Cambridge Assessment
International Education

Syllabus

Cambridge International AS & A Level
Chemistry 9701

Use this syllabus for exams in 2022, 2023 and 2024.
Exams are available in the June and November series.
Exams are also available in the March series in India only.

Cambridge Pathway

Weighting for assessment objectives

The approximate weightings allocated to each of the assessment objectives (AOs) are summarised below.

Assessment objectives as a percentage of each qualification

Assessment objective	Weighting in AS Level %	Weighting in A Level %
AO1 Knowledge and understanding	40	40
AO2 Handling, applying and evaluating information	40	40
AO3 Experimental skills and investigations	20	20
Total	100	100

Assessment objectives as a percentage of each component

Assessment objective	Weighting in components %				
	Paper 1	Paper 2	Paper 3	Paper 4	Paper 5
AO1 Knowledge and understanding	50	50	0	50	0
AO2 Handling, applying and evaluating information	50	50	0	50	0
AO3 Experimental skills and investigations	0	0	100	0	100
Total	100	100	100	100	100

3 Subject content

Candidates for Cambridge International AS Level should study topics 1–21.

Candidates for Cambridge International A Level should study all topics.

The content of the AS Level learning outcomes is assumed knowledge for the A Level components.

Teachers should refer to the social, environmental, economic and technological aspects of chemistry wherever possible throughout the syllabus. Some examples are included in the syllabus and teachers should encourage learners to apply the principles of these examples to other situations introduced during the course.

The syllabus content for practical skills is in the Practical assessment section.

Teachers should ensure that candidates are prepared for the assessment of both theory learning outcomes and practical expectations.

This syllabus gives you the flexibility to design a course that will interest, challenge and engage your learners. Where appropriate you are responsible for selecting suitable subject contexts, resources and examples to support your learners' study. These should be appropriate for the learners' age, cultural background and learning context as well as complying with your school policies and local legal requirements.

Nomenclature

Symbols, signs and abbreviations used in examination papers will follow the recommendations made in the ASE publication *Signs, Symbols and Systematics* (The ASE Companion to 16–19 Science, 2000) although the traditional names sulfite, nitrite, sulfur trioxide, sulfurous acid and nitrous acid will be used. Sulfur and all compounds of sulfur will be spelled with f, not ph.

Decimal markers

In accordance with current ASE convention, decimal markers in examination papers will be a single dot on the line. Candidates are expected to follow this convention in their answers.

Units

In practical work, candidates will be expected to use SI units or, where appropriate, units approved by the BIPM for use with the SI (e.g. minute). A list of SI units and units approved for use with the SI may be found in the SI brochure at **www.bipm.org**

The use of imperial/customary units such as the inch and degree Fahrenheit is not acceptable and should be discouraged. In all examinations, where data is supplied for use in questions, candidates will be expected to use units that are consistent with the units supplied, and should not attempt conversion to other systems of units unless this is a requirement of the question.

Special note regarding units and significant figures

Candidates should be aware that misuse of units and/or significant figures, i.e. failure to quote units where necessary, the inclusion of units in quantities defined as ratios or quoting answers to an inappropriate number of significant figures, is liable to be penalised in the assessments.

AS Level subject content

Physical chemistry

1   Atomic structure

1.1   Particles in the atom and atomic radius

Learning outcomes

Candidates should be able to:

1   understand that atoms are mostly empty space surrounding a very small, dense nucleus that contains protons and neutrons; electrons are found in shells in the empty space around the nucleus

2   identify and describe protons, neutrons and electrons in terms of their relative charges and relative masses

3   understand the terms atomic and proton number; mass and nucleon number

4   describe the distribution of mass and charge within an atom

5   describe the behaviour of beams of protons, neutrons and electrons moving at the same velocity in an electric field

6   determine the numbers of protons, neutrons and electrons present in both atoms and ions given atomic or proton number, mass or nucleon number and charge

7   state and explain qualitatively the variations in atomic radius and ionic radius across a period and down a group

1.2   Isotopes

Learning outcomes

Candidates should be able to:

1   define the term isotope in terms of numbers of protons and neutrons

2   understand the notation $^{x}_{y}A$ for isotopes, where x is the mass or nucleon number and y is the atomic or proton number

3   state that and explain why isotopes of the same element have the same chemical properties

4   state that and explain why isotopes of the same element have different physical properties, limited to mass and density
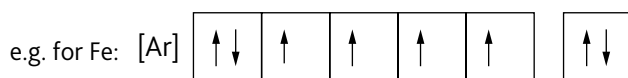
1.3   Electrons, energy levels and atomic orbitals

In 1.3 each atom or ion described will be in the ground state. Only the elements hydrogen to krypton will be assessed.

Learning outcomes

Candidates should be able to:

1   understand the terms:

   - shells, sub-shells and orbitals

   - principal quantum number (n)

   - ground state, limited to electronic configuration

2   describe 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

3   describe the order of increasing energy of the sub-shells within the first three shells and the 4s and 4p sub-shells

4   describe the electronic configurations to include the number of electrons in each shell, sub-shell and orbital

5   explain the electronic configurations in terms of energy of the electrons and inter-electron repulsion

6   determine the electronic configuration of atoms and ions given the atomic or proton number and charge, using either of the following conventions:

   e.g. for Fe: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$ (full electron configuration)

   or $[Ar]\, 3d^6\, 4s^2$ (shorthand electron configuration)

7   understand and use the electrons in boxes notation

   e.g. for Fe: [Ar] | ↑↓ | ↑ | ↑ | ↑ | ↑ | | ↑↓ |

8   describe and sketch the shapes of s and p orbitals

9   describe a free radical as a species with one or more unpaired electrons

1.4   Ionisation energy

In 1.4 each atom or ion described will be in the ground state. Only the elements hydrogen to krypton will be assessed.

Learning outcomes

Candidates should be able to:

1   define and use the term first ionisation energy, IE

2   construct equations to represent first, second and subsequent ionisation energies

3   identify and explain the trends in ionisation energies across a period and down a group of the Periodic Table

4   identify and explain the variation in successive ionisation energies of an element

5   understand that ionisation energies are due to the attraction between the nucleus and the outer electron

6   explain the factors influencing the ionisation energies of elements in terms of nuclear charge, atomic/ionic radius, shielding by inner shells and sub-shells and spin-pair repulsion

7   deduce the electronic configurations of elements using successive ionisation energy data

8   deduce the position of an element in the Periodic Table using successive ionisation energy data

2   Atoms, molecules and stoichiometry

2.1   Relative masses of atoms and molecules

Learning outcomes

Candidates should be able to:

1   define the unified atomic mass unit as one twelfth of the mass of a carbon-12 atom

2   define relative atomic mass, A_r, relative isotopic mass, relative molecular mass, M_r, and relative formula mass in terms of the unified atomic mass unit

2.2   The mole and the Avogadro constant

Learning outcomes

Candidates should be able to:

1   define and use the term mole in terms of the Avogadro constant

2.3   Formulae

Learning outcomes

Candidates should be able to:

1   write formulae of ionic compounds from ionic charges and oxidation numbers (shown by a Roman numeral), including:

  (a) the prediction of ionic charge from the position of an element in the Periodic Table

  (b) recall of the names and formulae for the following ions: NO_3^-, CO_3^{2-}, SO_4^{2-}, OH^-, NH_4^+, Zn^{2+}, Ag^+, HCO_3^-, PO_4^{3-}

2   (a) write and construct equations (which should be balanced), including ionic equations (which should not include spectator ions)

  (b) use appropriate state symbols in equations

3   define and use the terms empirical and molecular formula

4   understand and use the terms anhydrous, hydrated and water of crystallisation

5   calculate empirical and molecular formulae, using given data

2.4   Reacting masses and volumes (of solutions and gases)

Learning outcomes

Candidates should be able to:

1   perform calculations including use of the mole concept, involving:

  (a) reacting masses (from formulae and equations) including percentage yield calculations

  (b) volumes of gases (e.g. in the burning of hydrocarbons)

  (c) volumes and concentrations of solutions

  (d) limiting reagent and excess reagent

  (When performing calculations, candidates' answers should reflect the number of significant figures given or asked for in the question. When rounding up or down, candidates should ensure that significant figures are neither lost unnecessarily nor used beyond what is justified (see also Mathematical requirements section).)

  (e) deduce stoichiometric relationships from calculations such as those in 2.4.1 (a)–(d)

3   Chemical bonding

3.1   Electronegativity and bonding

Learning outcomes

Candidates should be able to:

1   define electronegativity as the power of an atom to attract electrons to itself

2   explain the factors influencing the electronegativities of the elements in terms of nuclear charge, atomic radius and shielding by inner shells and sub-shells

3   state and explain the trends in electronegativity across a period and down a group of the Periodic Table

4   use the differences in Pauling electronegativity values to predict the formation of ionic and covalent bonds (the presence of covalent character in some ionic compounds will not be assessed) (Pauling electronegativity values will be given where necessary)

3.2   Ionic bonding

Learning outcomes

Candidates should be able to:

1   define ionic bonding as the electrostatic attraction between oppositely charged ions (positively charged cations and negatively charged anions)

2   describe ionic bonding including the examples of sodium chloride, magnesium oxide and calcium fluoride

3.3   Metallic bonding

Learning outcomes

Candidates should be able to:

1   define metallic bonding as the electrostatic attraction between positive metal ions and delocalised electrons

3.4   Covalent bonding and coordinate (dative covalent) bonding

Learning outcomes

Candidates should be able to:

1   define covalent bonding as electrostatic attraction between the nuclei of two atoms and a shared pair of electrons

(a)   describe covalent bonding in molecules including:

- hydrogen, H_2
- oxygen, O_2
- nitrogen, N_2
- chlorine, Cl_2
- hydrogen chloride, HCl
- carbon dioxide, CO_2
- ammonia, NH_3
- methane, CH_4
- ethane, C_2H_6
- ethene, C_2H_4

(continued)

3.4   Covalent bonding and coordinate (dative covalent) bonding (continued)

Learning outcomes

Candidates should be able to:

(b) understand that elements in period 3 can expand their octet including in the compounds sulfur dioxide, SO_2, phosphorus pentachloride, PCl_5, and sulfur hexafluoride, SF_6

(c) describe coordinate (dative covalent) bonding, including in the reaction between ammonia and hydrogen chloride gases to form the ammonium ion, NH_4^+, and in the Al_2Cl_6 molecule

2   (a) describe covalent bonds in terms of orbital overlap giving σ and π bonds:

- σ bonds are formed by direct overlap of orbitals between the bonding atoms
- π bonds are formed by the sideways overlap of adjacent p orbitals above and below the σ bond

(b) describe how the σ and π bonds form in molecules including H_2, C_2H_6, C_2H_4, HCN and N_2

(c) use the concept of hybridisation to describe sp, sp^2 and sp^3 orbitals

3   (a) define the terms:

- bond energy as the energy required to break one mole of a particular covalent bond in the gaseous state
- bond length as the internuclear distance of two covalently bonded atoms

(b) use bond energy values and the concept of bond length to compare the reactivity of covalent molecules

3.5   Shapes of molecules

Learning outcomes

Candidates should be able to:

1   state and explain the shapes of, and bond angles in, molecules by using VSEPR theory, including as simple examples:

- BF_3 (trigonal planar, 120°)
- CO_2 (linear, 180°)
- CH_4 (tetrahedral, 109.5°)
- NH_3 (pyramidal, 107°)
- H_2O (non-linear, 104.5°)
- SF_6 (octahedral, 90°)
- PF_5 (trigonal bipyramidal, 120° and 90°)

2   predict the shapes of, and bond angles in, molecules and ions analogous to those specified in 3.5.1

3.6 Intermolecular forces, electronegativity and bond properties

Learning outcomes

Candidates should be able to:

1. (a) describe hydrogen bonding, limited to molecules containing N–H and O–H groups, including ammonia and water as simple examples

   (b) use the concept of hydrogen bonding to explain the anomalous properties of H_2O (ice and water):
   - its relatively high melting and boiling points
   - its relatively high surface tension
   - the density of the solid ice compared with the liquid water

2. use the concept of electronegativity to explain bond polarity and dipole moments of molecules

3. (a) describe van der Waals' forces as the intermolecular forces between molecular entities other than those due to bond formation, and use the term van der Waals' forces as a generic term to describe all intermolecular forces

   (b) describe the types of van der Waals' force:
   - instantaneous dipole – induced dipole (id-id) force, also called London dispersion forces
   - permanent dipole – permanent dipole (pd-pd) force, including hydrogen bonding

   (c) describe hydrogen bonding and understand that hydrogen bonding is a special case of permanent dipole – permanent dipole force between molecules where hydrogen is bonded to a highly electronegative atom

4. state that, in general, ionic, covalent and metallic bonding are stronger than intermolecular forces

3.7 Dot-and-cross diagrams

Learning outcomes

Candidates should be able to:

1. use dot-and-cross diagrams to illustrate ionic, covalent and coordinate bonding including the representation of any compounds stated in 3.4 and 3.5 (dot-and-cross diagrams may include species with atoms which have an expanded octet or species with an odd number of electrons)

4 States of matter

4.1 The gaseous state: ideal and real gases and $pV = nRT$

Learning outcomes

Candidates should be able to:

1. explain the origin of pressure in a gas in terms of collisions between gas molecules and the wall of the container

2. understand that ideal gases have zero particle volume and no intermolecular forces of attraction

3. state and use the ideal gas equation $pV = nRT$ in calculations, including in the determination of M_r

4.2   Bonding and structure

Learning outcomes

Candidates should be able to:

1   describe, in simple terms, the lattice structure of a crystalline solid which is:

    (a) giant ionic, including sodium chloride and magnesium oxide

    (b) simple molecular, including iodine, buckminsterfullerene C_{60} and ice

    (c) giant molecular, including silicon(IV) oxide, graphite and diamond

    (d) giant metallic, including copper

2   describe, interpret and predict the effect of different types of structure and bonding on the physical properties of substances, including melting point, boiling point, electrical conductivity and solubility

3   deduce the type of structure and bonding present in a substance from given information

5   Chemical energetics

5.1   Enthalpy change, ΔH

Learning outcomes

Candidates should be able to:

1   understand that chemical reactions are accompanied by enthalpy changes and these changes can be exothermic (ΔH is negative) or endothermic (ΔH is positive)

2   construct and interpret a reaction pathway diagram, in terms of the enthalpy change of the reaction and of the activation energy

3   define and use the terms:

    (a) standard conditions (this syllabus assumes that these are 298 K and 101 kPa) shown by $^{\ominus}$.

    (b) enthalpy change with particular reference to: reaction, ΔH_r, formation, ΔH_f, combustion, ΔH_c, neutralisation, ΔH_{neut}

4   understand that energy transfers occur during chemical reactions because of the breaking and making of chemical bonds

5   use bond energies (ΔH positive, i.e. bond breaking) to calculate enthalpy change of reaction, ΔH_r

6   understand that some bond energies are exact and some bond energies are averages

7   calculate enthalpy changes from appropriate experimental results, including the use of the relationships $q = mc\Delta T$ and $\Delta H = -mc\Delta T$

5.2   Hess's Law

Learning outcomes

Candidates should be able to:

1   apply Hess's Law to construct simple energy cycles

2   carry out calculations using cycles and relevant energy terms, including:

    (a) determining enthalpy changes that cannot be found by direct experiment

    (b) use of bond energy data

6   Electrochemistry

6.1   Redox processes: electron transfer and changes in oxidation number (oxidation state)

Learning outcomes

Candidates should be able to:

1   calculate oxidation numbers of elements in compounds and ions

2   use changes in oxidation numbers to help balance chemical equations

3   explain and use the terms redox, oxidation, reduction and disproportionation in terms of electron transfer and changes in oxidation number

4   explain and use the terms oxidising agent and reducing agent

5   use a Roman numeral to indicate the magnitude of the oxidation number of an element

7   Equilibria

7.1   Chemical equilibria: reversible reactions, dynamic equilibrium

Learning outcomes

Candidates should be able to:

1   (a) understand what is meant by a reversible reaction

    (b) understand what is meant by dynamic equilibrium in terms of the rate of forward and reverse reactions being equal and the concentration of reactants and products remaining constant

    (c) understand the need for a closed system in order to establish dynamic equilibrium

2   define Le Chatelier's principle as: if a change is made to a system at dynamic equilibrium, the position of equilibrium moves to minimise this change

3   use Le Chatelier's principle to deduce qualitatively (from appropriate information) the effects of changes in temperature, concentration, pressure or presence of a catalyst on a system at equilibrium

4   deduce expressions for equilibrium constants in terms of concentrations, K_c

5   use the terms mole fraction and partial pressure

6   deduce expressions for equilibrium constants in terms of partial pressures, K_p
(use of the relationship between K_p and K_c is not required)

7   use the K_c and K_p expressions to carry out calculations (such calculations will not require the solving of quadratic equations)

8   calculate the quantities present at equilibrium, given appropriate data

9   state whether changes in temperature, concentration or pressure or the presence of a catalyst affect the value of the equilibrium constant for a reaction

10   describe and explain the conditions used in the Haber process and the Contact process, as examples of the importance of an understanding of dynamic equilibrium in the chemical industry and the application of Le Chatelier's principle

7.2   Brønsted–Lowry theory of acids and bases

Learning outcomes

Candidates should be able to:

1   state the names and formulae of the common acids, limited to hydrochloric acid, HCl, sulfuric acid, H_2SO_4, nitric acid, HNO_3, ethanoic acid, CH_3COOH

2   state the names and formulae of the common alkalis, limited to sodium hydroxide, NaOH, potassium hydroxide, KOH, ammonia, NH_3

(continued)

7.2 Brønsted–Lowry theory of acids and bases (continued)

Learning outcomes

Candidates should be able to:

3 describe the Brønsted–Lowry theory of acids and bases

4 describe strong acids and strong bases as fully dissociated in aqueous solution and weak acids and weak bases as partially dissociated in aqueous solution

5 appreciate that water has pH of 7, acid solutions pH of below 7 and alkaline solutions pH of above 7

6 explain qualitatively the differences in behaviour between strong and weak acids including the reaction with a reactive metal and difference in pH values by use of a pH meter, universal indicator or conductivity

7 understand that neutralisation reactions occur when $H^+(aq)$ and $OH^-(aq)$ form $H_2O(l)$

8 understand that salts are formed in neutralisation reactions

9 sketch the pH titration curves of titrations using combinations of strong and weak acids with strong and weak alkalis

10 select suitable indicators for acid-alkali titrations, given appropriate data (pK_a values will not be used)

8 Reaction kinetics

8.1 Rate of reaction

Learning outcomes

Candidates should be able to:

1 explain and use the term rate of reaction, frequency of collisions, effective collisions and non-effective collisions

2 explain qualitatively, in terms of frequency of effective collisions, the effect of concentration and pressure changes on the rate of a reaction

3 use experimental data to calculate the rate of a reaction

8.2 Effect of temperature on reaction rates and the concept of activation energy

Learning outcomes

Candidates should be able to:

1 define activation energy, E_A, as the minimum energy required for a collision to be effective

2 sketch and use the Boltzmann distribution to explain the significance of activation energy

3 explain qualitatively, in terms both of the Boltzmann distribution and of frequency of effective collisions, the effect of temperature change on the rate of a reaction

8.3 Homogeneous and heterogeneous catalysts

Learning outcomes

Candidates should be able to:

1 explain and use the terms catalyst and catalysis

   (a) explain that, in the presence of a catalyst, a reaction has a different mechanism, i.e. one of lower activation energy

   (b) explain this catalytic effect in terms of the Boltzmann distribution

   (c) construct and interpret a reaction pathway diagram, for a reaction in the presence and absence of an effective catalyst

Inorganic chemistry

9 The Periodic Table: chemical periodicity

9.1 Periodicity of physical properties of the elements in Period 3

Learning outcomes

Candidates should be able to:

1 describe qualitatively (and indicate the periodicity in) the variations in atomic radius, ionic radius, melting point and electrical conductivity of the elements

2 explain the variation in melting point and electrical conductivity in terms of the structure and bonding of the elements

9.2 Periodicity of chemical properties of the elements in Period 3

Learning outcomes

Candidates should be able to:

1 describe, and write equations for, the reactions of the elements with oxygen (to give Na_2O, MgO, Al_2O_3, P_4O_{10}, SO_2), chlorine (to give $NaCl$, $MgCl_2$, $AlCl_3$, $SiCl_4$, PCl_5) and water (Na and Mg only)

2 state and explain the variation in the oxidation number of the oxides (Na_2O, MgO, Al_2O_3, P_4O_{10}, SO_2 and SO_3 only) and chlorides ($NaCl$, $MgCl_2$, $AlCl_3$, $SiCl_4$, PCl_5 only) in terms of their outer shell (valence shell) electrons

3 describe, and write equations for, the reactions, if any, of the oxides Na_2O, MgO, Al_2O_3, SiO_2, P_4O_{10}, SO_2 and SO_3 with water including the likely pHs of the solutions obtained

4 describe, explain, and write equations for, the acid / base behaviour of the oxides Na_2O, MgO, Al_2O_3, P_4O_{10}, SO_2 and SO_3 and the hydroxides $NaOH$, $Mg(OH)_2$, $Al(OH)_3$ including, where relevant, amphoteric behaviour in reactions with acids and bases (sodium hydroxide only)

5 describe, explain, and write equations for, the reactions of the chlorides $NaCl$, $MgCl_2$, $AlCl_3$, $SiCl_4$, PCl_5 with water including the likely pHs of the solutions obtained

6 explain the variations and trends in 9.2.2, 9.2.3, 9.2.4 and 9.2.5 in terms of bonding and electronegativity

7 suggest the types of chemical bonding present in the chlorides and oxides from observations of their chemical and physical properties

9.3 Chemical periodicity of other elements

Learning outcomes

Candidates should be able to:

1 predict the characteristic properties of an element in a given group by using knowledge of chemical periodicity

2 deduce the nature, possible position in the Periodic Table and identity of unknown elements from given information about physical and chemical properties

10   Group 2

10.1   Similarities and trends in the properties of the Group 2 metals, magnesium to barium, and their compounds

Learning outcomes

Candidates should be able to:

1   describe, and write equations for, the reactions of the elements with oxygen, water and dilute hydrochloric and sulfuric acids

2   describe, and write equations for, the reactions of the oxides, hydroxides and carbonates with water and dilute hydrochloric and sulfuric acids

3   describe, and write equations for, the thermal decomposition of the nitrates and carbonates, to include the trend in thermal stabilities

4   describe, and make predictions from, the trends in physical and chemical properties of the elements involved in the reactions in 10.1.1 and the compounds involved in 10.1.2, 10.1.3 and 10.1.5

5   state the variation in the solubilities of the hydroxides and sulfates

11   Group 17

11.1   Physical properties of the Group 17 elements

Learning outcomes

Candidates should be able to:

1   describe the colours and the trend in volatility of chlorine, bromine and iodine

2   describe and explain the trend in the bond strength of the halogen molecules

3   interpret the volatility of the elements in terms of instantaneous dipole–induced dipole forces

11.2   The chemical properties of the halogen elements and the hydrogen halides

Learning outcomes

Candidates should be able to:

1   describe the relative reactivity of the elements as oxidising agents

2   describe the reactions of the elements with hydrogen and explain their relative reactivity in these reactions

3   describe the relative thermal stabilities of the hydrogen halides and explain these in terms of bond strengths

11.3   Some reactions of the halide ions

Learning outcomes

Candidates should be able to:

1   describe the relative reactivity of halide ions as reducing agents

2   describe and explain the reactions of halide ions with:

(a)   aqueous silver ions followed by aqueous ammonia (the formation and formula of the $[Ag(NH_3)_2]^+$ complex is not required)

(b)   concentrated sulfuric acid, to include balanced chemical equations

11.4   The reactions of chlorine

Learning outcomes

Candidates should be able to:

1   describe and interpret, in terms of changes in oxidation number, the reaction of chlorine with cold and with hot aqueous sodium hydroxide and recognise these as disproportionation reactions

2   explain, including by use of an equation, the use of chlorine in water purification to include the production of the active species $HOCl$ and ClO which kill bacteria.

12   Nitrogen and sulfur

12.1   Nitrogen and sulfur

Learning outcomes

Candidates should be able to:

1   explain the lack of reactivity of nitrogen, with reference to triple bond strength and lack of polarity

2   describe and explain:

(a) the basicity of ammonia, using the Brønsted–Lowry theory

(b) the structure of the ammonium ion and its formation by an acid–base reaction

(c) the displacement of ammonia from ammonium salts by an acid–base reaction

3   state and explain the natural and man-made occurrences of oxides of nitrogen and their catalytic removal from the exhaust gases of internal combustion engines

4   understand that atmospheric oxides of nitrogen (NO and NO_2) can react with unburned hydrocarbons to form peroxyacetyl nitrate, PAN, which is a component of photochemical smog

5   describe the role of NO and NO_2 in the formation of acid rain both directly and in their catalytic role in the oxidation of atmospheric sulfur dioxide
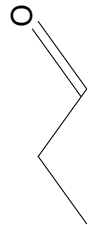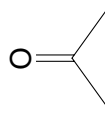
Organic chemistry

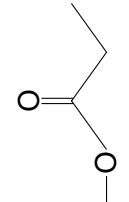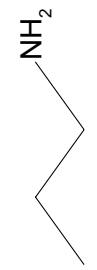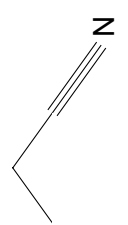## 13   An introduction to AS Level organic chemistry
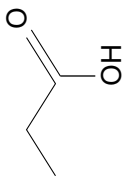
In this syllabus the following conventions are used:

X to represent a halogen atom

R and R´ to represent alkyl groups (or in some circumstances a hydrogen atom); R and R´ can be the same or different depending on the molecule.

class of compound	name of functional group	structural formula of functional group	displayed formula	skeletal formula	name
alkene	alkene	R₂C=CR₂ (R/C=C\R)	(displayed formula of propene)	(skeletal formula)	propene
halogenoalkane (primary, secondary and tertiary)	halogen	R—X	(displayed formula)	(skeletal formula, X)	1-chloropropane (when X is chlorine)
alcohol (primary, secondary and tertiary)	hydroxy	R—OH	(displayed formula)	(skeletal formula, OH)	propan-1-ol
aldehyde	carbonyl	R—C(=O)H	(displayed formula)	(skeletal formula, O)	propanal
ketone	carbonyl	R(R´)C=O	(displayed formula)	(skeletal formula, O)	propanone

class of compound	name of functional group	structural formula of functional group	displayed formula	skeletal formula	name
carboxylic acid	carboxyl	$R{-}C({=}O){-}OH$			propanoic acid
ester	ester	$R{-}C({=}O){-}O{-}R$			methyl propanoate
amine (primary only)	amine	$R{-}NH_2$			propylamine
nitrile	nitrile	$R{-}C{\equiv}N$			propanenitrile

13.1   Formulae, functional groups and the naming of organic compounds

Learning outcomes

Candidates should be able to:

1   define the term hydrocarbon as a compound made up of C and H atoms only

2   understand that alkanes are simple hydrocarbons with no functional group

3   understand that the compounds in the table on page 26 and 27 contain a functional group which dictates their physical and chemical properties

4   interpret and use the general, structural, displayed and skeletal formulae of the classes of compound stated in the table on page 26 and 27

5   understand and use systematic nomenclature of simple aliphatic organic molecules with functional groups detailed in the table on page 26 and 27, up to six carbon atoms (six plus six for esters, straight chains only for esters and nitriles)

6   deduce the molecular and/or empirical formula of a compound, given its structural, displayed or skeletal formula

13.2   Characteristic organic reactions

Learning outcomes

Candidates should be able to:

1   interpret and use the following terminology associated with types of organic compounds and reactions:

(a) homologous series

(b) saturated and unsaturated

(c) homolytic and heterolytic fission

(d) free radical, initiation, propagation, termination (the use of arrows to show movement of single electrons is not required)

(e) nucleophile, electrophile, nucleophilic, electrophilic

(f) addition, substitution, elimination, hydrolysis, condensation

(g) oxidation and reduction

(in equations for organic redox reactions, the symbol [O] can be used to represent one atom of oxygen from an oxidising agent and the symbol [H] one atom of hydrogen from a reducing agent)

2   understand and use the following terminology associated with types of organic mechanisms:

(a) free-radical substitution

(b) electrophilic addition

(c) nucleophilic substitution

(d) nucleophilic addition

(in organic reaction mechanisms, the use of curly arrows to represent movement of electron pairs is expected; the arrow should begin at a bond or a lone pair of electrons)

13.3   Shapes of organic molecules; σ and π bonds

Learning outcomes

Candidates should be able to:

1   describe organic molecules as either straight-chained, branched or cyclic

2   describe and explain the shape of, and bond angles in, molecules containing sp, sp^2 and sp^3 hybridised atoms

3   describe the arrangement of σ and π bonds in molecules containing sp, sp^2 and sp^3 hybridised atoms

4   understand and use the term planar when describing the arrangement of atoms in organic molecules, for example ethene