

AM SYLLABUS (2008-2010)

CHEMISTRY

AM 06

SYLLABUS

Chemistry AM06 Syllabus	(Available in September)
Paper I (3 hours) + Paper II (3 hours) + Paper III (3 hours)	

Chemistry is a core science and its importance to human well-being and continued development cannot be over-emphasised. The study of chemistry at this level requires a sound appreciation of basic principles without which it is difficult to understand the many facts of the science. Indeed, the syllabus stresses the principles, knowledge of which being considered to be of even greater importance than memorization of chemical facts. The syllabus assumes a knowledge of chemistry at a level equivalent to that defined by the Secondary Education Certificate syllabus.

The Examination

The examination consists of two written papers and a practical paper. Each of the three papers is of three hours' duration.

Paper I consists of six to ten compulsory structured questions and carries 40% of the total score. Students will write their answers on the examination paper in the spaces provided.

Paper II consists of two sections each containing four extended response questions. Five questions are to be answered from this paper, two questions from each section and any one other. This paper carries 40% of the total score.

Paper III is a practical examination. This is an open book examination and candidates may use any printed material which assists them in their work. This paper carries 20% of the total score.

Candidates may use an electronic calculator in all parts of the examination.

Familiarity with Periodic Table

Candidates are expected to be familiar with the structure of the Periodic Table and the group affiliation of each element with atomic number from 1 to 30. However, a copy of the Periodic Table will be provided for Paper II of the examination although not for Paper I. Relative atomic masses will always be given where necessary.

Mathematical skills

In order to understand certain concepts required by the syllabus, candidates will need to be able to:

- (i) Recognise and use expressions in decimal and standard form; use ratios, fractions and percentages and find arithmetic means; make estimates of results; use an appropriate number of significant figures; use calculators to find and use x^y , \sqrt{x} , $1/x$, $\log_{10} x$.

- (ii) Change the subject of an equation; substitute numerical values into algebraic equations using units for physical quantities; use logarithms in relation to quantities which range over several orders of magnitude.
- (iii) Appreciate angles and shapes in regular two and three dimensional structures and to represent three dimensional forms in two dimensions; understand the symmetry of two dimensional and three dimensional shapes.
- (iv) Plot two variables from given data; understand that $y = mx + c$ represents a linear relationship and be able to determine the slope and intercept of a line; draw and use the slope of a tangent to a curve as a measure of rate of change.

Language skills

Candidates are reminded of the importance of the use of good English in answering questions in the examination. In essay-type answers and also where calculations are involved, orderly and detailed presentation will be rewarded.

Syllabus

1. Atomic Structure

- 1.1 The fundamental particles: protons, electrons and neutrons, their charges and relative masses.
- 1.2 The nucleus of the atom. Proton (or atomic) number and nucleon (or mass) number. Isotopes and relative atomic masses. The carbon-12 scale. Use of isotopes as tracers in mechanistic studies exemplified by the use of O^{18} in esterification.
- 1.3 Radioactivity: alpha and beta particles and gamma rays. Half-life. Calculations requiring use of the equation for radioactive decay will not be set. Nuclear equations. Uses and application of radioactivity including a more detailed account of carbon dating.
- 1.4 Introductory treatment of quantised energy levels in atoms: evidence from atomic spectra; s, p and d orbitals; prediction of electronic configuration of isolated atoms of elements H to Kr using 1s, 2s, 2p notation and electrons-in-boxes notation using the 'building -up' (aufbau) principle. Classification of elements into periods, groups and 'blocks', including the transition elements. Shapes of s and p orbitals. Hybridised orbitals. Questions on interpretation of atomic spectra will not be set.
- 1.5 Definition of ionisation energies and electron affinity and relation to electronic configuration. Variation of ionisation energies across a period and down a group.
- 1.6 An elementary treatment of mass spectrometry: use of mass spectrometer to determine the relative atomic mass (see also §2.9).

2. *Electronic Theory and Chemical Bonding.*

- 2.1 The ionic (electrovalent), covalent and co-ordinate (or dative) bond. Electronegativity. Intermediate bonding: ion polarisation and bond polarisation. Electrical dipoles in molecules (qualitative treatment only). Polar covalent bonds which may, or may not, give rise to molecules with a permanent dipole.
- 2.2 Nature of forces in bonding. Comparison between ionic and covalent bonding. Multiple bonding.
- 2.3 Delocalisation of electrons.
- 2.4 Metallic bonding (electron sea model) and use of model to account for malleability, ductility and thermal and electrical conductivity.
- 2.5 Lattice structures of NaCl and CsCl as typical ionic solids. Coordination numbers.
- 2.6 Simple shapes of molecules and ions explained on the basis of the valence shell electron pair repulsion (VSEPR) theory.
- 2.7 Molecular crystals, e.g. iodine, and macromolecular structures, e.g. carbon (diamond), carbon (graphite) and silicon(IV) oxide; relationship between structure and physical properties.
- 2.8 Intermolecular forces: hydrogen bonding, dipole-dipole interactions and van der Waals forces (also known as induced dipole-induced dipole forces). Effects of these forces on specific properties of molecular compounds: e.g. boiling points of simple hydrides in Groups 4, 5, 6 and 7; formation of carboxylic acid dimers.
- 2.9 Mass spectrometer as a method of determining molecular structure: interpretation of simple mass spectra involving only singly charged ions; the molecular (parent) ion and fragment ions.

3. *States and Quantity of Matter*

- 3.1 The ideal gas law and its use in the determination of relative molecular mass for gases and volatile liquids: use of gas syringe.
- 3.2 Dalton's Law of Partial Pressures; diffusion and effusion (Graham's Law will not be examined).
- 3.3 Kinetic theory of gases (treated qualitatively). Distribution of molecular speeds.
- 3.4 Real gases and deviations from ideality. Familiarity with van der Waals equation will be assumed although no numerical questions will be set.

- 3.5 Vapour pressure and saturation vapour pressure; vaporisation and fusion in terms of the kinetic molecular model.
- 3.6 Moles of substance. Concentration in terms of moles dm^{-3} or mol L^{-1} ; mass concentration in terms of g dm^{-3} or g L^{-1} . Volume measured in dm^3 or L and cm^3 or mL. Avogadro constant. Empirical and molecular formulae; chemical equations, both full and ionic and the use of these equations in calculations on reacting substances in terms of amounts and concentrations measured in moles and mol dm^{-3} respectively. Concept of limiting reagent. Percentage yield. Acid-base and redox reactions in titrimetry. Back titration.

4. *Energetics*

- 4.1 Changes of energy accompanying phase changes and chemical changes. Energy level diagrams. Joule (J) as the unit of energy.
- 4.2 Standard enthalpy change of reaction, formation, atomisation, ionization, combustion, solution (or dissolution), neutralisation and solvation including hydration. Electron affinity including discussion of first and second electron affinity for dinegative ions. Bond enthalpy terms: distinction from bond dissociation enthalpies; lattice enthalpies; enthalpies of solution and relation to lattice enthalpy and enthalpies of hydration.
- 4.3 Hess's law and its use in simple calculations. Born-Haber cycle.
- 4.4 Comparison between theoretical and experimental values of enthalpy change and interpretation in terms of structure and bonding.
- 4.5 Calorimetry: determination of enthalpy change (heat) of neutralization, solution and reaction. Calorific value of fuels and food. Thermometric titrations.
- 4.6 Concept of system and its surroundings. Entropy, Gibbs free energy and spontaneity of chemical change. Familiarity and use of the relationship between free energy, enthalpy and entropy, namely, $\Delta G = \Delta H - T\Delta S$. Kinetic versus thermodynamic stability.

5. *Phase Equilibria*

- 5.1 Phase changes for one-component systems: H_2O , CO_2 . Pressure-temperature phase diagrams and boiling, melting and triple points. Critical temperature and supercritical fluids; use of supercritical fluids as solvents.
- 5.2 Two-component systems: mixtures of two miscible liquids and Raoult's Law. Pressure-composition and temperature-composition diagrams. Deviations from Raoult's Law. Azeotropic mixtures. Fractional distillation of ideal and non-ideal mixtures.
- 5.3 Immiscible liquids and steam distillation: numerical calculations involving vapour pressure may be set.

- 5.4 Osmosis and osmotic pressure: use of osmotic pressure, Π , for the determination of relative molecular mass from relation $\Pi V = nRT$ where n is the amount of particles of solute. Questions on the determination of partial degree of dissociation/association will not be set. Reverse osmosis.

6. *Chemical Equilibrium*

- 6.1 Concept of dynamic equilibrium. Characteristics of an equilibrium mixture. Equilibrium constant in terms of concentrations and of pressures: K_c and K_p including units as appropriate. Homogeneous and heterogeneous equilibria and their associated equilibrium constants. Relation of degree of dissociation to K_c and K_p (problems requiring solving quadratic equations using the formula will not be set).
- 6.2 Experimental methods of investigating chemical equilibrium, e.g. an esterification reaction.
- 6.3 Le Chatelier's Principle. The effect of concentration, temperature, and (where relevant) pressure on: (a) wholly gaseous equilibria, (b) solid-gas equilibria, (c) equilibria in the liquid phase (e.g. esterification) or in solution. Effects of presence of catalysts on equilibrium.
- 6.4 Distribution of a non-volatile solute between two immiscible solvents: partition constant; solvent extraction. Discussion of association and dissociation effects on partition equilibrium are not required.
- 6.5 Industrial processes where equilibria are involved: Haber process for ammonia; Contact process for sulfuric acid including the understanding of why compromise conditions are employed.

7. *Ionic Equilibria*

- 7.1 Acids and bases: Arrhenius, Bronsted-Lowry and Lewis definitions; conjugate pairs; proticity or 'basicity' of acids and 'acidity' of bases (e.g. sulfuric acid as a 'diprotic' or 'dibasic' acid and carbonate as a 'diacid' base); strong and weak acids and bases; amphoteric compounds; dissociation constants K_a and K_b and the pK convention. Ionic product of water; pH; hydrolysis of salts; indicators, typified by phenolphthalein and methyl orange (students are expected to recall the pH range of application of phenolphthalein and methyl orange); buffer solutions; pH curves restricted to monobasic acid - monoacid base reactions). Calculations involving pH, including those involving buffers, may be set.
- 7.2 Role of solvent in equilibria involving ionisation of molecular solutes. Degree of ionisation. Conductivity of solutions of strong and weak electrolytes (treated qualitatively).

- 7.3 Heterogeneous ionic equilibria: solubility product and relation to molar solubility; common ion effect. Numerical calculations on solubility product will not be set.

8. *Redox Equilibria*

- 8.1 Redox reactions: balancing of equations using half equations. Disproportionation reaction. Oxidation number.
- 8.2 Electrodes, galvanic cells and standard redox potentials. Standard hydrogen electrode as a reference electrode. Cell diagrams.
- 8.3 Application of redox potentials to the prediction of redox change. Electrochemical series. Simple treatment of Nernst Equation and non-standard electrode potentials (where needed equation will be given). Use of electrode potentials in deriving solubility product will not be examined. Corrosion as an electrochemical process: sacrificial protection.

9. *Reaction Kinetics*

- 9.1 The experimental investigation of reaction rate for simple reactions involving use of gas syringe, colorimeter, conductance measurements, etc.
- 9.2 Order and rate coefficients and their measurements, including the initial rate method; rate equation (fractional and pseudo orders will not be tested). Graphical presentation of kinetic results. Integrated rate equations are not required. Half-life for first order reactions.
- 9.3 Effects of pressure, concentration, surface area, temperature and catalysts on reaction rate. Activation energy of reactions. Maxwell-Boltzmann distribution of energies and the collision theory of reactions.
- 9.4 Concept of rate-determining step in a multistep reaction and its relevance to the mechanism of the reaction. Third order reactions and notion of the improbability of three-particle collisions.
- 9.5 Photochemical reactions and free radical mechanisms.
- 9.6 Homogeneous and heterogeneous catalysis; examples of the use of catalysts in industrial processes; catalytic converters; autocatalysis.

10. *Principles of Chemical Periodicity*

Periodic classification in terms of electronic structure. Periodic relationships amongst the elements Li to Ar. Variation in properties as illustrated by: trends in melting and boiling points, electrical conductivity and ionisation energy of elements; reaction of elements with oxygen, chlorine and water; formulae and acid-base character of oxides/hydroxides of metals and oxides of non-metals (limited to Na_2O , MgO , Al_2O_3 , NaOH , $\text{Mg}(\text{OH})_2$, $\text{Al}(\text{OH})_3$, SiO_2 , P_4O_{10} , SO_2 , SO_3 , Cl_2O); formulae of chlorides and

reactions with water; formulae of simple hydrides (excluding boron hydrides).

11. Descriptive Inorganic Chemistry

11.1 Comparative study of the s-block elements (i) lithium, sodium and potassium and (ii) beryllium, magnesium, calcium, strontium and barium.

Fixed oxidation states and electronic configuration. Physical properties of elements; flame tests; reaction with water and with oxygen; diagonal relationships. Oxides, peroxides, hydroxides, carbonates, nitrates(V), nitrates(III), sulfates(VI) and sulfates(IV): trends in thermal stability of nitrates(V) and carbonates and solubility of hydroxides and sulfates(VI). Amphotericity of BeO, covalency of BeCl₂.

11.2 Chemistry of carbon, silicon, tin and lead.

Allotropes of carbon, including fullerenes. Variation in ionisation energies of the elements and structure of the elements and their effect on physical properties. Main oxidation states and inert pair effect. Hydrolytic behaviour of chlorides. Principal oxides: preparation (excluding practical details) to include formation of lower oxides from thermal decomposition of ethanedioates; acid/base character of oxides. Hydrides. Halides of tin and lead and chloro complexes. Redox character of the compounds of Sn(II) and Pb(IV).

11.3 Chemistry of fluorine, chlorine, bromine and iodine.

Laboratory preparation of chlorine, bromine and iodine. Physical properties: trends in boiling points, ionisation energies, electron affinities and bond dissociation enthalpies for the series X₂ and HX (X = F, Cl, Br, I). Reactions between ionic halides X⁻ (X = Cl, Br, I) and silver ions, phosphoric(V) and sulfuric(VI) acids. Acidity of HX and anomalous value for HF. Variable oxidation states; relative oxidising strength of halogens X₂ (X = Cl, Br, I) and disproportionation reactions of chlorine and its oxoanions. Iodine in titrimetry; triiodide ion. Hydrated and anhydrous aluminium chloride and chlorides of iron to illustrate the difficulty of preparation of hydrolysable salts.

11.4 Transition elements

A general overview of the transition metals emphasising features they have in common; electronic configurations, metallic character, variable oxidation states, catalytic roles, coloured compounds and formation of complex ions. The formation of complex ions explained in terms of the electrostatic model or else in terms of dative covalency between ligand and central metal ion; shapes of di-, tetra-, and hexa-coordinated systems; elementary treatment of isomerism in complex ions (compare §12.5); stability of complex ions and ligand exchange; nomenclature of complex ions: refer to publication *Chemical Nomenclature for Use in Matriculation Examinations*.

11.5 More detailed chemistry of the following d-block elements: chromium, manganese, iron and copper.

- (i) *Chromium*: reaction of metal with acids, including passivation with HNO_3 , with alkalis to form chromate(III); chromium(II) state and its strong reducing character, including its formation from Cr(VI) by reduction with, e.g. Zn/H^+ ; chromium(III) compounds: oxide, hydroxide and chromates(III), CrCl_3 and its hexahydrates as examples of isomerism involving complex ions: chrome alums; chromium(VI): CrO_3 ; chromate(VI)-dichromate(VI) equilibrium, chromates(VI) (including ammonium chromate(VI)), toxicity of Cr(VI) compounds.
- (ii) *Manganese*: reactivity of metal exhibited by its reaction with dilute acids and water to form Mn^{2+} and hydrogen; manganese(II); oxide and hydroxide (instability with respect to oxidation to Mn(III) and Mn(IV)); oxidation of manganese(II) to manganese(VII); manganese(IV) oxide and its strong oxidising power; conversion of MnO_2 to manganate(VI) and disproportionation of MnO_4^{2-} ; KMnO_4 ; its preparation and reactions as an oxidising agent.
- (iii) *Iron*: reaction of metal with oxidising and non-oxidising acids and water; reaction with non-metals including C to form steels; rusting of iron and protection, including sacrificial anodizing with zinc; iron(II) compounds: oxide, hydroxide, salts, including ammonium iron(II) sulfate-6-water; iron(III) compounds; oxide, hydroxide, anhydrous FeCl_3 , acidity of hexaaquairon(III) chloride; iron(III) alums; reasons for the non-existence of iron(III) iodide and iron(III) carbonate; important complex ions of iron: $[\text{Fe}(\text{CN})_6]^{4-}$, $[\text{Fe}(\text{CN})_6]^{3-}$, $[\text{Fe}(\text{SCN})(\text{H}_2\text{O})_5]^{2+}$ (use in qualitative analysis), $[\text{Fe}(\text{NO})(\text{H}_2\text{O})_5]^{2+}$ (treated as an Fe(II)-NO complex) and the brown ring test for nitrate(V).
- (iv) *Copper*: reaction with oxidising acids; copper(I) state exemplified by Cu_2O (relation to Fehling's Test for aldehydes), CuCl and CuI , instability of $\text{Cu}^+(\text{aq})$ with respect to disproportionation; copper(II) compounds: oxide, hydroxide, sulfate(VI), nitrate(V) and basic carbonates; complexes of copper(I) and copper(II) with ammonia and chloride ions.

11.6 A more detailed study of selected non-metals.

- (i) *Hydrogen*: manufacture, laboratory preparation and properties as reducing agent; ionic, covalent and interstitial hydrides; complex hydrides (typified by LiAlH_4); water and its physical properties as determined by hydrogen bonding; deuterium oxide and deuteroderivatives (eg: DCl , ND_3 , C_2D_2).
- (ii) *Nitrogen*: unreactivity of nitrogen and relation to strength of the nitrogen-nitrogen triple bond; ammonia and its reactions as base, ligand and reducing agent; ammonium salts; oxides of nitrogen N_2O , NO , NO_2 , N_2O_4 , N_2O_5 ; preparation, electronic structures and molecular

shapes, relation to oxo acids of nitrogen; nitric(III) (nitrous) acid and nitrates(III) (nitrites) of s-block elements, decomposition of aqueous ammonium nitrate(III); nitric(V) (nitric) acid: industrial preparation, reactions as acid, oxidising and nitrating agent; nitrates(V) and their reactions: thermal decomposition, reduction, ammonium nitrate(V).

- (iii) *Oxygen*: allotropy; effects of stratospheric and tropospheric trioxygen (ozone); laboratory preparation of oxygen from oxides, peroxides and electrolysis; reaction of oxygen with metals and non-metals; acidic, basic, amphoteric and neutral oxides; normal and mixed oxides (typified by Fe_3O_4 and Pb_3O_4), hydrogen peroxide and peroxides.
- (iv) *Sulfur*: allotropy; hydrogen sulfide as an acid, a reducing agent and as a precipitant for insoluble sulfides; toxicity of hydrogen sulfide; sulfur dioxide and sulfuric(IV) (sulfurous) acid, sulfates(IV) (sulfites): reducing character; sulfur trioxide and sulfuric(VI) (sulfuric) acid: reactions of the concentrated acid as an involatile proton donor, oxidising agent and strong dehydrating agent, dilute sulfuric(VI) acid and sulfates(VI); thiosulfate: formation from sulfate(IV) and reaction with acids, iodine and chlorine. Oxides of sulfur as environmental pollutants.

12. Fundamental Principles of Organic Chemistry

- 12.1 Functional groups and homologous series. Molecular, empirical and structural formulae including graphical representations. Nomenclature of organic compounds. Systematic names: candidates are referred to the publication *Chemical Nomenclature for Use in Matriculation Examinations*.
- 12.2 Purification of compounds: solvent extraction, recrystallisation; drying, simple, fractional and steam distillation, sublimation; column, paper and thin layer chromatography. Preparation of derivatives and their use for characterisation. The determination of melting points as a test for purity: mixed melting point technique.
- 12.3 The determination of empirical, molecular and structural formulae from analytical information (experimental techniques not required). Structure deduced from chemistry (sodium fusion test excluded), physical properties and from information derived from instrumental techniques, namely, mass spectrometry and infra red spectrophotometry. Interpretation of infra red spectra for the identification of simple functional groups using tables of frequencies/wave numbers, details of instrumentation are not required.
- 12.4 Catenation in carbon compounds. Spatial distribution of the bonds in simple carbon compounds. Tetrahedral distribution of bonds in alkanes, planarity of $>\text{C}=\text{C}<$ in alkenes and of $>\text{C}=\text{O}$ in carbonyl compounds and linearity of $-\text{C}\equiv\text{C}-$ in alkynes. Hybridisation of orbitals and organic structure: sp^3 , sp^2 and sp orbitals.

- 12.5 Structural, geometrical and optical isomerism (restricted to enantiomerism); racemic mixtures (resolution of which will not be tested). Tautomerism. The influence of size, shape and polarity on physical properties of organic molecules.
- 12.6 Delocalisation of electrons in organic molecules and canonical formulae; sigma (σ) and pi (π) bonds. Delocalisation energy of conjugated systems.

13 Chemistry of Aliphatic and Aromatic Compounds

- 13.1 *Hydrocarbons and petroleum*: alkanes, alkenes and aromatic hydrocarbons from petroleum by cracking and reforming; use of hydrocarbons as fuels.
- 13.3 *Alkanes*: chlorination; cycloalkanes typified by cyclohexane, boat and chair conformation.
- 13.4 *Alkenes*: preparation by elimination reactions (from alcohols and haloalkanes); addition reactions (with halogen, hydrogen, halogen halides, sulfuric(VI) acid), Markownikoff's rule; ozonolysis and use of reaction for location of double bond, reaction with alkaline manganate(VII); polymerization reactions: high and low density polymers and Ziegler-Natta catalysts in the formation of stereoregular poly(propene), biodegradable plastics.
- 13.5 *Alkynes*: preparation of ethyne from CaC_2 and propyne by dehydrohalogenation reactions; addition reactions including hydration to produce carbonyl compounds.
- 13.6 *Arenes*: benzene and other alkylbenzenes. Substitution reactions: nitration, sulfonation, alkylation, acylation, halogenation (nuclear versus side chain for methylbenzene); influence of substituents on further substitution in the benzene ring: activating and deactivating groups exemplified by $-\text{CH}_3$, $-\text{OH}$, $-\text{NH}_2$, $-\text{NO}_2$, $-\text{CHO}$ and $-\text{COOH}$; addition reactions of hydrogen and chlorine; side chain oxidation of alkylbenzenes using alkaline manganate(VII) to produce aromatic carboxylic acid salts.
- 13.7 *Alcohols, ethers, phenols*: primary, secondary and tertiary alcohols and their differentiation on the basis of oxidation, Lucas test and haloform reaction; aromatic alcohols typified by phenylmethanol; preparative techniques of alcohols from halogenoalkanes, alkenes and by reduction of aldehydes, ketones and carboxylic acids; formation of alcohols by addition of Grignard reagent to aldehydes and ketones; formation of alkoxides, halogenoalkanes, esters; alkenes and ethers from alcohols; ethane-1,2-diol and propane-1,2,3-triol to typify polyhydric alcohols. Ethers: preparation from alkoxides and halogenoalkanes and their cleavage with HI. Phenols: preparation from sulfonic acids and from diazonium salts, acidity of phenols, ether formation, substitution reactions in the ring including tribromination and reaction with diazonium ions; reaction with FeCl_3 as a colour test for phenols; polymerization reaction of phenol with methanal.

- 13.7 *Halogenoalkanes and halogenoarenes*: preparation of halogenoalkanes from alcohols or alkenes; conversion of halogenoalkanes into alcohols, ethers, amines, nitriles, esters, alkanes, Grignard reagents and alkylarenes; reactivity of halogenoalkanes related to structure (primary, secondary, tertiary halogenoalkanes) and nature of halogen (Cl, Br, I); organohalogen compounds as solvents and biocides and their environmental impact; halogenoarenes: preparation by direct halogenation (where appropriate) and via diazonium compounds; unreactivity of halogen atom with respect to substitution reactions explained in terms of delocalization of electrons; (chloromethyl)benzene and comparison with nuclear halogenoarenes.
- 13.8 *Aldehydes and ketones*: preparation from corresponding alcohols, from CrCl_2/O_2 -oxidation of methylarenes (for aromatic aldehydes) and by Friedel-Crafts reaction (for aromatic ketones); difference in reactivity related to structure; reactions: reduction by reaction with hydrogen and hydridometallates; oxidation, a method for distinguishing between aldehydes and ketones (Fehling, Tollen); addition reactions (ROH , HCN , NaHSO_3), polymerisation of aldehydes; condensation reactions with hydrazine, phenylhydrazine and 2,4-dinitrophenylhydrazine and hydroxylamine, formation of derivatives for characterisation purposes; condensation polymers from methanal, halogenation and the haloform reaction; aldol reaction; Cannizzaro reaction.
- 13.9 *Carboxylic acids and derivatives*: carboxylic acids: weak protic acidity and effect on acidity of α -substitution with halogen atoms; salt formation and alkalinity of carboxylates; soaps from long chain (fatty) acids and comparison with detergents; decarboxylation reaction; reduction (LiAlH_4) to primary alcohols; conversion of carboxylic acids into esters, acid chlorides, acid anhydrides, acid amides and chloroacids; formation of esters from phenols; hydrolysis and interconversion reactions of acid derivatives; Hofmann degradation of amides; formation of nitriles from amides, halogenoalkanes and diazonium salts (isonitriles not required); conversion of nitriles to amines and hydrolysis to carboxylic acids.
- 13.10 *Amines and diazonium compounds*: primary, secondary and tertiary amines; quaternary ammonium compounds; basicity of amines, salt formation and acidity of alkylammonium salts: Hofmann preparation of amines; reduction of simple and N-substituted amides to form corresponding amines, hydrolysis of N-substituted amides, aromatic amines by reduction of nitroarenes; alkylation and acylation reactions of amines; reaction of primary amines with nitric(III) acid; halogenation of phenylamine; diazonium salts: formation; substitution reactions to form phenols, halogenoarenes, arenes and nitriles and coupling reactions with phenols or arylamines to form azo dyes.
- 13.11 *Difunctional molecules*: amino acids: zwitterion formation; relation to peptides and proteins, synthetic polyamides (eg: Nylon); intramolecular dehydration of dicarboxylic acids exemplified by butenedioic and benzene-1,2-dicarboxylic acids; polyesters from hydroxycarboxylic acids.

13.12 *Mechanistic aspects*: reaction mechanisms should be emphasised and used to rationalise the facts of organic chemistry. It is expected that students be familiar with the following mechanistic concepts: nucleophile, electrophile, free radical; inductive and mesomeric effects; bimolecular and unimolecular nucleophilic substitution reactions of halogenoalkanes, electrophilic substitution of arenes; mechanism of esterification reactions; nucleophilic addition to C=O group; electrophilic addition to >C=C< bond; homolytic reactions: the halogenation of alkanes and methylbenzene; polymerisation of alkenes.

13.13 Students are expected to be familiar with detailed laboratory preparation and isolation of pure samples of the following organic substances: (a) a bromoalkane from the corresponding alcohol; (b) nitrobenzene and 1,3-dinitrobenzene from benzene; (c) ethanal from ethanol; (d) an ester from the reaction of a carboxylic acid and an alcohol; (e) phenylamine from nitrobenzene.

14. PRACTICAL EXAMINATION

The practical examination will seek to test the ability of candidates to:

- (a) manipulate chemicals and simple apparatus in quantitative and qualitative exercises;
- (b) observe and record results of experimental work;
- (c) interpret these observations and deduce correct inferences and conclusions based both on qualitative and quantitative data.

The examination will primarily attempt to test practical skills. Quantitative exercises including the measurement of mass, volume, temperature and time may be set. Volumetric analysis involving acids and alkalis, redox titrations including iodimetry and silver nitrate may also be examined.

Qualitative exercises involving observations of reactions and requiring deductions on the chemical nature of the substances will also be set. These will involve both organic and inorganic materials. Inorganic materials set as unknowns will be chosen from the following types:

Anions: oxide, hydroxide, carbonate, hydrogencarbonate, nitrate(V), nitrate(III), carboxylate, sulfate(IV), sulfate(VI), thiosulfate, chloride, bromide, iodide, chromate(VI), dichromate(VI), phosphate, manganate(VII);

Cations: ammonium, potassium, sodium, silver, calcium, magnesium, strontium, barium, aluminium, lead(II), zinc, chromium(III), manganese(II), manganese(IV), manganate(VII), iron(II), iron(III), nickel(II), copper(I), copper(II).

Candidates are instructed to use labcoats and to wear eye protection (safety goggles) during the practical examination and to avoid wearing clothes or articles that increase the risk of accident in the laboratory.

Recommended Texts

Any of the following are suitable textbooks for the syllabus; the list is not comprehensive:

1. Lister, T and Renshaw, J, Understanding chemistry for Advanced Level, 4th Edition, Nelson Thornes Ltd., 1999.
2. Ramsden, EN, A-Level Chemistry, 4th Edition, Nelson Thornes Ltd., 2000.
3. Andrew, J and Rispoli, P, Chemistry in Focus, 2nd Edition, Hodder & Stoughton Ltd., 1999.
4. Jones, L and Atkins, P, Chemistry: molecules, matter and change, 4th Edition, W H Freeman, 1999.