

OCR AS (3882) and A (7882) GCE Chemistry

Modules 2811, 2812 and 2813 constitute the 'AS' course. Modules 2811, 2812, 2813, 2814, 2815 and 2816 constitute the 'A' course.

Syllabus reference	Syllabus content	Page reference in <i>Chemistry in Context 5th Edn</i>
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5.1 Module 2811: Foundation Chemistry

5.1.1 Atoms, Molecules and Stoichiometry

(a)	Define the terms 'relative atomic', 'isotopic', 'molecular' and 'formula masses', based on the ^{12}C scale.	4
(b)	Describe the basic principles of the mass spectrometer limited to ionisation, acceleration, deflection and detection.	2–4
(c)	Outline the use of mass spectrometry (i) in the determination of relative isotopic masses; (ii) as a method for identifying elements, for example; use in Mars space probe.	4
(d)	Interpret mass spectra in terms of isotopic abundances.	4
(e)	Calculate the relative atomic mass of an element given the relative abundances of its isotopes, or its mass spectrum.	4
(f)	Define the mole in terms of the Avogadro constant; molar mass as the mass of 1 mole of a substance.	5
(g)	Define the terms 'empirical formula' and 'molecular formula'.	6
(h)	Calculate empirical and molecular formulae using composition by mass.	6
(i)	Construct balanced chemical equations (full and ionic).	14, 17–18
(j)	Perform calculations (including use of the Mole Concept, formulae and equations) involving (i) reacting masses; (ii) volumes of gases; (iii) volumes and concentrations of solutions in simple acid–base titrations.	15–16 12–14 7, 11
(k)	Deduce stoichiometric relationships from calculations such as those in (j)	11–14

5.1.2 Atomic Structure

(a)	Recognise and describe protons, neutrons and electrons in terms of relative charge and relative mass.	60
(b)	Describe the distribution of mass and charge within an atom.	60
(c)	Describe the contribution of protons and neutrons to the nucleus of an atom, in terms of atomic number and mass number.	57, 60

(d)	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.	60
(e)	Distinguish between the isotopes of an element in terms of their different masses and different numbers of neutrons.	60–61
(f)	Explain the terms first ionisation energy and successive ionisation energy of an element in terms of 1 mole of gaseous atoms or ions (see also 5.1.4(e), (f)).	67
(g)	Explain that ionisation energies are influenced by nuclear charge, electron shielding and the distance of the outermost electron from the nucleus.	173–175
(h)	Predict the number of electrons in each principal quantum shell of an element from its successive ionisation energies.	68–69
(i)	Describe the shapes of s- and p- orbitals.	72
(j)	Describe the numbers and relative energies of s-, p- and d- orbitals for the principal quantum numbers 1, 2, 3 and also the 4s- and 4p- orbitals.	70–72
(k)	Deduce the electronic configuration 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 $Z=36$.	73–74

5.1.3 Chemical Bonding and Structure

(a)	Describe ionic bonding as the electrostatic attraction between two oppositely-charged ions.	84–85
(b)	Describe, including the use of 'dot-and-cross' diagrams, ionic bonding, for example, as in sodium chloride and magnesium oxide.	84–85
(c)	Describe, in simple terms, the lattice structure of sodium chloride.	125
(d)	Describe a covalent bond as a shared pair of electrons.	86
(e)	Describe, including the use of 'dot-and-cross' diagrams, (i) covalent bonding, for example, as in hydrogen, chlorine, oxygen, hydrogen chloride, water, ammonia, methane, carbon dioxide and ethene; (ii) dative covalent (co-ordinate) bonding, for example, as in the ammonium ion.	86–90
(f)	Explain the shapes of, and bond angles in, molecules and ions by using the qualitative model of electron-pair repulsion for up to 4 electron pairs (including lone pairs), for example, as in BF_3 (trigonal), CO_2 (linear), CH_4 and NH_4^+ (tetrahedral), NH_3 (pyramidal) and H_2O (non-linear).	90–92
(g)	Predict the shapes of, and bond angles in, molecules and ions analogous to those specified in (f).	92
(h)	Appreciate that, between the extremes of ionic and covalent bonding, there is a gradual transition from one extreme to the other.	98
(i)	Describe electronegativity as the ability of an atom to attract the bonding electrons in a covalent bond.	97–98
(j)	Explain that: (i) bond polarity may arise when covalently-bonded atoms have different electronegativities; (ii) polarisation may occur between cations of high charge density and anions of low charge density.	97–98
(k)	Describe intermolecular forces based on permanent dipoles, as in hydrogen chloride, and instantaneous dipoles (van der Waals' forces), as in the noble gases.	99–103
(l)	Describe hydrogen bonding between molecules containing —OH and —NH groups, typified by water and ammonia.	103–105
(m)	Describe and explain the anomalous properties of water resulting from hydrogen bonding for example: (i) the density of ice compared with water; (ii) its relatively high freezing point and boiling point.	106–108
(n)	Describe, in simple terms, the giant molecular structures of diamond and graphite.	123–124
(o)	Describe metallic bonding, present in a giant metallic lattice structure, as the attraction of a lattice of positive ions to a sea of mobile electrons.	45–46

(p)	Describe, interpret and/or predict physical properties, for example: melting and boiling points, electrical conductivity and solubility in terms of (i) the types, motion and arrangement of particles (atoms, molecules and ions) and the forces between them; (ii) the different types of bonding (ionic bonding, covalent bonding, hydrogen bonding, other intermolecular interactions, metallic bonding).	126–128 130–132
(q)	Deduce the type of bonding present from given information.	131–132

5.1.4 The Periodic Table: Introduction

(a)	Describe the Periodic Table in terms of the arrangement of elements (i) by increasing atomic number; (ii) in periods showing repeating trends in physical and chemical properties; (iii) in groups having similar physical and chemical properties.	35–45
(b)	Describe, for the elements of Period 3, the variation in electronic configurations, atomic radii, electrical conductivities, melting points and boiling points.	42–44
(c)	Explain variations in (b) in terms of the structure and bonding of the elements.	169, 170–177
(d)	Classify the elements in s-, p- and d- blocks.	39–41
(e)	Interpret successive ionisation energies of an element in terms of its position in the Periodic Table (see also 5.1.2(f)–(h)).	67
(f)	Describe and explain the variation of the first ionisation energies of elements shown by (i) a decrease down a group in terms of increasing atomic radius and electron shielding; (ii) a general increase across a period, in terms of increasing nuclear charge; (iii) the periodic decrease between Groups 2 and 3, in terms of the higher energy level of the p sub-shell compared with that of the s sub-shell; (iv) the periodic decrease between Groups 5 and 6, in terms of an increase in energy from mutual repulsion of paired electrons in a Group 6 p-orbital.	} 67–71, 173–175
(g)	Interpret data on electronic configurations, atomic radii, electrical conductivities, first ionisation energies, melting points and boiling points to demonstrate periodicity.	36, 169–177

5.1.5 The Periodic Table: The Group 2 elements and their compounds

(a)	Describe and explain the trends in electronic configurations, atomic radii and ionisation energies of the Group 2 elements, Mg to Ba.	212
(b)	Use the rules for assigning oxidation state (number) with elements, compounds and ions.	211, 29–32
(c)	Describe oxidation and reduction in terms of (i) electron transfer; (ii) changes in oxidation state.	20–21 32
(d)	Describe the redox reactions of the elements (Mg to Ba) with oxygen and with water and explain the trend in reactivity in terms of ionisation energies.	213
(e)	Describe the reactions of Mg, MgO and MgCO ₃ with hydrochloric acid (see also 5.3.3(f), (g)).	
(f)	Describe the thermal decomposition of CaCO ₃ (limestone) to form CaO (lime) and the subsequent formation of Ca(OH) ₂ (slaked lime) with water.	214, 292
(g)	Describe lime water as an aqueous solution of Ca(OH) ₂ and state its appropriate pH.	
(h)	Describe the reaction of lime water (i) with carbon dioxide forming CaCO ₃ (s); (ii) with excess carbon dioxide, forming Ca(HCO ₃) ₂ (aq), as a hard water.	
(i)	Interpret and make predictions from the chemical and physical properties of the Group 2 elements and their compounds.	213–215
(j)	Show awareness of the importance and use of Group 2 elements and their compounds, with appropriate chemical explanations, for example: 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.	220

5.1.6 The Periodic Table: The Group 7 elements and their compounds

(a)	Explain the trend in the volatilities of chlorine, bromine and iodine in terms of van der Waals' forces	102, 228
(b)	Describe the relative reactivity of the elements Cl ₂ , Br ₂ and I ₂ in displacement reactions.	230–231
(c)	Explain the trend in (b) in terms of oxidising power, i.e. the relative ease with which an electron can be captured.	231
(d)	Describe the characteristic reactions of the ions Cl ⁻ , Br ⁻ and I ⁻ with aqueous silver ions followed by aqueous ammonia (knowledge of complex formulae not required).	233–234
(e)	Describe and interpret, in terms of changes in oxidation number, (i) the reaction of chlorine with water, as used in water purification to prevent life-threatening diseases; (ii) the reaction of chlorine with cold, dilute aqueous sodium hydroxide, as used to form bleach.	231 231–233

5.2 Module 2812: Chains and Rings

5.2.1 Basic Concepts

(a)	Interpret and use the terms: nomenclature, molecular formula, general formula, structural formula, displayed formula, skeletal formula, homologous series and functional group;	380
(b)	Describe and explain (i) structural isomerism in compounds with the same molecular formula but different structural formulae; (ii) <i>cis-trans</i> isomerism in alkenes, in terms of restricted rotation about a double bond.	391–392 419–420
(c)	Determine the possible structural and/or <i>cis-trans</i> isomers of an organic molecule of given molecular formula.	419–420
(d)	Perform calculations, involving use of the Mole Concept and reacting quantities, to determine the percentage yield of a reaction (see also 5.1.1(j),(k)).	532–534

5.2.2 Hydrocarbons: Alkanes

(a)	State that alkanes are saturated hydrocarbons.	400
(b)	Explain, in terms of van der Waals' forces, the variations in boiling points of alkanes with different carbon chain length and branching.	402
(c)	Describe the lack of reactivity of alkanes, in terms of the non-polarity of C—H bonds.	407
(d)	Describe the chemistry of alkanes, typified by the following reactions of methane: (i) combustion (see also, 5.2.3(c)); (ii) substitution by chlorine and by bromine to form halogenoalkanes.	408 407, 408, 409
(e)	Describe how homolytic fission leads to the mechanism of free-radical substitution in alkanes, typified by methane and chlorine, in terms of initiation, propagation and termination reactions.	404, 407–408

5.2.3 Hydrocarbons: Fuels

(a)	Explain the use of crude oil as a source of hydrocarbons (separated by fractional distillation) which can be used directly as fuels or for processing into petrochemicals.	397–400 410–411
(b)	Describe the use of (i) cracking to obtain more useful alkanes and alkenes; (ii) isomerisation to obtain branched alkanes; (iii) reforming to obtain cycloalkanes and arenes.	410–411 410 440
(c)	Describe and explain how the combustion reactions of alkanes (see also, 5.2.2(d)) lead to their use as fuels in industry, in the home and in transport.	408
(d)	State that branched alkanes, cycloalkanes and arenes are used in petrol to promote efficient combustion (see also 5.2.5(g), alcohols as fuels; 5.3.2(i)–(k), catalytic converters),	409, 412, 413

(e)	Outline	} 164–166
	(i) the value to society of fossil fuels in relation to needs for energy and raw materials;	
	(ii) the non-renewable nature of fossil fuel reserves;	
	(iii) the need to develop renewable fuels, for example biofuels, which do not further deplete finite energy resources.	

5.2.4 Hydrocarbons: Alkenes

(a)	State that alkenes are unsaturated hydrocarbons.	418
(b)	State and explain the bonding in alkenes in terms of the overlap of adjacent p-orbitals to form a π -bond.	419
(c)	State and explain the shape of ethene and other related molecules (see also 5.1.3(f)).	419
(d)	Describe the chemistry of alkenes, for example, by the addition reactions of ethene and propene with:	421,
	(i) hydrogen in the presence of a suitable catalyst, for example Ni, to form alkanes;	422
	(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;	422
	(iii) hydrogen halides to form halogenoalkanes;	423,
	(iv) steam in the presence of an acid catalyst, for example H_3PO_4 to form alcohols (see also 5.2.5(b)).	424–425
(e)	Define the term electrophile as an electron pair acceptor.	405
(f)	Describe how, heterolytic fission leads to the mechanism of electrophilic addition in alkenes, typified by bromine and ethene to form 1, 2-dibromoethane.	421, 422
(g)	Describe the addition polymerisation of alkenes, for example: ethene and propene.	428–430
(h)	Deduce the repeat unit of an addition polymer obtained from a given monomer.	431
(i)	Identify, in a given section of an addition polymer, the monomer from which it was obtained.	431
(j)	Outline the use of alkenes in the industrial production of organic compounds, typified by:	
	(i) the manufacture of margarine by catalytic hydrogenation of unsaturated vegetable oils using hydrogen and a nickel catalyst;	422,
	(ii) the formation of a range of polymers using unsaturated monomer units based on the ethene molecule, for example, CH_2CHCl , CF_2CF_2 (see also, 5.2.6(f)).	431
(k)	Outline the difficulties in disposing of polymers, for example: non-biodegradability or toxic combustion products.	286, 426

5.2.5 Alcohols

(a)	Explain in terms of hydrogen bonding, the water solubility and the relatively low volatility of alcohols.	468–469
(b)	Describe and explain the industrial production of ethanol by	
	(i) fermentation from sugars, for example glucose;	466,
	(ii) the reaction of steam with ethene in the presence of H_3PO_4 (see also 5.2.4(d)).	425
(c)	Describe the classification of alcohols into primary, secondary and tertiary alcohols.	467
(d)	Describe the chemistry of alcohols, typified by the following reactions of ethanol:	
	(i) combustion,	476,
	(ii) substitution using HBr (e.g. $\text{NaBr}/\text{H}_2\text{SO}_4$ to form a bromoalkane;	472–473,
	(iii) reaction with sodium to form a sodium alkoxide and hydrogen;	471,
	(iv) dehydration with hot, concentrated sulphuric acid or hot pumice/ Al_2O_3 to form an alkene;	473–474,
	(v) esterification with carboxylic acids in the presence of an acid catalyst.	472

(e)	Describe the action of $\text{Cr}_2\text{O}_7^{2-}/\text{H}^+$ (e.g. $\text{K}_2\text{Cr}_2\text{O}_7/\text{H}_2\text{SO}_4$) on alcohols, typified by (i) the oxidation of primary alcohols to form aldehydes and carboxylic acids, and 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.	475–476
(f)	Identify, using an infra-red spectrum, (i) an alcohol from absorption of the O—H bond; (ii) a carbonyl compound from absorption of the C=O bond; (iii) a carboxylic acid from absorption of the C=O bond and broad absorption of the O=H bond.	386
(g)	Outline the use of (i) ethanol in alcoholic drinks, as a solvent in the form of methylated spirits, and as a fuel, particularly as a petrol substitute in countries with limited oil reserves; (ii) methanol as a petrol additive to improve combustion and its increasing importance as a feedstock in the production of organic chemicals. (See also, 5.2.3(d), petrol as a fuel).	466, 476, 476

5.2.6 Halogenoalkanes

(a)	Describe substitution reactions of halogenoalkanes, typified by the following reactions of bromoethane: (i) hydrolysis with hot aqueous alkali to form alcohols (ii) reaction with excess ethanolic ammonia to form primary amines.	456, 461–462
(b)	Define the term nucleophile as an electron pair donor.	405
(c)	Describe the mechanism of nucleophilic substitution in the hydrolysis of primary halogenoalkanes.	455
(d)	Explain the rates of hydrolysis of primary halogenoalkanes in terms of the bond enthalpies of carbon-halogen bonds (C—F, C—Cl, C—Br and C—I). (See also 5.3.1(f).)	456
(e)	Describe the elimination of hydrogen bromide from halogenoalkanes, typified by bromoethane, with hot ethanolic sodium hydroxide.	462
(f)	Outline the uses of (i) fluoroalkanes and fluorohalogenoalkanes, for example: chlorofluorocarbons, CFCs (refrigerants, propellants, blowing polystyrene, dry cleaning, degreasing agents); (ii) chloroethene and tetrafluoroethene to produce the plastics <i>pvc</i> and <i>ptfe</i> . (See also 5.2.4(g)–(m) and 5.4.6(a).) (iii) halogenoalkanes as synthetic intermediates in chemistry.	453, 430–431, 459
(g)	Outline the role of chemists in minimising damage to the environment by, for example, the development of alternatives to CFCs so that depletion of the ozone layer (see also 5.3.2(i), (1)) can be reversed.	454

5.3 Module 2813, Component 01: How Far, How Fast?

5.3.1 Enthalpy changes

(a)	Explain that some chemical reactions are accompanied by enthalpy changes, principally in the form of heat energy; the enthalpy changes can be exothermic (ΔH , negative) or endothermic (ΔH , positive).	146–147
(b)	Recognise the importance of oxidation as an exothermic process, for example, in the combustion of fuels and the oxidation of carbohydrates such as glucose in respiration.	110
(c)	Recognise that endothermic processes require an input of heat energy, for example, the thermal decomposition of calcium carbonate (see also 5.1.5(f)) and in photosynthesis.	355
(d)	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.	147
(e)	Explain chemical reactions in terms of enthalpy changes associated with the breaking and making of chemical bonds.	155–157

(f)	Explain and use the terms: (i) <i>enthalpy change of reaction and standard conditions</i> , with particular reference to formation and combustion; (ii) <i>average bond enthalpy</i> (ΔH positive; bond breaking of 1 mole of bonds).	148–150
(g)	Calculate enthalpy changes from appropriate experimental results directly, including the use of the relationship; energy change = $mc\Delta T$.	147–148
(h)	Use Hess's Law to construct enthalpy cycles and carry out calculations using such cycles and relevant enthalpy terms, with particular reference to enthalpy changes that cannot be found by direct experiment, for example: (i) an enthalpy change of formation from enthalpy changes of combustion; (ii) an enthalpy change of reaction from enthalpy changes of formation; (iii) an enthalpy change of reaction from average bond enthalpies.	150–155, 151, 152–153, 156–160

5.3.2 Reaction rates

(a)	Describe qualitatively, in terms of collision theory, the effect of concentration changes on the rate of a reaction.	354
(b)	Explain why an increase in the pressure of a gas, increasing its concentration, may increase the rate of a reaction involving gases.	
(c)	Explain qualitatively, using the Boltzmann distribution and enthalpy profiles diagrams, what is meant by the term <i>activation energy</i> .	366–367
(d)	Describe qualitatively, using the Boltzmann distribution, the effect of temperature changes on the rate of a reaction.	369
(e)	Explain what is meant by a <i>catalyst</i> .	369–371
(f)	Describe catalysts as having great economic importance, for example: in fertiliser production (see also 5.3.3(c), (d)), petroleum processing (see also 5.2.3(b)) and margarine production (see also 5.2.4(j)).	314–315, 410, 422
(g)	Explain that, in the presence of a catalyst, a reaction proceeds via a different route, i.e. one of lower activation energy, giving rise to an increased reaction rate.	371
(h)	Interpret the catalytic behaviour in (g) in terms of the Boltzmann distribution and enthalpy profile diagrams.	371
(i)	State what is meant by (i) homogeneous catalysis, for example: $\text{H}^+(\text{aq})$ in esterification (see also 5.2.5(d)) and chlorine free radicals with ozone (see also (l) below); (ii) heterogeneous catalysis, for example: Fe in the Haber process (see also 5.3.3(c)) and Rh/Pt/Pd in catalytic converters (see also (k) below).	371, 453–454, 314–315
(j)	For carbon monoxide, oxides of nitrogen and unburnt hydrocarbons; describe their presence and/or formation from the internal combustion engine; state their environmental consequences in terms of low-level ozone and photochemical smog (equations not required).	411–412
(k)	Outline, as an example of heterogeneous catalysis, how a catalytic converter decreases carbon monoxide and nitrogen monoxide emissions from internal combustion engines (see also 5.3.2(i); 5.7.1(h)) by: (i) adsorption of carbon monoxide and nitrogen monoxide to the catalyst surface; (ii) chemical reaction; (iii) subsequent desorption of carbon dioxide and nitrogen from the catalyst surface.	413
(l)	Outline, as an example of homogeneous catalysis, how gaseous chlorine free radicals, formed by the action of ultraviolet radiation on CFCs, catalyse the breakdown of the gaseous ozone layer into oxygen (see also 5.2.2(g); 5.2.6(g); 5.7.1(e)) by a reaction route via ClO radicals (as the intermediate), for example: $\text{Cl} + \text{O}_3 \rightarrow \text{ClO} + \text{O}_2$; $\text{ClO} + \text{O} \rightarrow \text{Cl} + \text{O}_2$.	453–454, 453–454

5.3.3 Chemical Equilibrium

(a)	Explain the features of a <i>dynamic equilibrium</i> .	292
(b)	State Le Chatelier's principle and apply it to deduce qualitatively (from appropriate information) the effect of a change in temperature, concentration or pressure, on a homogeneous system in equilibrium.	305–309

(c)	Describe and explain the conditions used in the Haber process for the formation of ammonia, as an example of the importance of a compromise between chemical equilibrium and reaction rate in the chemical industry.	312–315
(d)	Outline the importance of ammonia and nitrogen compounds derived from ammonia, for example, fertilisers, polyamides and explosives.	315–318
(e)	Describe an acid as a species that can donate a proton.	198–199
(f)	Describe the reactions of an acid, typified by hydrochloric acid with metals, carbonates, bases and alkalis (see also 5.1.5(e)).	17–18, 199–200
(g)	Interpret the reactions in (f) using ionic equations to emphasise the role of $H^+(aq)$.	17–18
(h)	Explain qualitatively, in terms of dissociation, the differences between strong and weak acids.	200
(i)	Describe ammonia as a base, in terms of its reaction with an acid (e.g. sulphuric acid) to form ammonium salts, used in fertilisers.	199

5.4 Module 2814: Chains, Rings and Spectroscopy

5.4.1 Arenes

(a)	Show understanding of the concept of delocalisation of electrons as used in a model of benzene.	159–160, 436–438
(b)	Describe the electrophilic substitution of arenes with: (i) concentrated nitric acid in the presence of sulphuric acid; (ii) a halogen in the presence of a halogen carrier; (iii) a halogenoalkane such as chloromethane in the presence of a halogen carrier; (Friedel-Crafts reaction).	442–445
(c)	Describe the mechanism of electrophilic substitution in arenes, using the mononitration of benzene as an example.	441–442
(d)	Understand that reactions of arenes, such as those in (b), are used by industry during the synthesis of commercially important materials, for example: explosives, pharmaceuticals and dyes (from nitration) – see also 5.4.4 (d), and polymers such as polystyrene (from alkylation)- see also 5.4.6 (a).	448, 522, 444
(e)	Explain the relative resistance to bromination of benzene compared with cyclohexene (see also 5.2.4(d), in terms of delocalisation of the benzene ring.	440–441
(f)	Describe the reactions of phenol (i) with bases and with sodium to form salts; (ii) with bromine to form 2,4,6-tribromophenol.	471, 476–477
(g)	Explain the relative ease of bromination of phenol compared with benzene, in terms of activation of the benzene ring.	477
(h)	State the uses of phenols in antiseptics and disinfectants.	471

5.4.2 Carbonyl Compounds

(a)	Describe the reduction of carbonyl compounds using $NaBH_4$ to form alcohols.	486
(b)	Describe the mechanism for nucleophilic addition reactions of hydrogen cyanide (in the presence of potassium cyanide) with aldehydes and ketones. (see also 5.4.5(c)).	484–485
(c)	Describe the use of 2,4-dinitrophenylhydrazine (i) to detect the presence of a carbonyl group in an organic compound; (ii) to identify a carbonyl compound from the melting point of the derivative.	487
(d)	Describe the use of Tollen's reagent (ammoniacal silver nitrate) (i) to detect the presence of an aldehyde group; (ii) to distinguish between aldehydes and ketones, explained in terms of the oxidation of aldehydes to carboxylic acids with reduction of silver ions to silver.	487–488, 488

5.4.3 Carboxylic Acids and Esters

(a)	Describe carboxylic acids as proton donors.	500
(b)	Describe the reactions of carboxylic acids, typified by ethanoic acid; (i) With aqueous alkalis to form carboxylates (salts); (ii) With alcohols, in the presence of an acid catalyst, to form esters (see also 5.2.5(d)).	500–501, 505–506
(c)	State the uses of esters in perfumes and flavourings.	507
(d)	Describe the acid and base hydrolysis of esters to form carboxylic acids and carboxylates respectively.	508, 509–510

5.4.4 Nitrogen Compounds

(a)	Describe the formation of phenylamine by reduction of nitrobenzene using tin and concentrated hydrochloric acid.	518
(b)	Describe the reactions of primary amines with acids to form salts.	519
(c)	Explain the basicity of primary amines and the relative basicities of ethylamine and phenylamine in terms of the inductive effect and the influence of the delocalised electrons in the benzene ring.	520
(d)	Describe the synthesis of an azo-dye by reaction of phenylamine with nitrous acid (HNO_2/HCl , $<10^\circ\text{C}$) with the formation of a diazonium salt, followed by coupling with phenols under alkaline conditions; the use of such reactions in the formation of dyestuffs.	521, 522–524
(e)	State the general formula for an α -amino acid as $\text{RCH}(\text{NH}_2)\text{COOH}$.	524
(f)	Describe the acid-base properties of α -amino acids and the formation of zwitterions.	524
(g)	Explain the formation of a peptide linkage between α -amino acids leading to the idea that polypeptides and proteins are condensation polymers (see also Section 5.4.6 (b)).	524–525
(h)	Describe the acid hydrolysis, for example with hot $\text{HCl}(\text{aq})$, of proteins and peptides to form α -amino acids.	525

5.4.5 Stereoisomerism and organic synthesis

(a)	Interpret and use the term stereoisomerism in terms of <i>cis-trans</i> and optical isomerism.	392–393, 420
(b)	Explain the term chiral centre and identify any chiral centres in a molecule of given structural formula (for example, amino acids (see also 5.4.4(e)) and 2-hydroxypropanoic acid (lactic acid)).	392–393
(c)	Describe the two-stage synthesis of hydroxypropanoic acid (lactic acid) by the addition of hydrogen cyanide to ethanal (see also 5.4.2(b)) followed by acid hydrolysis; explain the use of such reactions in synthesis by providing a route for lengthening a carbon chain.	484, 531, 499
(d)	Understand that chiral molecules prepared synthetically in the laboratory may contain a mixture of optical isomers, whereas molecules of the same compound produced naturally in living systems will often be present as one optical isomer only (for example: L-amino acids).	393
(e)	Understand that the synthesis of pharmaceuticals often requires the production of chiral drugs containing a single optical isomer, resulting in smaller doses (only half the drug is needed), reduced side effects and improved pharmacological activity.	393

5.4.6 Polymerisation

(a)	Describe the characteristics of an addition polymerisation, typified by poly(phenylethene) (see also 5.2.4 (g)–(m), 5.2.6(f)).	428–431
(b)	Identify that some alkenes, typified by propene, can produce addition polymers that are atactic, isotactic and/or syndiotactic.	
(c)	Describe the characteristics of condensation polymerisation: (i) in polyesters, typified by <i>Terylene</i> (from benzene-1,4-dicarboxylic acid and ethane-1,2-diol); (ii) in polyamides, typified by nylon-6, 6 (from 1-6-diaminohexane and hexane-1,6-dicarboxylic acid) and Kevlar (from benzene-1,4 diamine and benzene-1-4-dicarboxylic acid); (iii) in polypeptides and proteins (see also Section 5.4.4(g)).	428, 507, 525–526, 109, 524

(d)	Suggest the type of polymerisation reaction from (i) a given monomer or pair of monomers; (ii) a given section of a polymer molecule.	426–431
(e)	Deduce the repeat unit of a polymer obtained from a given monomer or pair of monomers.	426–431
(f)	Identify, in a given section of polymer, the monomer(s) from which it was obtained.	426–431
(g)	State the use of polyesters and polyamides as fibres in clothing.	526, 507

5.4.7 Spectroscopy

(a)	Use a simple infra-red spectrum to identify the presence of functional groups in a molecule (limited to alcohols (O—H), carbonyl compounds C=O), carboxylic acids (COOH) and esters (COOR) see also 5.2.5(f)).	385–386
(b)	Use the molecular ion peak in a mass spectrum to determine the relative molecular mass of an organic molecule.	383
(c)	Predict, from the high resolution n.m.r. spectrum of a simple molecule containing carbon, hydrogen and/or oxygen, (i) the different types of proton present from chemical shift values; (ii) the relative numbers of each type of proton present from the relative peak area; (iii) the number of protons adjacent to a given proton from the spin-spin splitting pattern, limited to splitting patterns up to a quadruplet only; (iv) possible structures for the molecule.	387–390
(d)	Predict the chemical shifts and splitting patterns of the protons in a given molecule. <ul style="list-style-type: none"> ● Background theory will not be tested on examination papers: the emphasis is on the interpretation of spectra. Thus, candidates will not be tested on why nuclear magnetic resonance takes place, the reasons for difference chemical shift values or why spin-spin splitting occurs. ● The relative peak areas will be given on any provided spectra. ● For splitting patterns, the $n+1$ rule can be used, where n is the number of H atoms on adjacent carbon atoms. Limited to singlet, doublet, triplet and quadruplet. 	387–390
(e)	Describe the use of D ₂ O to identify the n.m.r signal from —OH groups.	

5.5 Module 2815, Component 01: Trends and Patterns

5.5.1 Lattice enthalpy

(a)	Explain and use the term <i>lattice enthalpy</i> (ΔH negative, i.e. gaseous ions to the solid lattice) as a measure of ionic bond strength.	86, 160
(b)	Construct Born-Haber cycles to calculate the lattice enthalpy of a simple ionic solid (e.g. NaCl, MgCl ₂) using relevant energy terms (enthalpy change of formation, ionisation energy, enthalpy change of atomisation and electron affinity).	161
(c)	Explain, in qualitative terms, the effect of ionic charge and of ionic radius on the numerical magnitude of a lattice enthalpy.	161–162
(d)	Describe the thermal decomposition of the Group 2 carbonates (MgCO ₃ to BaCO ₃) and explain the trend in terms of lattice enthalpy, the polarisation of the anions and the charge density of the cations (see also 5.1.3 (j), 5.1.5 (f)).	213–214
(e)	Relate the high lattice enthalpy of MgO to its use as a refractory lining.	161

5.5.2 Periodic Table: Period 3

(a)	Describe redox reactions of Period 3 elements (i) with oxygen to give MgO, Al ₂ O ₃ and SO ₂ ; (ii) with chlorine to give NaCl, MgCl ₂ , AlCl ₃ , SiCl ₄ , PCl ₅ ; (iii) with water to give NaOH and Mg(OH) ₂ .	47, 47, 213
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(b)	Describe the action of water, if any, on the compounds in (a) and the pH of any resulting solutions.	179, 180, 180–182
(c)	Explain the trends in (a) and (b) above in terms of the structure and bonding of the oxides and chlorides involved.	47–49 179–184

5.5.3 Periodic Table: Transition Elements

(a)	Describe a <i>transition element</i> as a d-block element forming one or more stable ions with incompletely filled d-orbitals.	249–251
(b)	Deduce the electronic configuration of atoms and ions of the d-block elements of Period 4 (Sc → Cu), given the atomic number and charge.	249
(c)	Illustrate, using the transition elements iron and copper as appropriate, (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.	254–262, 254–257, 269–270 259–261 261–262
(d)	Describe the simple precipitation reactions and the accompanying colour changes of $\text{Cu}^{2+}(\text{aq})$, $\text{Fe}^{2+}(\text{aq})$ and $\text{Fe}^{3+}(\text{aq})$ with aqueous sodium hydroxide to form $\text{Cu}(\text{OH})_2$, $\text{Fe}(\text{OH})_2$ and $\text{Fe}(\text{OH})_3$ respectively.	270, 283–284
(e)	Explain and use the terms <i>complex ion</i> and <i>ligand</i> in terms of co-ordinate bonding.	201–202, 257
(f)	Describe the process of ligand substitution and the accompanying colour changes, for example in the formation of: (i) $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$ and CuCl_4^{2-} from $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$; (ii) $[\text{Fe}(\text{SCN})(\text{H}_2\text{O})_5]^{2+}$ from $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$	203–204, 257 269, 305
(g)	Predict from data, the formula and possible shape of a complex ion, limited to tetrahedral and octahedral complexes.	257–258
(h)	Explain how colorimetry can be used to determine the formula of a complex ion, for example $[\text{Fe}(\text{SCN})(\text{H}_2\text{O})_5]^{2+}$.	204
(i)	Predict the colour of a transition element complex from its ultraviolet/visible spectrum.	259
(j)	Describe redox behaviour in transition elements, for example by $\text{Fe}^{3+}/\text{Fe}^{2+}$ and $\text{MnO}_4^-/\text{Mn}^{2+}$.	31–32
(k)	Construct redox equation, such as those in (j) above, using relevant half-equations.	23–24
(l)	Perform calculations involving simple redox titrations, for example $\text{MnO}_4^-/\text{Fe}^{2+}$ in acid solution.	207–208

5.11 Module 2816, Component 01: Unifying Concepts in Chemistry

5.11.1 How Fast?

(a)	Explain and use the terms: rate of reaction, order, rate constant, half-life, rate-determining step.	361, 362, 374
(b)	Deduce, from a concentration-time graph, the rate of a reaction and the half-life of a first order reaction.	359–360
(c)	Recall that the half-life of a first-order reaction is independent of the concentration.	363
(d)	Deduce, from a rate-concentration graph, the order (0, 1 or 2) with respect to a reactant.	361
(e)	Calculate, using the initial rates method, the order (0, 1 or 2) with respect to a reactant.	362
(f)	Construct a rate equation of the form: $\text{rate} = k[\text{A}]^m[\text{B}]^n$, for which m and n are 0, 1 or 2.	362
(g)	Calculate a rate constant from a rate equation.	362
(h)	Explain qualitatively, the effect of temperature change on a rate constant and hence the rate of a reaction.	365–369

(i)	For a multistep reaction, (i) given the rate-determining step, predict an expression for the rate equation; (ii) show that a rate equation enables a rate-determining step to be proposed; (iii) use a rate equation and the balanced equation for a reaction to suggest possible steps in a reaction mechanism.	373–375
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5.11.2 How Far?

(a)	Understand and use the terms concentration, mole fraction and partial pressure.	7, 140
(b)	Calculate a concentration or partial pressure present at equilibrium, given appropriate data.	300
(c)	Deduce, for homogeneous reactions, expressions for the equilibrium constants: K_c in terms of concentrations, and K_p in terms of partial pressures.	298, 301
(d)	Calculate the values of the equilibrium constants K_c or K_p , including determination of units, given appropriate data.	299–301
(e)	Recall that, for an equilibrium system, (i) changes in concentration and pressure have no effect on the magnitude of the equilibrium constant; (ii) an increase in temperature decreases the value of K_c or K_p for an exothermic reaction and increases the value of K_c or K_p for an endothermic reaction.	299, 305–307 308–309
(f)	Understand that a large value of K_c or K_p indicates a high theoretical yield of products and vice versa.	299
(g)	Appreciate that most organic reactions are in equilibrium.	

5.11.3 Acids, Bases and Buffers

(a)	Describe and use the Brønsted–Lowry theory of acids and bases, to include conjugate acid-base pairs (see also 5.3.3(e)–(i)).	198–200
(b)	Define the terms pH, K_w , K_a and pK_a .	323, 326, 329, 330
(c)	Calculate pH from $[H^+(aq)]$ and $[H^+(aq)]$ from pH (i) for strong monobasic acids and bases; (ii) for weak monobasic acids. For a weak acid HA, it can be assumed that ● $H^+(aq) = A^-(aq)$; ● the equilibrium acid concentration is approximately equal to the undissociated acid concentration.	325, 327, 330 330
(d)	Using acid-base titration pH curves for strong and weak acids and bases, (i) recognise their shapes; (ii) deduce suitable indicators, from supplied pH ranges; (iii) explain why phenolphthalein is unsuitable for titrations involving weak bases and why methyl orange is unsuitable for titrations involving weak acids.	333–334 } 332–334
(e)	Explain the choice of suitable indicators for acid-base titrations, given the pH range of the indicator.	331–332
(f)	Explain what is meant by a <i>buffer solution</i> (as a system that minimises pH changes on addition of an acid or a base).	334
(g)	Explain the role of each component in a buffer solution (for example: CH_3COOH/CH_3COO^- and NH_4^+/NH_3) in the control of pH.	334–335
(h)	Calculate the pH of a buffer solution, for example: from the K_a value of a weak acid and the equilibrium concentrations of the conjugate acid-base pair.	335–336
(i)	State the importance of buffer solutions for controlling pH in blood and shampoos.	337

