



SUBJECT CONTENT

CORE IDEA 1 – MATTER

1. Atomic Structure

Content

- The nucleus of the atom: neutrons and protons, isotopes, proton and nucleon numbers
- Electrons: electronic energy levels, ionisation energies, atomic orbitals, extranuclear structure

Learning Outcomes

Candidates should be able to:

- identify and describe protons, neutrons and electrons in terms of their relative charges and relative masses
- deduce the behaviour of beams of protons, neutrons and electrons in an electric field
- describe the distribution of mass and charges within an atom
- deduce the numbers of protons, neutrons and electrons present in both atoms and ions given proton and nucleon numbers (and charge)
- describe the contribution of protons and neutrons to atomic nuclei in terms of proton number and nucleon number
 - distinguish between isotopes on the basis of different numbers of neutrons present
- describe the number and relative energies of the s, p and d orbitals for the principal quantum numbers 1, 2 and 3 and also the 4s and 4p orbitals
- describe the shapes of s, p and d orbitals (see also Section 13)
[knowledge of wave functions is **not** required]
- state the electronic configuration of atoms and ions given the proton number (and charge)
- explain the factors influencing the ionisation energies of elements (see the *Data Booklet*) (see also Section 5)
- deduce the electronic configurations of elements from successive ionisation energy data
- interpret successive ionisation energy data of an element in terms of the position of that element within the Periodic Table

CORE IDEA 2 – STRUCTURE AND PROPERTIES

2. Chemical Bonding

Content

- Ionic bonding, metallic bonding, covalent bonding and co-ordinate (dative covalent) bonding
- Shapes of simple molecules and bond angles
- Bond polarities and polarity of molecules
- Intermolecular forces, including hydrogen bonding
- Bond energies and bond lengths
- Lattice structure of solids
- Bonding and physical properties

Learning Outcomes

Candidates should be able to:

- (a) show understanding that all chemical bonds are electrostatic in nature and describe:
- ionic bond as the electrostatic attraction between oppositely charged ions
 - covalent bond as the electrostatic attraction between a shared pair of electrons and positively charged nuclei
 - metallic bond as the electrostatic attraction between a lattice of positive ions and delocalised electrons
- (b) describe, including the use of 'dot-and-cross' diagrams,
- ionic bonding as in sodium chloride and magnesium oxide
 - covalent bonding as in hydrogen; oxygen; nitrogen; chlorine; hydrogen chloride; carbon dioxide; methane; ethene
 - co-ordinate (dative covalent) bonding, as in formation of the ammonium ion and in the Al_2Cl_6 molecule
- (c) describe covalent bonding in terms of orbital overlap (limited to s and p orbitals only), giving σ and π bonds (see also Section 11.1)
- (d) explain the shapes of, and bond angles in, molecules such as BF_3 (trigonal planar); CO_2 (linear); CH_4 (tetrahedral); NH_3 (trigonal pyramidal); H_2O (bent); SF_6 (octahedral) by using the Valence Shell Electron Pair Repulsion theory
- (e) predict the shapes of, and bond angles in, molecules analogous to those specified in (d)
- (f) explain and deduce bond polarity using the concept of electronegativity [quantitative treatment of electronegativity is **not** required]
- (g) deduce the polarity of a molecule using bond polarity and its molecular shape (analogous to those specified in (d));
- (h) describe the following forces of attraction (electrostatic in nature):
- intermolecular forces, based on permanent and induced dipoles, as in $CHCl_3(l)$; $Br_2(l)$ and the liquid noble gases
 - hydrogen bonding, using ammonia and water as examples of molecules containing $-NH$ and $-OH$ groups
- (i) outline the importance of hydrogen bonding to the physical properties of substances, including ice and water

- (j) explain the terms *bond energy* and *bond length* for covalent bonds
- (k) compare the reactivities of covalent bonds in terms of bond energy, bond length and bond polarity
- (l) describe, in simple terms, the lattice structure of a crystalline solid which is:
- ionic, as in sodium chloride and magnesium oxide
 - simple molecular, as in iodine
 - giant molecular, as in graphite and diamond
 - hydrogen-bonded, as in ice
 - metallic, as in copper
- [the concept of the 'unit cell' is **not** required]
- (m) describe, interpret and/or predict the effect of different types of structure and bonding on the physical properties of substances
- (n) suggest the type of structure and bonding present in a substance from given information

3. The Gaseous State

Content

- Ideal gas behaviour and deviations from it
- $pV = nRT$ and its use in determining a value for M_r
- Dalton's Law and its use in determining the partial pressures of gases in a mixture

Learning Outcomes

Candidates should be able to:

- (a) state the basic assumptions of the kinetic theory as applied to an ideal gas
- (b) explain qualitatively in terms of intermolecular forces and molecular size:
- the conditions necessary for a gas to approach ideal behaviour
 - the limitations of ideality at very high pressures and very low temperatures
- (c) state and use the general gas equation $pV = nRT$ in calculations, including the determination of M_r
- (d) use Dalton's Law to determine the partial pressures of gases in a mixture (see also Section 9)

4. Theories of Acids and Bases

Content

- Arrhenius, Brønsted-Lowry and Lewis theories of acids and bases

Learning Outcomes

Candidates should be able to:

- (a) show understanding of, and apply the Arrhenius theory of acids and bases
- (b) show understanding of, and apply the Brønsted-Lowry theory of acids and bases, including the concept of conjugate acids and conjugate bases
- (c) show understanding of, and apply the Lewis theory of acids and bases (including non-aqueous system, e.g. reaction between BF_3 and NH_3)

5. The Periodic Table

Content

- Periodicity of atomic and physical properties of the elements: variation with proton number across the third period (sodium to chlorine) and down the group (Group 2 and Group 17) of:
 - (i) electronic configuration
 - (ii) atomic radius and ionic radius
 - (iii) ionisation energy
 - (iv) electronegativity
 - (v) melting point
 - (vi) electrical conductivity
- Periodicity of chemical properties of the elements in the third period:
 - (i) variation in oxidation number and bonding of the oxides (sodium to sulfur only) and of the chlorides (sodium to phosphorus only)
 - (ii) reactions of these oxides and chlorides with water
 - (iii) acid/base behaviour of these oxides and the corresponding hydroxides
- Periodicity of chemical properties of the elements down the group (Group 2 and Group 17):
 - (i) as reducing agents (Group 2) and oxidising agents (Group 17)
 - (ii) thermal stability of Group 2 carbonates and Group 17 hydrides

Learning Outcomes

Trends and variations in atomic and physical properties

For elements in the third period (sodium to chlorine), and in Group 2 (magnesium to barium) and Group 17 (chlorine to iodine) candidates should be able to:

- (a) recognise variation in the electronic configurations across a Period and down a Group
- (b) describe and explain qualitatively the general trends and variations in atomic radius, ionic radius, first ionisation energy and electronegativity:
 - (i) across a Period in terms of shielding and nuclear charge
 - (ii) down a Group in terms of increasing number of electronic shells, shielding and nuclear charge
- (c) interpret the variation in melting point and in electrical conductivity across a Period in terms of structure and bonding in the elements (metallic, giant molecular, or simple molecular)
- (d) describe and explain the trend in volatility of the Group 17 elements in terms of instantaneous dipole-induced dipole attraction

Trends and variations in chemical properties

For elements in the third period (sodium to chlorine) candidates should be able to:

- (e)
 - (i) state and explain the variation in the highest oxidation number of the elements in oxides (for Na_2O ; MgO ; Al_2O_3 ; SiO_2 ; P_4O_{10} ; SO_3) and chlorides (for NaCl ; MgCl_2 ; AlCl_3 ; SiCl_4 ; PCl_5)
 - (ii) state and explain the variation in bonding in oxides and chlorides in terms of electronegativity (with the exception of AlCl_3)
 - (iii) describe the reactions of the oxides with water (for Na_2O ; MgO ; Al_2O_3 ; SiO_2 ; P_4O_{10} ; SO_3)
 - (iv) describe and explain the acid/base behaviour of oxides (for Na_2O ; MgO ; Al_2O_3 ; SiO_2 ; P_4O_{10} ; SO_3) and hydroxides (for NaOH ; $\text{Mg}(\text{OH})_2$; $\text{Al}(\text{OH})_3$), including, where relevant, amphoteric behaviour in reaction with sodium hydroxide (only) and acids
 - (v) describe and explain the reactions of the chlorides with water (for NaCl ; MgCl_2 ; AlCl_3 ; SiCl_4 ; PCl_5)
 - (vi) suggest the types of structure and bonding present in the oxides and chlorides from observations of their chemical and physical properties

For elements in Group 2 (magnesium to barium) and Group 17 (chlorine to iodine) candidates should be able to:

- (f) describe and deduce from E^\ominus values the relative reactivity of elements of:
 - (i) Group 2 as reducing agents
 - (ii) Group 17 as oxidising agents
- (g) describe and explain the trend in thermal stability of:
 - (i) Group 2 carbonates in terms of the charge density of the cation and the polarisability of the large anion
 - (ii) Group 17 hydrides in terms of bond energies

In addition, candidates should be able to:

- (h) predict the characteristic properties of an element in a given Group by using knowledge of chemical periodicity
- (i) deduce the nature, possible position in the Periodic Table, and identity of unknown elements from given information of physical and chemical properties

CORE IDEA 3 – TRANSFORMATION

6. The Mole Concept and Stoichiometry

Content

- Relative masses of atoms and molecules
- The mole, the Avogadro constant
- The calculation of empirical and molecular formulae
- Reacting masses and volumes (of solutions and gases)

Learning Outcomes

[the term relative formula mass or M_r will be used for ionic compounds]

Candidates should be able to:

- (a) define the terms *relative atomic*, *isotopic*, *molecular* and *formula mass*
- (b) define the term *mole* in terms of the Avogadro constant
- (c) calculate the relative atomic mass of an element given the relative abundances of its isotopes
- (d) define the terms *empirical* and *molecular formula*
- (e) calculate empirical and molecular formulae using combustion data or composition by mass
- (f) write and/or construct balanced equations
- (g) perform calculations, including use of the mole concept, involving:
 - (i) reacting masses (from formulae and equations)
 - (ii) volumes of gases (e.g. in the burning of hydrocarbons)
 - (iii) volumes and concentrations of solutions[when performing calculations, candidates' answers should reflect the number of significant figures given or asked for in the question]
- (h) deduce stoichiometric relationships from calculations such as those in (g)

7. Chemical Energetics: Thermochemistry and Thermodynamics (Gibbs Free Energy and Entropy)

Content

- Enthalpy changes: ΔH , of formation; combustion; hydration; solution; neutralisation; atomisation; bond energy; lattice energy; electron affinity
- Hess' Law, including Born-Haber cycles
- Entropy and Free Energy

Learning Outcomes

Candidates should be able to:

- (a) explain that most chemical reactions are accompanied by energy changes, principally in the form of heat usually associated with the breaking and forming of chemical bonds; the reaction can be exothermic (ΔH negative) or endothermic (ΔH positive)
- (b) construct and interpret an energy profile diagram, in terms of the enthalpy change of the reaction and of the activation energy (see also Section 8)
- (c) explain and use the terms:
- enthalpy change of reaction* and *standard conditions*, with particular reference to: formation; combustion; hydration; solution; neutralisation; atomisation
 - bond energy* (ΔH positive, i.e. bond breaking) (see also Section 2)
 - lattice energy* (ΔH negative, i.e. gaseous ions to solid lattice)
- (d) calculate enthalpy changes from appropriate experimental results, including the use of the relationship: $\text{heat change} = mc\Delta T$
- (e) explain, in qualitative terms, the effect of ionic charge and of ionic radius on the numerical magnitude of a lattice energy
- (f) apply Hess' Law to construct simple energy cycles, e.g. Born-Haber cycle, and carry out calculations involving such cycles and relevant energy terms (including ionisation energy and electron affinity), with particular reference to:
- determining enthalpy changes that cannot be found by direct experiment, e.g. an enthalpy change of formation from enthalpy changes of combustion
 - the formation of a simple ionic solid and of its aqueous solution
 - average bond energies
- (g) explain and use the term *entropy*
- (h) discuss the effects on the entropy of a chemical system by the following:
- change in temperature
 - change in phase
 - change in the number of particles (especially for gaseous systems)
 - mixing of particles
- [quantitative treatment is **not** required]
- (i) predict whether the entropy change for a given process or reaction is positive or negative
- (j) state and use the equation involving *standard Gibbs free energy change of reaction*, ΔG^\ominus :
$$\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus$$

[the calculation of standard entropy change, ΔS^\ominus , for a reaction using standard entropies, S^\ominus , is **not** required]
- (k) state whether a reaction or process will be spontaneous by using the sign of ΔG^\ominus
- (l) understand the limitations in the use of ΔG^\ominus to predict the spontaneity of a reaction
- (m) predict the effect of temperature change on the spontaneity of a reaction, given standard enthalpy and entropy changes

8. Reaction Kinetics

Content

- Simple rate equations; orders of reaction; rate constants
- Concept of activation energy
- Effect of concentration, temperature, and catalysts on reaction rate
- Homogeneous and heterogeneous catalysis
- Enzymes as biological catalysts

Learning Outcomes

Candidates should be able to:

- (a) explain and use the terms: *rate of reaction*; *rate equation*; *order of reaction*; *rate constant*; *half-life of a reaction*; *rate-determining step*; *activation energy*; *catalysis*
- (b) construct and use rate equations of the form $\text{rate} = k[\text{A}]^m[\text{B}]^n$ (limited to simple cases of single-step reactions and of multi-step processes with a rate-determining step, for which m and n are 0, 1 or 2), including:
- deducing the order of a reaction by the initial rates method
 - justifying, for zero- and first-order reactions, the order of reaction from concentration-time graphs
 - verifying that a suggested reaction mechanism is consistent with the observed kinetics
 - predicting the order that would result from a given reaction mechanism
 - calculating an initial rate using concentration data
- [integrated forms of rate equations are **not** required]
- (c) (i) show understanding that the half-life of a first-order reaction is independent of concentration
(ii) use the half-life of a first-order reaction in calculations
- (d) calculate a rate constant using the initial rates method
- (e) devise a suitable experimental technique for studying the rate of a reaction, from given information
- (f) explain qualitatively, in terms of frequency of collisions, the effect of concentration changes on the rate of a reaction
- (g) show understanding, including reference to the Boltzmann distribution, of what is meant by the term *activation energy*
- (h) explain qualitatively, in terms both of the Boltzmann distribution and of collision frequency, the effect of temperature change on a rate constant (and hence, on the rate) of a reaction
- (i) (i) explain that, in the presence of a catalyst, a reaction has a different mechanism, i.e. one of lower activation energy, giving a larger rate constant
(ii) interpret this catalytic effect on a rate constant in terms of the Boltzmann distribution
- (j) outline the different modes of action of homogeneous and heterogeneous catalysis, including:
- the Haber process
 - the catalytic removal of oxides of nitrogen in the exhaust gases from car engines (see also Section 11.3)
 - the catalytic role of atmospheric oxides of nitrogen in the oxidation of atmospheric sulfur dioxide
 - catalytic role of Fe^{2+} in the $\text{I}^-/\text{S}_2\text{O}_8^{2-}$ reaction

- (k) describe enzymes as biological catalysts which may have specific activity
- (l) explain the relationship between substrate concentration and the rate of an enzyme-catalysed reaction in biochemical systems

9. Chemical Equilibria

Content

- Chemical equilibria: reversible reactions; dynamic equilibrium
 - (i) factors affecting chemical equilibria
 - (ii) equilibrium constants
 - (iii) the Haber process

Learning Outcomes

Candidates should be able to:

- (a) explain, in terms of rates of the forward and reverse reactions, what is meant by a *reversible reaction* and *dynamic equilibrium*
- (b) state Le Chatelier's Principle and apply it to deduce qualitatively (from appropriate information) the effects of changes in concentration, pressure or temperature, on a system at equilibrium
- (c) deduce whether changes in concentration, pressure or temperature or the presence of a catalyst affect the value of the equilibrium constant for a reaction
- (d) deduce expressions for equilibrium constants in terms of concentrations, K_c , and partial pressures, K_p [treatment of the relationship between K_p and K_c is **not** required]
- (e) calculate the values of equilibrium constants in terms of concentrations or partial pressures from appropriate data
- (f) calculate the quantities present at equilibrium, given appropriate data (such calculations will not require the solving of quadratic equations)
- (g) show understanding that the position of equilibrium is dependent on the standard Gibbs free energy change of reaction, ΔG^\ominus
[Quantitative treatment is **not** required]
- (h) describe and explain the conditions used in the Haber process, as an example of the importance of an understanding of chemical equilibrium in the chemical industry

EXTENSION TOPICS

10. Chemistry of Aqueous Solutions

10.1 Acid-base Equilibria

This topic deals with acid-base equilibria in aqueous solution. Hence the Brønsted-Lowry definitions of acid/base are primarily used in the understanding of pH of solutions. (See Section 4)

Content

- Acid dissociation constants, K_a and the use of pK_a
- Base dissociation constants, K_b and the use of pK_b
- The ionic product of water, K_w
- pH: choice of pH indicators
- Buffer solutions

Learning Outcomes

Candidates should be able to:

- explain qualitatively the differences in behaviour between strong and weak acids and bases in terms of the extent of dissociation
- explain the terms pH; K_a ; pK_a ; K_b ; pK_b ; K_w and apply them in calculations, including the relationship $K_w = K_a K_b$
- calculate $[H^+(aq)]$ and pH values for strong acids, weak monobasic (monoprotic) acids, strong bases, and weak monoacidic bases
[Calculations involving weak acids/bases will **not** require solving of quadratic equations]
- describe the changes in pH during acid-base titrations and explain these changes in terms of the strengths of the acids and bases
- explain the choice of suitable indicators for acid-base titrations, given appropriate data
- explain how buffer solutions control pH
 - describe and explain their uses, including the role of H_2CO_3/HCO_3^- in controlling pH in blood
- calculate the pH of buffer solutions, given appropriate data

10.2 Solubility Equilibria

Content

- Solubility product; the common ion effect and complex ion formation

Learning Outcomes

Candidates should be able to:

- show understanding of, and apply, the concept of solubility product, K_{sp}
- calculate K_{sp} from concentrations and *vice versa*

- (c) discuss the effects on the solubility of ionic salts by the following:
- common ion effect
 - formation of complex ion, as exemplified by the reactions of halide ions with aqueous silver ions followed by aqueous ammonia (see also Section 13)

11. Organic Chemistry

Preamble

Although there are features of organic chemistry topics that are distinctive, it is intended that appropriate cross-references with other sections/topics in the syllabus should be made.

In their study of organic chemistry, candidates may wish to group the organic reactions in terms of the mechanisms in the syllabus where possible. Candidates may wish to compare and contrast the different mechanisms.

When describing preparative reactions, candidates will be expected to quote the reagents, e.g. aqueous NaOH, the essential practical conditions, e.g. reflux, high temperature and pressure, and the identity of each of the major products. Detailed conditions involving specific temperature and pressure values are **not** required.

Detailed knowledge of practical procedures is also **not** required; however, candidates may be expected to suggest (from their knowledge of the reagents, essential conditions and products) what steps may be needed to purify/extract a required product from the reaction mixture. In equations for organic redox reactions, the symbols [O] and [H] are acceptable.

11.1 Introduction

Content

- Empirical, molecular and structural formulae
- Functional groups and the naming of organic compounds
- Common terms for organic reactions and reactivities
- Shapes of organic molecules; σ and π bonds

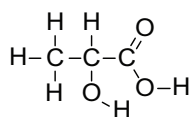
Candidates are expected to be able to interpret and use the following types of representations in the description of organic molecules. The examples given are for the compound (+)-lactic acid.

Empirical Formula: simplest ratio of the number of atoms of the elements present in one molecule, e.g. CH₂O

Molecular Formula: actual number of atoms of the elements present in one molecule, e.g. C₃H₆O₃

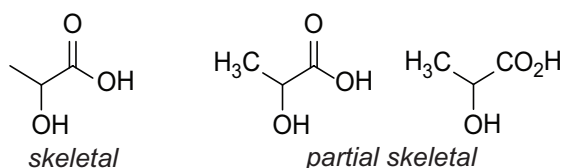
Structural Formula: shows how the constituent atoms of a molecule are joined together with minimal detail, using conventional groups, for an unambiguous structure, e.g. CH₃CH(OH)CO₂H


Full Structural or Displayed Formula: detailed structure of molecule showing the relative placing of atoms and the number of bonds between them, e.g.



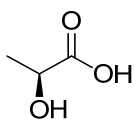
Skeletal Formula: simplified representation of an organic formula derived from the structural formula by removing hydrogen atoms (and their associated bonds) and carbon atoms from alkyl chains, leaving just the carbon-carbon bonds in the carbon skeleton and the associated functional groups

Skeletal or partial skeletal representations may be used in question papers and are acceptable in candidates' answers where they are unambiguous, e.g.

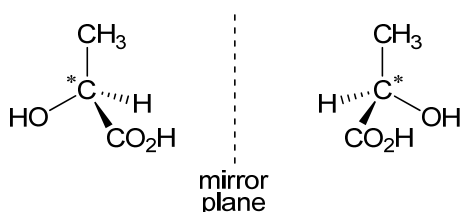


The convention  for representing the aromatic ring is preferred.

Stereochemical Formula: show spatial arrangement of bonds, atoms and groups in molecule in 3-D, e.g.



When drawing a pair of optical isomers, candidates should indicate the three-dimensional structures according to the convention used in the example below.



Candidates are expected to be able to interpret and use the curly arrow notation to represent the movement of electrons in organic reaction mechanisms.

For movement of a *pair* of electrons (**full arrow**) : $Z:\ominus + \overset{\delta+}{X}-\overset{\delta-}{Y} \longrightarrow Z-X + \ominus:Y$

For movement of a *single* unpaired electron (**half arrow**) : $Z\cdot + \overset{\delta+}{X}-\overset{\delta-}{Y} \longrightarrow Z-X + \cdot Y$

Learning Outcomes

Candidates should be able to:

- (a) interpret, and use the nomenclature, general formulae and displayed formulae of the following classes of compound:
- hydrocarbons (alkanes, alkenes and arenes)
 - halogen derivatives (halogenoalkanes and halogenoarenes)
 - hydroxyl compounds (alcohols and phenols)
 - carbonyl compounds (aldehydes and ketones)
 - carboxylic acids and derivatives (acyl chlorides and esters)
 - nitrogen compounds (amines, amides, amino acids and nitriles)
- (b) interpret, and use the following terminology associated with organic reactions:
- functional group
 - degree of substitution: primary, secondary, tertiary, quaternary
 - homolytic and heterolytic fission
 - carbocation
 - free radical, initiation, propagation, termination
 - electrophile (Lewis acid), nucleophile (Lewis base)
 - addition, substitution, elimination, condensation, hydrolysis
 - oxidation and reduction
- [in equations for organic redox reactions, the symbols [O] and [H] are acceptable]
- (c) interpret, and use the following terminology associated with organic reactivities:
- delocalisation
 - electronic effect (electron-donating and electron-withdrawing effect)
 - steric effect (steric hindrance)
- (d) describe sp^3 hybridisation, as in ethane molecule, sp^2 hybridisation, as in ethene and benzene molecules, and sp hybridisation, as in ethyne molecule
- (e) explain the shapes of, and bond angles in, the ethane, ethene, benzene, and ethyne molecules in relation to σ and π carbon-carbon bonds
- (f) predict the shapes of, and bond angles in, molecules analogous to those specified in (e)
- (g) apply (b) and (c) to the understanding of mechanisms in terms of organic structure and bonding
- (h) recognise that the mechanisms of polar reactions involve the flow of electrons from electron-rich to electron-poor sites

11.2 Isomerism

Content

- Isomerism: *constitutional (structural)*; *cis-trans*; *enantiomerism*

Learning Outcomes

Candidates should be able to:

- describe constitutional (structural) isomerism
- describe *cis-trans* isomerism in alkenes, and explain its origin in terms of restricted rotation due to the presence of π bonds
[use of *E*, *Z* nomenclature is **not** required]
- explain what is meant by a chiral centre
- deduce whether a given molecule is chiral based on the presence or absence of chiral centres and/or a plane of symmetry
- recognise that an optically active sample rotates plane-polarised light and contains chiral molecules
- recognise that enantiomers have identical physical properties except in the direction in which they rotate plane-polarised light
[usage of the term diastereomers is **not** required]
- recognise that enantiomers have identical chemical properties except in their interactions with another chiral molecule
- recognise that different stereoisomers exhibit different biological properties, for example in drug action
- deduce the possible isomers for an organic molecule of known molecular formula
- identify chiral centres and/or *cis-trans* isomerism in a molecule of given structural formula

In each of the sections below, 11.3 to 11.8, candidates will be expected to be able to predict the reaction products of a given compound in reactions that are chemically similar to those specified.

11.3 Hydrocarbons

Content

- Alkanes (exemplified by ethane)
 - free-radical substitution reactions
- Alkenes (exemplified by ethene)
 - electrophilic addition, including Markovnikov's rule
 - reduction and oxidation reactions
- Arenes (exemplified by benzene and methylbenzene)
 - influence of delocalised π electrons on structure and properties
 - electrophilic substitution reactions
 - oxidation of side-chain
- Hydrocarbons as fuels

Learning Outcomes

Candidates should be able to:

- (a) explain the general unreactivity of alkanes, including towards polar reagents
- (b) describe the chemistry of alkanes as exemplified by the following reactions of ethane:
 - (i) combustion
 - (ii) free-radical substitution by chlorine and by bromine
- (c) describe the mechanism of free-radical substitution with particular reference to the initiation, propagation and termination reactions
- (d) explain the general reactivity of alkenes towards electrophilic reagents/electrophiles
- (e) describe the chemistry of alkenes as exemplified, where relevant, by the following reactions of ethene:
 - (i) electrophilic addition of water/steam, hydrogen halides and halogens
 - (ii) reduction via catalytic hydrogenation (catalytic addition of hydrogen; see also 8(j))
 - (iii) oxidation by cold, alkaline solution of manganate(VII) ions to form the diol
 - (iv) oxidation by hot, acidified solution of manganate(VII) ions leading to the rupture of the carbon-to-carbon double bond in order to determine the position of alkene linkages in larger molecules
- (f) describe the mechanism of electrophilic addition in alkenes, using bromine with ethene as an example
- (g) apply Markovnikov's rule to the addition of hydrogen halides to unsymmetrical alkenes, and explain the composition of products in terms of the stability of carbocation intermediates
- (h) explain, in terms of delocalisation of π electrons, the difference between benzene and alkene:
 - (i) reactivity towards electrophiles
 - (ii) preference of benzene to undergo substitution rather than addition reaction
- (i) describe the chemistry of the benzene ring as exemplified by the following reactions of benzene and methylbenzene:
 - (i) electrophilic substitution reactions with chlorine and with bromine
(recognise the use of Lewis acid as catalysts; see also Section 4)
 - (ii) nitration with concentrated nitric acid
(recognise concentrated sulfuric acid as a Brønsted-Lowry acid catalyst; see also Section 4)
 - (iii) Friedel-Crafts alkylation with halogenoalkanes
(recognise the use of Lewis acid as catalysts; see also Section 4)
- (j)
 - (i) describe the mechanism of electrophilic substitution in arenes, using the mono-bromination of benzene as an example
 - (ii) describe the effect of the delocalisation of electrons in arenes in such reactions
- (k) describe the chemistry of the alkyl side-chain of benzene ring as exemplified by the following reactions of methylbenzene:
 - (i) free-radical substitution by chlorine and by bromine
 - (ii) complete oxidation to give benzoic acid
- (l) predict whether halogenation will occur in the side-chain or aromatic nucleus in arenes depending on reaction conditions
- (m) apply the knowledge of positions of substitution in the electrophilic substitution reactions of mono-substituted arenes

- (n) recognise the environmental consequences of:
- carbon monoxide, oxides of nitrogen and unburnt hydrocarbons arising from the internal combustion engine and of their catalytic removal
 - gases that contribute to the enhanced greenhouse effect
- (o) recognise that petroleum, a chemical feedstock, is a finite resource and the importance of recycling

11.4 Halogen Derivatives

Content

- Halogenoalkanes
 - nucleophilic substitution
 - elimination
- Relative strength of the C-Hal bond
- Unreactivity of halogenoarenes

Learning Outcomes

Candidates should be able to:

- (a) recall the chemistry of halogenoalkanes as exemplified by:
- the following nucleophilic substitution reactions of bromoethane: hydrolysis; formation of nitriles; formation of primary amines by reaction with ammonia
 - the elimination of hydrogen bromide from 2-bromopropane
- (b) describe and explain the mechanisms of nucleophilic substitutions in halogenoalkanes:
- S_N1 , in terms of stability of the carbocation intermediates
 - S_N2 , in terms of steric hindrance in the halogenoalkanes
- (c) explain the stereochemical outcome in nucleophilic substitution involving optically active substrates:
- inversion of configuration in S_N2 mechanism
 - racemisation in S_N1 mechanism
- (d) interpret the different reactivities of halogenoalkanes, with particular reference to hydrolysis, and to the relative strengths of the C-Hal bonds
- (e) explain the unreactivity of chlorobenzene compared to halogenoalkanes towards nucleophilic substitution, in terms of the delocalisation of the lone pair of electrons on the halogen and steric hindrance
- (f) suggest characteristic reactions to differentiate between:
- different halogenoalkanes
 - halogenoalkanes and halogenoarenes
- e.g. hydrolysis, followed by testing of the halide ions
- (g) explain the uses of fluoroalkanes and fluorohalogenoalkanes in terms of their relative chemical inertness
- (h) recognise the concern about the effect of chlorofluoroalkanes (CFCs) on the ozone layer
[the mechanistic details of how CFCs deplete the ozone layer are **not** required]

11.5 Hydroxy Compounds

Content

- Alcohols (exemplified by ethanol)
 - formation of halogenoalkanes
 - reaction with sodium; oxidation; dehydration
 - the tri-iodomethane test
- Phenol
 - its acidity; reaction with bases and sodium
 - nitration of, and bromination of, the aromatic ring

Learning Outcomes

Candidates should be able to:

- recall the chemistry of alcohols, exemplified by ethanol:
 - combustion
 - nucleophilic substitution to give halogenoalkanes
 - reaction with sodium
 - oxidation to carbonyl compounds and carboxylic acids
 - dehydration to alkenes
- suggest characteristic distinguishing reactions for the different classes of alcohols (primary, secondary and tertiary alcohols), e.g. mild oxidation
- deduce the presence of a $\text{CH}_3\text{CH}(\text{OH})-$ group in an alcohol from its reaction with alkaline aqueous iodine to form tri-iodomethane
- recall the chemistry of phenol, as exemplified by the following reactions:
 - with bases
 - with sodium
 - nitration of, and bromination of, the benzene ring
- explain the relative acidities of water, phenol and ethanol in aqueous medium (interpret as Brønsted-Lowry acids)

11.6 Carbonyl Compounds

Content

- Aldehydes (exemplified by ethanal)
 - oxidation to carboxylic acid
 - nucleophilic addition with hydrogen cyanide
 - characteristic tests for aldehydes
- Ketones (exemplified by propanone and phenylethanone)
 - nucleophilic addition with hydrogen cyanide
 - characteristic tests for ketones

Learning Outcomes

Candidates should be able to:

- describe the formation of aldehydes and ketones from, and their reduction to, primary and secondary alcohols respectively
- describe the mechanism of the nucleophilic addition reactions of hydrogen cyanide with aldehydes and ketones
- explain the differences in reactivity between carbonyl compounds and alkenes towards nucleophilic reagents, such as lithium aluminium hydride and hydrogen cyanide
- describe the use of 2,4-dinitrophenylhydrazine (2,4-DNPH) to detect the presence of carbonyl compounds
- deduce the nature (aldehyde or ketone) of an unknown carbonyl compound from the results of simple tests (i.e. Fehling's and Tollens' reagents; ease of oxidation)
- deduce the presence of a $\text{CH}_3\text{CO}-$ group in a carbonyl compound from its reaction with alkaline aqueous iodine to form tri-iodomethane

11.7 Carboxylic Acids and Derivatives

Content

- Carboxylic acids (exemplified by ethanoic acid and benzoic acid)
 - formation from primary alcohols and nitriles
 - salt, ester and acyl chloride formation
- Acyl chlorides (exemplified by ethanoyl chloride)
 - ease of hydrolysis compared with alkyl and aryl chlorides
 - reaction with alcohols, phenols and primary amines
- Esters (exemplified by ethyl ethanoate and phenyl benzoate)
 - formation from carboxylic acids and from acyl chlorides
 - hydrolysis (under acidic and under basic conditions)

Learning Outcomes

Candidates should be able to:

- describe the formation of carboxylic acids from alcohols, aldehydes and nitriles
- describe the reactions of carboxylic acids in the formation of:
 - salts
 - esters on condensation with alcohols, using ethyl ethanoate as an example
 - acyl chlorides, using ethanoyl chloride as an example
 - primary alcohols, via reduction with lithium aluminium hydride, using ethanol as an example
- explain the acidity of carboxylic acids and of chlorine-substituted ethanoic acids in terms of their structures
- describe the hydrolysis of acyl chlorides
- describe the condensation reactions of acyl chlorides with alcohols, phenols and primary amines
- explain the relative ease of hydrolysis of acyl chlorides, alkyl chlorides and aryl chlorides

- (g) describe the formation of esters from the condensation reaction of acyl chlorides, using phenyl benzoate as an example
- (h) describe the acid and base hydrolysis of esters

11.8 Nitrogen Compounds

Content

- Amines (exemplified by ethylamine and phenylamine)
 - (i) their formation
 - (ii) salt formation
 - (iii) other reactions of phenylamine
- Amides (exemplified by ethanamide)
 - (i) formation from acyl chlorides
 - (ii) neutrality of amides
 - (iii) hydrolysis (under acidic and under basic conditions)
- Amino acids (exemplified by aminoethanoic acid)
 - (i) their acid and base properties
 - (ii) zwitterion formation
- Proteins
 - (i) formation of proteins
 - (ii) hydrolysis of proteins

Learning Outcomes

Candidates should be able to:

- (a) describe the formation of amines as exemplified by ethylamine (through amide and nitrile reduction; see also Section 11.4) and by phenylamine (through the reduction of nitrobenzene)
- (b) describe the reaction of amines in the formation of salts
- (c) describe and explain the basicity of primary, secondary and tertiary amines in the gaseous phase (interpret as Lewis bases)
- (d) explain the relative basicities of ammonia, ethylamine and phenylamine in aqueous medium, in terms of their structures
- (e) describe the reaction of phenylamine with aqueous bromine
- (f) describe the formation of amides from the condensation reaction between RNH_2 and $\text{R}'\text{COCl}$
- (g) explain why an amide is neutral in terms of delocalisation of the lone pair of electrons on nitrogen
- (h) describe the chemistry of amides, exemplified by the following reactions:
 - (i) hydrolysis on treatment with aqueous alkali or acid
 - (ii) reduction to amines with lithium aluminium hydride
- (i) describe the acid/base properties of amino acids and the formation of zwitterions
[knowledge of isoelectric points is **not** required]
- (j) describe the formation of peptide (amide) bonds between α -amino acids, and hence explain protein formation
- (k) describe the hydrolysis of proteins

12. Electrochemistry

Content

- Redox processes: electron transfer and changes in oxidation number (oxidation state)
- Electrode potentials
 - (i) standard electrode (redox) potentials, E^\ominus ; the redox series
 - (ii) standard cell potentials, E_{cell}^\ominus , and their uses
 - (iii) batteries and fuel cells
- Electrolysis
 - (i) factors affecting the amount of substance liberated during electrolysis
 - (ii) the Faraday constant; the Avogadro constant; their relationship
 - (iii) industrial uses of electrolysis

Learning Outcomes

Candidates should be able to:

- describe and explain redox processes in terms of electron transfer and/or of changes in oxidation number (oxidation state)
- define the terms:
 - standard electrode (redox) potential*
 - standard cell potential*
- describe the standard hydrogen electrode
- describe methods used to measure the standard electrode potentials of:
 - metals or non-metals in contact with their ions in aqueous solution
 - ions of the same element in different oxidation states
- calculate a standard cell potential by combining two standard electrode potential
- use standard cell potentials to:
 - explain/deduce the direction of electron flow from a simple cell
 - predict the spontaneity of a reaction
- understand the limitations in the use of standard cell potentials to predict the spontaneity of a reaction
- construct redox equations using the relevant half-equations (see also Section 13)
- state and apply the relationship $\Delta G^\ominus = -nFE^\ominus$ to electrochemical cells, including the calculation of E^\ominus for combined half reactions
- predict qualitatively how the value of an electrode potential varies with the concentration of the aqueous ion
- state the possible advantages of developing other types of cell, e.g. the H_2/O_2 fuel cell and improved batteries (as in electric vehicles) in terms of smaller size, lower mass and higher voltage
- state the relationship, $F = Le$, between the Faraday constant, the Avogadro constant and the charge on the electron
- predict the identity of the substance liberated during electrolysis from the state of electrolyte (molten or aqueous), position in the redox series (electrode potential) and concentration

- (n) calculate:
- the quantity of charge passed during electrolysis
 - the mass and/or volume of substance liberated during electrolysis
- (o) explain, in terms of the electrode reactions, the industrial processes of:
- the anodising of aluminium
 - the electrolytic purification of copper
- [technical details are **not** required]

13. An Introduction to the Chemistry of Transition Elements

Content

- General physical and characteristic chemical properties of the first set of transition elements, titanium to copper
- Colour of complexes

Learning Outcomes

Candidates should be able to:

- explain what is meant by a transition element, in terms of d block elements forming one or more stable ions with partially filled d subshells
- state the electronic configuration of a first row transition element and its ions
- explain why atomic radii and first ionisation energies of the transition elements are relatively invariant
- contrast, qualitatively, the melting point and density of the transition elements with those of calcium as a typical s block element
- describe the tendency of transition elements to have variable oxidation states
- predict from a given electronic configuration, the likely oxidation states of a transition element
- describe and explain the use of $\text{Fe}^{3+}/\text{Fe}^{2+}$, $\text{MnO}_4^-/\text{Mn}^{2+}$ and $\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}$ as examples of redox systems (see also Section 12)
- predict, using E^\ominus values, the likelihood of redox reactions
- define the terms *ligand* and *complex* as exemplified by the complexes of copper(II) ions with water, ammonia and chloride ions as ligands (including the transition metal complexes found in the Qualitative Analysis Notes)
- explain qualitatively that ligand exchange may occur, as exemplified by the formation of the complexes in (i), including the colour changes involved, and CO/O_2 exchange in haemoglobin
- describe, using the shape and orientation of the d orbitals, the splitting of degenerate d orbitals into two energy levels in octahedral complexes
- explain, in terms of d orbital splitting and d-d transition, why transition element complexes are usually coloured
[knowledge of the relative order of ligand field strength is **not** required]
- explain how some transition elements and/or their compounds can act as catalysts (see also 8(j))