

**IM SYLLABUS (2018)**

**CHEMISTRY**

**IM 06**

***SYLLABUS***

<b>Chemistry IM 06 Syllabus</b>	(Available in September) 1 paper (3 hours)
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The syllabus assumes candidates have already followed an introductory course in chemistry and it presents an overview of important concepts in the subject. Most of the concepts studied at intermediate level will not be developed to the same depth as at the Advanced Matriculation level but the syllabus is intended to cover key ideas that allow the student to understand better the nature of chemicals and the forces that drive chemical change.

### The Examination

The Examination consists of one three-hour paper. The paper will be divided into three sections: Section A will contain between eight and ten compulsory questions of the fill-in type requiring short answers; Section B will consist of between four and six compulsory structured questions; Section C will require candidates to choose two out of four long questions. Each of section A and B will carry 30 marks while section C will carry 40 marks. The minimum mathematical requirements of the syllabus are similar to those for the SEC examination in Chemistry and additional mathematical skills will be indicated in the relevant section of the syllabus. 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, ions and the 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  $^{12}\text{C}$ . 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 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. The general gas equation  $pV = nRT$  and its use in the determination of the relative molecular mass of a gas or vapour (calculations limited to simple examples). Use of gas syringe for the determination of relative molecular mass.

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 Electron gain and electron loss in atoms: ionization energies and relation to electronic configuration and valency (experimental determination of ionization energies not required). Periodic Table of elements: valency as a periodic property. Electron affinity.
4. Chemical bonding.
  - 4.1 Ionic bonding as electrostatic attraction between oppositely charged ions. Lattice structure as exemplified by NaCl.
  - 4.2 Covalent bonding as sharing of electrons between two atoms. Single and multiple bonds. Dative covalent (coordinate) bonds. Simple molecules. e.g. CO<sub>2</sub>, and macromolecules (network covalent) as exemplified by SiO<sub>2</sub>, graphite and diamond.
  - 4.3 Octet rule and use in predicting formulae: exceptions to octet rule to include electron deficient molecules (e.g. AlCl<sub>3</sub>) and 'octet expansion' (e.g. PCl<sub>5</sub>).
  - 4.4 Polarity of bonds. Concept of electronegativity and intermediate bonding: limited to polar covalent bonds.
  - 4.5 Concept of resonance: delocalisation of electrons as exemplified by benzene. The concept of orbitals will not be tested.
  - 4.6 Shapes of molecules and simple ions in terms of electron pair repulsion theory e.g. BeF<sub>2</sub>, BF<sub>3</sub>, CH<sub>4</sub>, SF<sub>6</sub>, NH<sub>3</sub>, H<sub>2</sub>O, NH<sub>4</sub><sup>+</sup>, H<sub>3</sub>O<sup>+</sup>.
  - 4.7 Intermolecular forces: van der Waals' forces to include permanent dipole-permanent dipole and induced dipole-induced dipole forces. 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 as evidence of the presence of a vapour in contact with the evaporating liquid/solid; the saturated vapour. 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  $\Delta H$  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 based 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 (rate equation will not be examined). 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 (effect on equilibrium of addition of inert gas at constant volume will not be tested). Formation of ammonia from its elements (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 as a useful tool for predicting redox reaction involving metal ions.
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 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. Students are expected to be familiar with the logarithmic function as required for calculation of pH.
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 and chlorides Trends in melting points of the oxides.

## II DESCRIPTIVE CHEMISTRY

12. Chemistry of transition metals. Colour and variable oxidation states as shown by the species  $\text{Mn}^{2+}$ ,  $\text{MnO}_2$ ,  $\text{MnO}_4^{2-}$  and  $\text{MnO}_4^-$ ,  $\text{Fe}^{2+}$  and  $\text{Fe}^{3+}$ ,  $\text{Cu}^+$  and  $\text{Cu}^{2+}$ . Catalytic properties as exemplified by use of  $\text{MnO}_2$  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 (coordinate) bonding; examples of complex ions, including the molecular shape to include:  $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$ ,  $[\text{Zn}(\text{NH}_3)_6]^{2+}$ ,  $[\text{Cu}(\text{NH}_3)_4]^{2+}$ ,  $[\text{CuCl}_2]^-$ ,  $[\text{Fe}(\text{CN})_6]^{3-}$ . Nomenclature of coordination compounds and discussion of isomerism are not required. Extraction of metals from ores is not required.
13. Chemistry of the halogens: chlorine, bromine, iodine. Occurrence. Uses of chlorine. Relative oxidising power of elements. Reaction of  $\text{Cl}_2$  and  $\text{I}_2$  with thiosulfate. Hydrogen halides: preparation from salts, acidity of aqueous solutions. The chlorate(I) ion and its application as a bleaching and disinfecting agent: discussion of other 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. Addition reactions of alkenes with halogens and hydrogen halides as examples of ionic mechanisms. Markownikov's Rule to predict addition of HX to unsymmetrical alkenes, preparation of alkenes by dehydration of alcohols; ethyne from calcium dicarbide. Hydrogenation of alkenes. Alkanes and alkenes from petroleum by cracking.
- 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  $\text{PCl}_5$ , 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  $\text{PCl}_5$ ; (d) amines: salt formation from primary amines and comparison with ammonia, 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 (qualitative treatment only).
- 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: standard solutions, acid-base and redox titrations and related calculations in terms of moles and concentrations in terms of  $\text{mol dm}^{-3}$ . Redox titrations are limited to manganate(VII) – iron(II) and iodine - thiosulfate. Back titration involving addition of known excess of acid or alkali to a reactant solution and determining the excess by titration. Gravimetric methods of analysis.

### 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., AQA Chemistry AS: Students' Book, Nelson Thornes Ltd., 2008.
2. Lewis, R. and Evans W., Chemistry, 4<sup>th</sup> Edition, Palgrave Macmillan 2011.
3. Hill GC and Holman JS, Chemistry in Context, 5<sup>th</sup> Edition, Nelson Thornes, 2000.