

Edexcel (8080) AS and A (9080) GCE Chemistry

Units 1, 2 and 3 constitute the 'AS' course. Units 1, 2, 3, 4, 5 and 6 constitute the 'A' course.

Syllabus
reference

Syllabus content

Page reference in *Chemistry in Context 5th Edn*

Unit 1: Structure, Bonding and Main Group Chemistry

Topic 1.1: Atomic Structure

a	Recall the simple model of an atom in terms of electrons, protons and neutrons.	60
b	Recall the definition of relative isotopic, relative atomic, and relative molecular masses and understand that they are measured on a scale in which $^{12}\text{C} = 12$ exactly.	4
c	Understand the relationships between the numbers of protons and neutrons in an atom and its atomic number, mass number and relative isotopic mass.	61
d	Recall the existence of isotopes and their relevance to the relative atomic mass of an atom.	60–62
e	Understand the principles and use of the low-resolution mass spectrometer to determine relative atomic mass; interpret mass spectrometric data to make simple deductions concerning the isotopic composition of a sample of an element and to calculate its relative atomic mass.	2–4
f	Recall the definitions of first and successive ionisation energies.	64, 67
g	Understand that successive ionisation energies provide evidence for the existence of quantum shells.	67–69
h	Understand that first ionisation energies of successive elements provide evidence for the existence of characteristic energy levels of <i>s</i> , <i>p</i> and <i>d</i> orbitals.	70–72
i	Understand the terms <i>s</i> , <i>p</i> and <i>d</i> -block elements.	39–41
j	Predict the electronic configuration of isolated atoms of the elements from hydrogen to krypton inclusive (using <i>1s</i> , <i>2p</i> ... notation and electrons-in-boxes notations) using the building-up principle.	71
k	Understand that electronic structure determines the chemical properties of an element.	73–74
l	Define first and second electron affinities and understand that the second electron affinity is endothermic.	85–86

Topic 1.2: Formulae, Equations and Moles

a	Determine the empirical formula of a compound from reacting mass data and hence the molecular formula.	6
b	Write full and ionic equations.	17–18
c	Use chemical equations to deduce reacting masses and vice versa.	15–17
d	Define the Avogadro constant and understand the concept of amount of substance.	5
e	Use solution concentration data expressed in g dm^{-3} and mol dm^{-3} to include volumetric analysis.	7, 11–12
f	Understand the concept of the molar volume of a gas and apply this to simple calculations.	13–14
g	Deduce reacting gas volumes from chemical equations.	14

Topic 1.3: Structure and Bonding

a	Understand the nature of ionic, covalent and dative covalent bonds, and the simple charge cloud representation of σ and π bonds.	84–85, 86–90
b	Understand the intermediate nature of most bonds in terms of: (i) electronegativity difference leading to polarity in bonds; (ii) polarising power of cations and polarisability of anions and the factors affecting these.	175–176 96–98 98–99
c	Understand that polar bonds may or may not give rise to permanent dipoles within molecules.	100
d	Understand the nature of intermolecular forces, resulting from interactions between permanent dipoles and induced dipoles, and from the formation of hydrogen bonds.	100 102–112
e	Show how these various types of bond give rise to giant atomic structures (e.g. diamond and graphite), hydrogen-bonded molecular structures (e.g. ice), ionic structures (e.g. sodium chloride), simple molecular structures (e.g. iodine) and polymers (e.g. polyethene), and how the properties of solids are related to the structure and bonding.	123 106 125 130, 429 45–47, 131–132
f	Understand the existence of interparticle forces in the liquid state and hence explain the trends in the boiling temperatures of the noble gases, and in the boiling temperatures of the hydrides of the elements of Groups 4, 5, 6 and 7.	137 100 103–104
g	Interpret changes of state in terms of the types, motion and arrangement of particles present (atoms, molecules or ions) and explain associated energy changes.	123, 106, 125, 130, 45–47, 131–132
h	Recall the shapes of the following molecules and ions: BeCl_2 , BCl_3 , CH_4 , HCl , NH_3 , NH_4^+ , H_2O , CO_2 , SO_2 , CO_3^{2-} , SO_4^{2-} , NO_3^- , PCl_5 , SF_6 .	90–92 (except sulphate, carbonate, nitrate, PCl_5)
i	Interpret these shapes in (h) in terms of the Valence Shell Electron Pair Repulsion Theory, and predict the shapes of related molecules and ions (e.g. from a knowledge of NH_3 the shape of PH_3 can be predicted by analogy).	90–92
j	Describe metallic bonding and explain the electrical conductivity of metals and of graphite in terms of the mobility of electrons.	119–123 124 45–46

Topic 1.4: The Periodic Table 1

a	Demonstrate an understanding that elements in the Periodic Table are arranged in order of ascending atomic number (proton number) and of the electronic configuration from hydrogen to krypton inclusive.	73–74
b	Recall and explain, in terms of structure and bonding, trends in the following properties of the elements across the period sodium to argon of the Periodic Table: (i) melting temperature and boiling temperature; (ii) electrical conductivity; (iii) ionisation energy.	42–45 169–177

Topic 1.5: Introduction to Oxidation and Reduction

a	Define and use the concept of oxidation number.	29
b	Recall that there is a variation in the strength of oxidising and reducing agents.	187, 193
c	Understand oxidation, reduction, oxidising and reducing agents in terms of electron transfer.	20–23
d	Construct ionic half-equations and use them to deduce equations for simple redox processes.	23–24
e	Deduce the oxidation number of an element from the formula of an inorganic species.	30
f	Interpret changes in oxidation number in terms of oxidation and reduction.	31

Topic 1.6: Group 1 (lithium to caesium) and Group 2 (beryllium to barium)

a	Recall the characteristic physical properties of the elements.	212
b	Recall the characteristic flame colours of the compounds and the use of these in analysis (restricted to lithium, sodium, potassium, calcium, strontium and barium compounds) and recognise that the flame colours arise from electronic transitions within the atom.	221 64
c	Understand the trends in ionisation energies of the elements in both groups.	212, 174
d	Recall the reactions of the elements of Group 1 with oxygen and water.	213
e	Recall the reactions of the elements of Group 2 with oxygen, chlorine and water.	213
f	Recall the reactions of the oxides of Groups 1 and 2 with water and dilute acid.	213
g	Recall that the oxidation numbers of elements in Groups 1 and 2 in their compounds are +1 and +2 respectively.	211
h	Recall the trends in solubility of the sulphates and hydroxides of Group 2 elements.	215
i	Recall the trends in thermal stability of the nitrates and carbonates of the elements in Groups 1 and 2; relate these to the sizes and charges of the cations involved.	213–214

Topic 1.7: Group 7 (chlorine to iodine)

a	Recall the physical properties of the elements.	227–228
b	Recall the simple chemical tests for the elements.	
c	Recall and explain the fact that hydrogen halides are water soluble and acidic in solution.	235
d	Recall the use of silver ions and aqueous ammonia in the identification of chloride, bromide and iodide ions.	233–234
e	Understand the reactions of halide salts with concentrated sulphuric acid in relation to the reducing power of the hydrogen halides.	234
f	Recall the existence of the +1 and +5 oxidation states of chlorine in the oxo-anions chlorate(I) and chlorate(V).	231–232
g	Recall the disproportionation of chlorine and of the chlorate(I) ion.	232
h	Recall the relative strengths of the elements as oxidising agents and understand the application of this to the extraction of bromine from sea water.	230–231

Unit 2: Introductory Organic Chemistry, Energetics, Kinetics, Equilibrium and Applications**Topic 2.1: Energetics 1**

a	Understand the concept of enthalpy change, ΔH , and the necessity to express this under standard conditions.	148
b	Recall that ΔH is negative for exothermic reactions, and positive for endothermic reactions.	146–147
c	Recall the definitions of standard enthalpy of formation, ΔH_f^\ominus , standard enthalpy of combustion ΔH_c^\ominus and standard enthalpy of neutralisation, ΔH_{neu}^\ominus .	148–150
d	Construct simple enthalpy level diagrams.	149, 151, 152
e	Understand that the sign of the enthalpy change does not always indicate the direction of spontaneous change.	154–155
f	Recall and apply Hess's Law to calculate enthalpy changes.	151–153
g	Describe experiments to measure values of ΔH for simple laboratory reactions and perform calculations based on the results of these experiments.	148, 150
h	Use average bond enthalpy data to calculate approximate enthalpy changes.	155–160

Topic 2.2: Organic Chemistry 1 (Introduction, Alkanes, Alkenes, Halogenoalkanes and Alcohols)

a	Introduction (i) Understand the concept of homologous series. (ii) Apply the rules for IUPAC nomenclature for compounds relevant to this specification. (iii) Recognise and predict the existence of structural isomerism within the types of compounds found in this topic. (iv) Recognise the existence of geometric (cis-trans) isomerism resulting from restricted rotation about a carbon-carbon double bond. (v) Be able to classify reactions as free radical, electrophilic addition, nucleophilic substitution, elimination, hydrolysis, reduction, oxidation or polymerisation.	380–381 400–402, 418, 453, 467 390–392 419–420 403–406
b	Reactions of organic compounds Recall, in terms of reagents and general reaction conditions, the reactions of: (i) alkanes with oxygen (air), chlorine and bromine; (ii) alkenes (to include unsymmetrical alkenes) with hydrogen, halogens, hydrogen halides and potassium manganate (VII). The use of bromine solution as a test for alkenes; (iii) halogeno-compounds with potassium hydroxide (to produce alcohols or alkenes) potassium cyanide and ammonia; identification of the halide group by alkaline hydrolysis, subsequent acidification and testing with aqueous silver nitrate; (iv) primary, secondary and tertiary alcohols with acidified potassium dichromate(VI), dehydrating agents (e.g. concentrated sulphuric acid) and halogenating agents (phosphorus pentachloride, sodium bromide and concentrated sulphuric acid, phosphorus and iodine); the use of phosphorus pentachloride as a test for the —OH group.	407–409 422–425, 422 455–456, 460–462, 233, 455 472–476
c	Bonding and reactivity Use the concepts of single and double covalent bonds, bond polarity and bond enthalpy as factors influencing relative reactivity in appropriate reactions of alkanes, alkenes, halogenoalkanes and alcohols.	407, 419, 421, 455
d	Quantitative chemistry related to this topic Calculate: (i) empirical and molecular formulae; (ii) theoretical yield and percentage yield in reactions from reacting mass data.	138, 381 532–533
e	Applied organic chemistry Discuss (i) the advantages and disadvantages of liquid and gaseous fuels; (ii) the structures and properties of the polymers of simple alkenes and their uses; (iii) recall that halogens are used in the manufacture of herbicides and polymers; understand that the strength of the C—X bond gives a useful life to many of these compounds.	426–432 235–236

Topic 2.3: Kinetics 1

a	Recall the factors which control the rate of a chemical reaction.	352–356
b	Give an explanation of the changes in rate based on a qualitative understanding of the collision theory.	
c	Relate, in a qualitative way, the changes in the Maxwell-Boltzmann distribution of molecular energies with changes of temperature to the alteration in rate of reaction.	366–369
d	Understand the concept of activation energy and its qualitative relationship to the effect of temperature changes on the rate of reaction.	366–369
e	Recall the role of catalysts in providing alternative reaction routes of lower activation energy.	369–372
f	Understand the concepts of thermodynamic and kinetic stability.	153–154

Topic 2.4: Chemical Equilibria I

a	Understand that equilibria are dynamic.	290–292
b	Deduce the qualitative effects of changes of temperature, pressure and concentration on the position of a homogeneous equilibrium.	305–309
c	Understand the importance of temperature control to the establishment of the optimum reaction conditions for an industrial process.	310

Topic 2.5: Industrial Inorganic Chemistry

a	Recall the conditions used in the manufacture of ammonia by the Haber Process and its conversion to nitric acid.	312–315
b	Recall conditions used in the manufacture of sulphuric acid by the Contact Process.	310–311
c	Apply the knowledge gained in this unit concerning enthalpy change, equilibria and kinetics to justify the conditions used to obtain an economic yield in the processes in (a) and (b).	310–315
d	Recall the use of ammonia, nitric acid and sulphuric acid in the manufacture of inorganic fertilisers.	312, 315
e	Recall and explain the electrolytic extraction of aluminium from purified bauxite and the fact that it is economically important yet expensive.	274–276
f	Recall the production of chlorine and sodium chlorate(I) by the electrolysis of sodium chloride solution.	226, 235
g	Recall the uses of chlorine and sodium chlorate(I).	235

Unit 4: Periodicity, Quantitative Equilibria and Functional Group Chemistry**Topic 4.1: Energetics II**

a	Understand the definition of enthalpy of atomisation, ΔH_a^θ , enthalpy of hydration, $\Delta H_{\text{hyd}}^\theta$ and lattice enthalpy $\Delta H_{\text{latt}}^\theta$.	149, 160, 163
b	Construct a Born-Haber cycle and carry out associated calculations.	160–162
c	Understand that values of lattice enthalpies calculated from the Born-Haber cycle may differ from those calculated from a purely ionic model	162
d	Understand the factors that influence the value of lattice enthalpy	161
e	Understand the part played by lattice enthalpy and enthalpy of hydration in rationalising the variation in solubilities of the hydroxides and the sulphates of Group 2.	162–163, 215

Topic 4.2: The Periodic Table II (Period 3 and Group 4)

a	<p>Variation of properties across a period</p> <p>Show the variation in properties across Period 3 (sodium to argon) as illustrated by:</p> <p>(i) reactions of the elements with oxygen, chlorine and water;</p> <p>(ii) the formulae and acid-base character of the oxides and hydroxides of the metals and oxides of the non-metals;</p> <p>(iii) the formulae of the chlorides, and their reactions with water;</p> <p>(iv) interpret the reactions in (a)(ii) and (a)(iii) in terms of the structure and bonding of the oxides and chlorides.</p>	<p>47</p> <p>177, 180–184</p> <p>177–180</p> <p>48, 177–180, 180–184</p>
b	<p>Variation of properties down a group</p> <p>(i) recall the reasons for the increase in metallic character with increase in atomic number in Group 4;</p> <p>(ii) recall that the +2 oxidation state in Group 4 becomes more stable as the atomic number increases;</p> <p>(iii) recall and explain the structure of carbon tetrachloride;</p> <p>(iv) recall and explain the behaviour of carbon tetrachloride with water and contrast this behaviour with that of silicon tetrachloride with water.</p>	<p>241–243</p> <p>243–246</p> <p>244, 245</p> <p>180 (not an explanation)</p>

Topic 4.3: Chemical Equilibria II

a	Define the terms <i>partial pressure</i> and <i>concentration</i> .	140, 7
b	Deduce expressions for the equilibrium constants K_c and K_p for homogenous reactions from given equations and calculate their numerical values with units, given suitable data.	296–301
c	Recall that expressions for K_p and K_c for heterogeneous equilibria do not include values for solid and liquid phases.	302–303
d	Recall that the presence of a catalyst does not affect the position of equilibrium.	308
e	Determine the equilibrium partial pressures in a simple binary gaseous dissociation (eg that of dinitrogen tetroxide) given the value of K_p .	
f	Understand that changes in temperature result in a change in the value of K_c and K_p and that the position of equilibrium will change with change in temperature.	308–309

Topic 4.4: Acid-base Equilibria

a	Recall the Brønsted-Lowry theory and use it to identify acid-base behaviour, and identify acid-base conjugate pairs and relate them by means of suitable equations.	198–200
b	Define pH.	326–327
c	Understand the terms 'strong' and 'weak' as applied to acids and bases.	325
d	Define K_a and K_w and recall their units.	326, 329
e	Define pK_a and pK_w .	330
f	Calculate the pH of solutions of strong acids and strong bases.	327
g	Calculate pH of weak acids given K_a and vice versa.	330
h	Understand the principles involved in acid-base titrations.	333–334
i	Recall the sketch curves for the variation in pH during the following titrations: strong acid-strong base, weak acid- strong base and strong acid- weak base.	333–334
j	Use titration curves to determine K_a for a weak acid.	
k	Explain the choice of a suitable indicator for an acid-base titration given pK_{Ind} values.	331–332
l	Explain the action of a buffer solution and calculate its pH from suitable data.	334–337
m	Demonstrate an understanding of how enthalpy of neutralisation is related to the strength of acids and bases.	

Topic 4.5: Organic Chemistry II (Acids, Esters, Carbonyl Compounds, Acid Chlorides, Nitrogen Compounds and Further Halogeno-compounds)

a	Further ideas in organic chemistry	
	(i) apply the concept of functional groups to the range of organic compounds found in Units 2 and 4;	380–381
	(ii) recognise and predict the existence of structural isomers within the types of organic compound found in Units 2 and 4;	391–392
	(iii) recognise stereoisomerism as geometrical (cis-trans) or optical;	392, 418
	(iv) explain the existence of geometrical (cis-trans) isomerism resulting from restricted rotation about a carbon-carbon double bond;	419–420
	(v) understand the existence of optical isomerism resulting from a chiral centre in molecules with a single asymmetric carbon atom, and understand optical isomers as object and non-superimposable mirror images;	392–393
	(vi) recall optical activity as the ability of a single optical isomer to rotate the plane of polarisation of plane polarised monochromatic light and understand the nature of a racemic mixture.	392, 393

b	Further reactions of organic compounds	
	Recall, in terms of reagents and general reaction conditions, the following reactions and classify reactions in this topic as oxidation, reduction, nucleophilic substitution or nucleophilic addition:	
	(i) halogeno-compounds with magnesium to form Grignard reagents and the reactions of the latter with water, carbon dioxide and carbonyl compounds;	505–506, 500, 504, 505
	(ii) carboxylic acids with alcohols, lithium tetrahydridoaluminate(III) (lithium aluminium hydride), phosphorus pentachloride, sodium carbonate and sodium hydrogencarbonate;	
	(iii) esters with acids and alkalis;	
	(iv) carbonyl compounds with hydrogen cyanide, 2,4-dinitrophenylhydrazine, alkaline ammoniacal silver nitrate solution, Fehling's solution, iodine in the presence of alkali (or potassium iodide and sodium chlorate(I)), sodium tetrahydridoborate(III) (sodium borohydride) and lithium tetrahydridoaluminate(III) (lithium aluminium hydride);	484–489
	(v) ethanoyl chloride with water, alcohols, ammonia and primary amines;	478–479
	(vi) primary amines with aqueous hydrogen ions, acid chlorides;	
	(vii) nitriles undergoing hydrolysis and undergoing reduction;	463, 511
(viii) amides with phosphorus(V) oxide and bromine in aqueous alkali;	519–521	
(ix) amino acids with acids and bases, and their zwitterion structures.	531, 499, 519 531, 532 524	

Unit 5: Transition Metals, Quantitative Kinetics and Applied Organic Chemistry

Topic 5.1: Redox Equilibria (application)

a	Relate changes in oxidation number to reaction stoichiometry.	23–24
b	Understand the procedures and principles involved in the use of potassium manganate (VII) to estimate reducing agents and potassium iodide and sodium thiosulphate to estimate oxidising agents.	207–208
c	Recall the definition of standard electrode potential (standard reduction potential) and understand the need for a standard electrode.	187–192
d	Predict the likely direction of spontaneous change of redox reactions, using standard electrode potential data, and understand why these predictions may not be borne out in practice.	193–195
e	Understand disproportionation reactions in terms of standard electrode potentials.	
f	Understand the applications of electrode potentials in connection with corrosion and to the solution of problems caused by corrosion.	271–272
g	Understand the application of electrode potentials to the construction of simple storage cells.	196–198

Topic 5.2: Transition Metal Chemistry

a	Derive the electronic configurations of atoms of the <i>d</i> -block elements (Sc to Zn), and their simple ions, from their position in the Periodic Table.	249–250
b	Recall the transition metals as <i>d</i> -block elements forming one or more stable ions which have incompletely filled <i>d</i> -orbitals.	250–251
c	Recall the characteristic properties of the transition elements, such as: (i) the formation of coloured aqueous ions, and other complex ions; (ii) the formation of a range of compounds in which they are present in different and stable oxidation states.	254–261
d	Understand the nature of the bonding in complex ions, including the aqua-ions, their shape and the cause of their colour.	257, 259–260
e	Understand simple ligand exchange processes.	203
f	Recall the formation of hydroxide precipitates on the addition of aqueous solutions of sodium hydroxide or ammonia, and that some hydroxide precipitates react with an excess of strong alkali, and some react with an excess of ammonia; limited to Cr^{3+} , Mn^{2+} , Fe^{2+} , Fe^{3+} , Co^{2+} , Ni^{2+} , Cu^{2+} , Zn^{2+} .	284–285

g	Recall the oxidation states of vanadium (+2, +3, +4, +5) in its compounds, and deduce given E^θ values, reagents for the interconversion of the metal ions, oxo anions and oxo cations of the element in these oxidation states.	255
h	Describe reactions for the interconversion of the oxidation states of vanadium in aqueous solution.	255
i	Recall that transition elements and their compounds are important catalysts in industrial processes, and that their catalytic activity is often associated with the variable oxidation states of the elements.	261–262
j	Recall examples of catalytic action by vanadium, iron, nickel and/or their compounds.	261–262

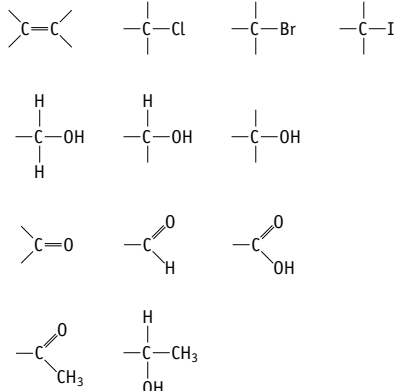
Topic 5.3: Organic Chemistry III (Reaction Mechanisms and Aromatic Compounds)

a	<p>Structure of benzene and reactions of aromatic compounds</p> <p>Use the concepts of the different types of covalent bonds, and bond enthalpy to explain the structure and stability of the benzene ring.</p> <p>Recall, in terms of reagents and reaction conditions, the reactions of:</p> <p>(i) benzene with a nitrating mixture, bromine, chloroalkanes or acid chlorides in the presence of anhydrous aluminium chloride;</p> <p>(ii) aromatic compounds with carbon-containing side chains with alkaline potassium manganate(VII) solution, resulting in the oxidation of the side-chains;</p> <p>(iii) phenol with sodium hydroxide, bromine and ethanoyl chloride;</p> <p>(iv) reduction of aromatic nitro-compounds to amines using tin and concentrated hydrochloric acid;</p> <p>(v) phenylamine with nitrous acid; and the subsequent coupling reaction of benzenediazonium ions with phenol.</p>	<p>436–438</p> <p>442–445</p> <p>447</p> <p>471, 477, 463</p> <p>518–519</p> <p>521, 522–523</p>
b	<p>Reaction mechanisms</p> <p>Recall the following reaction mechanisms together with reagents and general conditions for the reactions shown and apply them to simple allied reactions:</p> <p>(i) homolytic, free radical substitution (alkanes with chlorine);</p> <p>(ii) homolytic, free radical addition (polymerisation of ethene);</p> <p>(iii) heterolytic, electrophilic addition (symmetrical and unsymmetrical alkenes with halogens and hydrogen halides);</p> <p>(iv) heterolytic, electrophilic substitution (benzene with a nitrating mixture, with bromine and with chloroalkane and acid chlorides);</p> <p>(v) heterolytic, nucleophilic substitution (halogenoalkanes with hydroxide ions and cyanide ions) S_N1 and S_N2;</p> <p>(vi) heterolytic, nucleophilic addition (carbonyl compounds with hydrogen cyanide).</p>	<p>407–408</p> <p>429</p> <p>421, 423–424</p> <p>441</p> <p>455, 456, 457</p> <p>484–485</p>

Topic 5.4: Chemical Kinetics II

a	Recall that rates of reaction may be expressed by empirical rate equations of the form: $\text{rate} = k[A]^m[B]^n$, where m and n are 0, 1 or 2.	358–362
b	Define the terms <i>rate constant</i> and <i>order of reaction</i> and understand that these are experimentally determined.	360–362
c	Deduce rate equations from given experimental initial rate data.	361–362
d	Recall that reactions with a large activation energy will have a small rate constant.	366–367
e	Understand that many reactions take place in several steps, one of which will be the rate-determining step.	374
f	Understand that it is sometimes possible to deduce information regarding the mechanism of a chemical reaction from kinetic data.	373–375
g	Understand that many reactions proceed through a transition state.	369–370
h	Select and describe a suitable experimental technique for following a given reaction.	356–358
i	Present and interpret the results of kinetic measurements in graphical form.	358–361
j	Define the term <i>half-life</i> and recall that this is constant for any given first order reaction.	363–365

Topic 5.5: Organic Chemistry IV (Analysis, Synthesis and Application)

<p>a</p>	<p>Organic analysis</p> <p>(i) describe practical tests or combinations of tests to confirm the presence of the following functional groups:</p> <div style="text-align: center;">  </div> <p>(ii) interpret physical data and chemical information, including information relating to derivatives where appropriate, to arrive at the structural formula of a compound;</p> <p>(iii) (a) interpret simple fragmentation patterns from a mass spectrometer;</p> <p>(b) interpret simple infra-red spectra;</p> <p>(c) interpret simple low-resolution nuclear magnetic resonance spectra;</p> <p>(d) interpret simple ultra-violet/visible spectra.</p>	<p>382</p> <p>383</p> <p>385–387</p> <p>387–389</p> <p>384, 259</p>
<p>b</p>	<p>Organic synthesis</p> <p>(i) propose practical pathways for the synthesis of organic molecules;</p> <p>(ii) propose suitable apparatus, conditions and safety precautions for carrying out organic syntheses, given suitable information;</p> <p>(iii) demonstrate familiarity with a range of practical techniques used in organic chemistry;</p> <p>(iv) demonstrate an understanding of the principles of fractional distillation in terms of the graphs of boiling point against composition.</p>	<p>531, 533</p>
<p>c</p>	<p>Applied organic chemistry</p> <p>Appreciate the importance of organic compounds in pharmaceuticals, agricultural products and materials, confined to the following aspects:</p> <p>(i) changes to the relative lipid/water solubility of pharmaceuticals by the introduction of non-polar side-chains or ionic groups;</p> <p>(ii) the use of organic compounds such as urea as sources of nitrogen in agriculture and their advantages as compared with inorganic compounds containing nitrogen;</p> <p>(iii) the use of esters, oils and fats;</p> <p>(iv) properties and uses of addition polymers of ethene, propene, chloroethene, tetrafluoroethene and phenylethene, and of the condensation polymers (polyesters and polyamides).</p>	<p>501–504</p> <p>509–510</p> <p>429–431, 507, 525</p>