IM SYLLABUS (2013)

CHEMISTRY IM 06

SYLLABUS

Chemistry IM 06 (Available in September)
Syllabus 1 paper (3 hours)

The syllabus is intended for candidates who have followed an introductory course in chemistry. It attempts to consolidate this knowledge by extending it beyond the limits usually set at Secondary Education Certificate (SEC) level to cover more advanced concepts in chemistry which are dealt with in a broad manner. Most of the concepts studied at intermediate level will not be developed to the same depth and with the same rigour as at the Advanced Matriculation level. It is hoped that a general overview of chemistry is conveyed which covers in sufficient detail key concepts in the subject.

The Examination

The Examination consists of one three-hour paper. The paper will be divided into three sections: Section A will contain between fifteen and twenty compulsory questions of the fill-in type requiring short answers; Section B will consist of five compulsory structured questions; Section C will require candidates to choose two out of four long questions. Each section carries equal marks. The minimum mathematical requirements of the syllabus are the same as those for the SEC examination in Chemistry. Questions will be set in SI units except that pressures may be expressed in atmospheres (atm.).

The Periodic Table, complete with atomic numbers, will be provided. Relative atomic masses and the molar volume for gases will also be given when required.

Syllabus

Chemistry is an experimental science and it is essential that students spend time in a laboratory to see for themselves how chemists work. It may not be possible for students following this course to participate actively in laboratory work: to make up for this, tutors should be very generous with experimental demonstrations.

The syllabus is organised in two sections, namely Chemical Principles and Descriptive Chemistry. This organisation is not intended to serve as a teaching scheme for the subject matter.

The influence of chemicals and chemical industry on the quality of life and on the environment should be discussed throughout the course. Examples of everyday applications of chemistry and its products are numerous and they should be invoked by tutors as frequently as possible.

I. CHEMICAL PRINCIPLES

1. Atoms, molecules and mole concept.

1.1 Atomic nature of matter. Idea of size of atoms. Avogadro's constant and moles of particles. Relative atomic masses with respect to ^{12}C. Molecules and relative molecular mass as the mass of a mole of molecules. Experimental determination of Avogadro constant not examinable.

1.2 Molar volume of gases and Avogadro's Law.

2. Stoichiometry.

2.1 Experimental determination of chemical formulae from empirical data.

2.2 Balanced chemical equations to represent relative numbers of particles in chemical reactions. States of substances and use of symbols (s), (l), (g) and (aq) for solid, liquid, gas and aqueous solutions respectively.

2.3 Combining ratios of reactants and products established experimentally to include gravimetry, volumetric work and measurements involving gas volumes. Use of $pV=nRT$ and use in the determination of the relative molecular mass of a gas or vapour (calculations limited to simple examples). Use of gas syringe.

3. Atomic structure and the Periodic Table.

3.1 The nuclear atom: proton, neutron and electron. Isotopes and relation of isotopy to relative atomic masses. Mass number and atomic number and atomic symbols ${}^{A}_{Z}X$.

3.2 Radioactivity: alpha and beta particles and gamma radiation. Half life. Radioisotopes in medicine and industry.

3.3 Electrons in atoms; dot and cross diagrams for electron configurations. Electrons in energy levels; the s, p, d notation.

3.4 Ionization energies and relation to electronic configuration and valence (experimental determination of ionization energies not required). Periodic Table of elements: valence as a periodic property. Electron affinity.

4. Chemical bonding.

4.1 Ionic bonding as electrostatic attraction between ions formed by transfer of electrons. Lattice structure as exemplified by NaCl.

4.2 Covalent bonding as sharing of electrons. Single and multiple bonds. Dative covalent bonds. Simple molecules. e.g. CO_2, and macromolecules (network covalent) as exemplified by SiO_2, graphite and diamond.

4.3 Octet rule and use in predicting formulae: exceptions to octet rule to include electron deficient molecules and 'octet expansion'.

4.4 Polarity of bonds. Concept of electronegativity and intermediate bonding.

4.5 Concept of resonance: delocalisation of electrons as exemplified by benzene.

4.6 Shapes of molecules and simple ions in terms of electron pair repulsion theory e.g. BeF_2, BF_3, CH_4, SF_6, NH_3, H_2O, NH_4^+, H_3O^+.

4.7 Intermolecular forces: permanent dipole-permanent dipole; van der Waals' forces (also known as induced dipole-induced dipole forces) and hydrogen bonds.

4.8 Metallic bonding: electron sea model.

4.9 Relationship between structure and physical properties: volatility, melting and boiling point; hardness, malleability, electrical conductivity. Solubility in polar and non-polar solvents.

5. Kinetic molecular theory and states of matter.

5.1 Diffusion and Brownian motion in terms of simple kinetic theory. Interconversions between the three states of matter in terms of kinetic theory. Gaseous state: distribution of molecular kinetic energies and relation with temperature (qualitative treatment only).

5.2 Intermolecular forces and the liquid and solid states. Vapour pressure. Changes of state and associated enthalpy changes as related to structure. Hydrogen bonding and the open structure of ice: density of ice compared to that of liquid water.

5.3 Elementary treatment of mass spectrometry as a method of determining the relative atomic/molecular mass. Fragmentation of simple molecules in mass spectrometer.

6. Energetics

6.1 Changes of energy accompanying phase changes and chemical changes. Energy level diagrams.

6.2 Exothermic and endothermic processes; the Joule as a unit of energy. Thermochemical equations and the ΔH notation. Energy changes accompanying neutralisation, solution, combustion and atomisation reactions. Experimental determination of energy changes not required. Calorific value of fuels and food; hydrogen as a clean fuel.

6.3 Hess's law and calculation of enthalpy change of formation from enthalpy change of combustion.

6.4 Enthalpy of reaction explained in terms of bond breaking and bond making processes: bond enthalpy terms.

7. Rates of reactions.

7.1 Concept of reaction rate as the increase of product concentration or decrease of reactant concentration with time.
Factors influencing rate of reactions, e.g. concentration, temperature, pressure, particle size, electromagnetic radiation, catalysts.

7.2 Theory of reactions on the collision model (qualitative). Change in distribution of molecular kinetic energies to explain the relation between reaction rate and temperature.

7.3 Concept of reaction mechanism and rate determining step. Activation energy. Role of catalyst. Addition of HBr to ethene could be used as an example of a multistep reaction mechanism.

8. Reversible reactions and chemical equilibrium.

8.1 Reversible reactions and kinetic picture of dynamic equilibrium. Use of symbol $\rightleftharpoons$

8.2 The equilibrium law and equilibrium constant K_c and K_p; only homogeneous equilibria will be examined and calculations requiring application of the quadratic formula will not be set.

8.3 Le Chatelier's principle: effect of concentration, temperature and (where relevant) pressure on equilibria. Haber process as an industrial application of a reversible process.

9. Ionic theory and redox reactions.
Oxidation and reduction as the loss and gain of electrons respectively.
Disproportionation. Ionic half equations; concept of spectator ions. Oxidation number. Electrochemical series.

10. Acids and bases.

10.1 Acids, bases, alkalis and salts. Strong and weak acids and bases. Amphoteric character.

10.2 Proton transfer equilibria and Lowry-Bronsted definition. Concept of conjugate pair; K_a and K_b. Ionic product of water, K_w; pH scale. Numerical calculations involving K_a and K_b will not be set but the ability to calculate pH from molar concentrations of hydroxonium or hydroxide ions will be expected. The use of the pH meter and universal indicator in the measurement of pH. Theory of indicators will not be tested.

11. Principles of Chemical Periodicity

11.1 Periodic classification in terms of electronic structure; periodic relationships among the elements Li to Ar; variation in physical properties with atomic number. Periodic properties to include melting point, valency, atomic radius and first ionization energy.

11.2 Periodicity of the elements lithium to chlorine as shown by their formulae and reaction with water of the oxides, chlorides and simple hydrides. Trends in melting points of the oxides and chlorides.

II DESCRIPTIVE CHEMISTRY

12. Chemistry of transition metals. Colour and variable oxidation states as shown by the species Mn^{2+}, MnO_2, MnO_4^{2-} and MnO_4^{-}, Fe^{2+} and Fe^{3+}, Cu^{+} and Cu^{2+}. Catalytic properties as exemplified by use of MnO_2 in the decomposition of hydrogen peroxide; Fe in the Haber Process; Raney Ni in hydrogenation of unsaturated hydrocarbons. Coordination compounds: metal-ligand bonding described in terms of electrostatic model or dative covalent bonding; examples of complex ions, including the molecular shape to include: $[Fe(H_2O)_6]^{2+}$, $[CuCl_4]^{2-}$, $[Cu(NH_3)_4]^{2+}$, $[CuCl_2]^{2-}$, $[Fe(CN)_6]^{3-}$.

Nomenclature of coordination compounds and discussion of stereostructure are not required. Extraction of metals from ores is not required.

13. Chemistry of the halogens: chlorine, bromine, iodine. Occurrence. Manufacture and uses of chlorine. Relative oxidising power of elements. Reaction of Cl_2 and I_2 with thiosulfate. Hydrogen halides: preparation from salts, acidity of aqueous solutions. Discussion of oxoanions is not required.

14. Organic chemistry

14.1 Catenation related to strength of C-C bond. Concepts of homologous series, functional groups and structural and geometric isomerism. Nomenclature of organic compounds: candidates are referred to the publication *Chemical Nomenclature for Use in Matriculation Examinations.*

14.2 Aliphatic hydrocarbons: alkanes, alkenes, alkynes; cyclic structures. Nomenclature. Substitution reaction of alkanes with halogens to exemplify a free radical mechanism and addition reaction of alkenes with halogens and hydrogen halides as examples of an ionic mechanism; Markownikov's Rule to predict addition of HX to unsymmetrical alkenes, preparation of alkenes by dehydration of alcohols; ethyne from calcium dicarbide. Alkanes and alkenes from petroleum by cracking. Hydrogenation of alkenes.

14.3 Benzene and aromatic hydrocarbons: stable delocalised structure. Substitution reactions of benzene with chlorine, sulfuric(VI) acid, nitrating mixture (monosubstitution only). Hydrogenation of benzene. Mechanisms of reactions of benzene are not expected.

14.4 Chemistry of some common functional groups: alcohols; ethers; aldehydes and ketones; carboxylic acids; amines. The treatment should be an *outline survey* and should include the following topics (mechanisms will not be tested): (a) alcohols: oxidation, dehydration to alkenes and ethers, esterification, reaction with PCl_5, reaction with sodium; (b) aldehydes and ketones: Fehling's test, reduction, addition of HCN to form hydroxynitriles and subsequent hydrolysis to hydroxycarboxylic acid; (c) carboxylic acids: salt formation, esterification, decarboxylation and reaction with PCl_5; (d) amines: salt formation from primary, secondary and tertiary amines and comparison with ammonia (relative basic strengths of different amine types not required), conversion of nitrobenzene to phenylamine (details of preparation not required).

14.5 Polymers. Formation by addition reaction of alkenes and by condensation reactions of difunctional molecules, e.g. formation of polyesters (structure of condensation polymers shown by block diagrams).

15. Purification Techniques

15.1 Pure substances and mixtures. Solvent, solute, solution. Separation of mixtures by filtration, distillation, sublimation, crystallisation and chromatography as typified by paper chromatography. Solvent extraction: partitioning of a solute between two immiscible solvents treated as an equilibrium process.

15.2 Boiling point and melting point as criteria of purity.

16. Qualitative and quantitative analysis.

16.1 Simple tests which can be used for the identification of the following ions: sodium, potassium, magnesium, calcium, barium, ammonium, copper(II), iron(II), iron(III), manganese(II), chromium(III), zinc, lead(II), silver(I), aluminium; chloride, bromide, iodide, sulfate(VI), sulfate(IV), thiosulfate, carbonate, nitrate(V), nitrate(III), oxide, hydroxide.

16.2 Quantitative analysis by titration: standard solutions, acid-base and redox titrations and related calculations in terms of moles and concentrations in terms of mol dm^{-3}.

Recommended Texts

Any of the following textbooks may be used with profit for this syllabus although the treatment of subject matter as covered in these books will generally be more detailed than that required for the purposes of the syllabus:

1. Lister, T. and Renshaw, J., Understanding Chemistry for Advanced Level, 3rd Edition, Nelson Thornes Ltd., 2000.
2. Andrew, J. and Rispoli, P., Chemistry in Focus, 2nd Edition, Hodder and Stoughton Ltd., 1999.
3. Jones, L. and Atkins, P., Chemistry: molecules, matter and change, 4th Edition, WH Freeman, 1999.
4. Hill GC and Holman JS, Chemistry in Context, 5th Edition, Nelson Thornes, 2000.