kinetics.			
5.2.3 - Redox and Electrode Potentials	(a) explanation and use of the terms *oxidising agent* and *reducing agent*			
	(b) construction of redox equations using half equations and oxidation numbers			
	(c) interpretation and prediction of reactions involving electron transfer			
	(d) the techniques and procedures used when carrying out redox titrations including those involving Fe^{2+}/MnO_4^- and $I_2/S_2O_3^{2-}$			
	(e) structured and non-structured titration calculations, based on experimental results of redox titrations involving: (i) Fe^{2+}/MnO_4^- and $I_2/S_2O_3^{2-}$ (ii) non-familiar redox systems			
	(f) use of the term standard electrode (redox) potential, $E^\ominus$ including its measurement using a hydrogen electrode			
	(g) the techniques and procedures used for the measurement of cell 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 in contact with a Pt electrode			
	(h) calculation of a standard cell potential by combining two standard electrode potentials			
	(i) prediction of the feasibility of a reaction using standard cell potentials and the limitations of such predictions in terms of kinetics and concentration			
	(j) application of principles of electrode potentials to modern storage cells			
	(k) explanation 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.			
5.3.1 - Transition Elements	(a) the electron configuration of atoms and ions of the d-block elements of Period 4 (Sc–Zn), given the atomic number and charge			
	(b) the elements Ti–Cu as transition elements i.e. d-block elements that have an ion with an incomplete d-sub-shell			
	(c) illustration, using at least two transition elements, of: (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 their compounds and their importance in the manufacture of chemicals by industry			
	(d) explanation and use of the term ligand in terms of coordinate (dative covalent) bonding to a metal ion or metal, including bidentate ligands			
	(e) use of the terms complex ion and coordination number and examples of complexes with: (i) six-fold coordination with an octahedral shape (ii) four-fold coordination with either a planar or tetrahedral shape			
	(f) 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. $[Ni(NH_2CH_2CH_2NH_2)_3]^{2+}$			
	(g) use of cis-platin as an anti-cancer drug and its action by binding to DNA preventing cell division			
	(h) ligand substitution reactions 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 $[Cr(H_2O)_6]^{3+}$			

	(i) explanation of the biochemical importance of iron in haemoglobin, including ligand substitution involving O_2 and CO			
	(j) reactions, including ionic equations, and the accompanying colour changes of aqueous Cu^{2+}, Fe^{2+}, Fe^{3+}, Mn^{2+} and Cr^{3+} with aqueous sodium hydroxide and aqueous ammonia, including: (i) precipitation reactions (ii) complex formation with excess aqueous sodium hydroxide and aqueous ammonia			
	(k) redox reactions and accompanying colour changes for: (i) interconversions between Fe^{2+} and Fe^{3+} (ii) interconversions between Cr^{3+} and $Cr_2O_7^{2-}$ (iii) reduction of Cu^{2+} to Cu^{+} and disproportionation of Cu^{+} to Cu^{2+} and Cu			
	(l) interpretation and prediction of unfamiliar reactions including ligand substitution, precipitation, redox.			
5.3.2 - Qualitative Analysis	(a) qualitative analysis of ions on a test-tube scale: processes and techniques needed to identify the following ions in an unknown compound: (i) anions: CO_3^{2-}, Cl^-, Br^-, I^-, SO_4^{2-} (ii) cations: NH_4^+, Cu^{2+}, Fe^{2+}, Fe^{3+}, Mn^{2+}, Cr^{3+}			
6.1.1 - Aromatic Compounds	(a) the comparison of the Kekulé model of benzene with the subsequent delocalised models for benzene in terms of p-orbital overlap forming a delocalised π-system			
	(b) the experimental evidence for a delocalised, rather than Kekulé, model for benzene in terms of bond lengths, enthalpy change of hydrogenation and resistance to reaction			
	(c) use of IUPAC rules of nomenclature for systematically naming substituted aromatic compounds			
	(d) 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			
	(e) the mechanism of electrophilic substitution in arenes for nitration and halogenation			
	(f) the explanation of 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			
	(g) the interpretation of unfamiliar electrophilic substitution reactions of aromatic compounds, including prediction of mechanisms			
	(h) the weak acidity of phenols shown by the neutralisation reaction with NaOH but absence of reaction with carbonates			
	(i) the electrophilic substitution reactions of phenol: (i) with bromine to form 2,4,6-tribromophenol (ii) with dilute nitric acid to form 2-nitrophenol			
	(j) the relative ease of electrophilic substitution of phenol compared with benzene, in terms of electron pair donation to the π-system from an oxygen p-orbital in phenol			
	(k) 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			
	(l) the prediction of substitution products of aromatic compounds by directing effects and the importance to organic synthesis.			
6.1.2 - Carbonyl Compounds	(a) oxidation of aldehydes using $Cr_2O_7^{2-}/H^+$ (i.e. $K_2Cr_2O_7/H_2SO_4$) to form carboxylic acids			
	(b) nucleophilic addition reactions of carbonyl compounds with: (i) $NaBH_4$ to form alcohols (ii) HCN [i.e. $NaCN_{(aq)}/H+_{(aq)}$], to form hydroxynitriles			
	(c) the mechanism for nucleophilic addition reactions of aldehydes and ketones with $NaBH_4$ and HCN			
	(d) 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			
	(e) 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	(a) explanation of the water solubility of carboxylic acids in terms of hydrogen bonding			
	(b) reactions in aqueous conditions of carboxylic acids with metals and bases (including carbonates, metal oxides and alkalis)			
	(c) esterification of: (i) carboxylic acids with alcohols in the presence of an acid catalyst (e.g. conc H_2SO_4) (ii) acid anhydrides with alcohols			
	(d) 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			
	(e) the formation of acyl chlorides from carboxylic acids using $SOCl_2$			
	(f) use of acyl chlorides in synthesis in formation of esters, carboxylic acids and primary and secondary amides.			
6.2.1 - Amines	(a) the basicity of amines in terms of proton acceptance by the nitrogen lone pair and the reactions of amines with dilute acids, e.g. HCl(aq), to form salts			
	(b) the preparation of: (i) aliphatic amines by substitution of haloalkanes with excess ethanolic ammonia and amines (ii) aromatic amines by reduction of nitroarenes using tin and concentrated hydrochloric acid.			
6.2.2 - Amino Acids, Amides and Chirality	(a) the general formula for an α-amino acid as $RCH(NH_2)COOH$ and the following reactions of amino acids: (i) reaction of the carboxylic acid group with alkalis and in the formation of esters (ii) reaction of the amine group with acids			
	(b) structures of primary and secondary amides			
	(c) optical isomerism (an example of stereoisomerism, in terms of non-superimposable mirror images about a chiral centre)			
	(d) identification of chiral centres in a molecule of any organic compound.			
6.2.3 - Polyesters and Polyamides	(a) condensation polymerisation to form: (i) polyesters (ii) polyamides			
	(b) the acid and base hydrolysis of: (i) the ester groups in polyesters (ii) the amide groups in polyamides			
	(c) prediction from addition and condensation polymerisation of: (i) the repeat unit from a given monomer(s) (ii) the monomer(s) required for a given section of a polymer molecule (iii) the type of polymerisation.			
6.2.4 - Carbon-carbon bond formation	(a) the use of C–C bond formation in synthesis to increase the length of a carbon chain			
	(b) formation of C-CN by reaction of: (i) haloalkanes with CN^- and ethanol, including nucleophilic substitution mechanism (ii) carbonyl compounds with HCN, including nucleophilic addition mechanism			
	(c) reaction of nitriles from (b): (i) by reduction (e.g. with H_2/Ni) to form amines (ii) by acid hydrolysis to form carboxylic acids			
	(d) 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	(a) 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 • recrystallization • measurement of melting points			
	(b) for an organic molecule containing several functional groups: (i) identification of individual functional groups (ii) prediction of properties and reactions			
	(c) multi-stage synthetic routes for preparing organic compounds.			
6.3.1 - Chromatography and Qualitative Analysis	(a) interpretation of one-way TLC chromatograms in terms of Rf values			
	(b) interpretation of gas chromatograms in terms of: (i) retention times (ii) the amounts and proportions of the components in a mixture.			
	(c) 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 carboxylic acids by reaction with CO_3^{2-}			
6.3.2 - Spectroscopy	(a) 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			
	(b) 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			
	(c) prediction of a carbon-13 or proton NMR spectrum for a given molecule			
	(d) (i) 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			
	(e) deduction of the structures of organic compounds from different analytical data including: (i) elemental analysis (ii) mass spectra (iii) IR spectra (iv) NMR spectra.			