Physics and Chemistry

Note: SI Units will be used in the examinations.

1. Displacement. Velocity. Acceleration. Mass. Newton's laws of motion and gravitation. Conservation of momentum. Force. Weight.

   Coulomb's law. Electric fields. Field strength and potential. Capacitors and capacitance.

   Simple use of equations of motion, including vertical motion under gravity. Newton's three laws of motion, their meaning and application. Laws of universal gravitation: relation between G and g. Force as rate of change of momentum and a special case of force equals mass multiplied by acceleration. Simple problems. Weight as a force. Electrification by friction, charging by induction. Coulomb's law of force. Electric field patterns. Capacitators in series and parallel: factors affecting capacitance.

2 Work. Energy. Kinetic and Potential Energy. Conservation of energy. Concept of temperature. Celsius Scale. Gas thermometer. Definition of gas scale. Absolute scale. Gas laws. General gas equation. Qualitative treatment of pressure and temperature on the basis of the kinetic theory.

   Definition of work; its units (joules). Energy as ability to do work. Conservation of energy: heat a form of energy. Equivalence of work and heat. Definition of temperature in terms of change in physical properties of materials e.g. Celsius scale as given by

$$\frac{t}{100} = \frac{X_t - X_o}{X_{100} - X_o}$$

   where X is some property which varies as heat is added.

   Superiority of gas thermometer as being independent of thermometric substance i.e. Charles' law. The gas scale. Absolute zero as the origin of the extrapolated volume/temperature graph for all gases; its fundamental nature. Boyle's law. Equivalence of constant pressure and constant volume gas scales.

   General gas equation and simple calculations. Demonstration of molecular agitation. (Brownian movement).

3. Laws of reflection and refraction of light. Formation of images by mirrors and lenses; telescope and microscope.

   Laws of reflection and refraction of light treated experimentally. Refractive index and total internal reflection. Real and virtual images. Simple exercises on single lens, mirror, by either ray-tracing or application of formula. Sign convention optional. Method of measuring focal lengths of lenses and mirrors, method optional. Magnification. Optical system of simple telescope, and, of compound microscope, without calculations.

4. Transverse and longitudinal waves. Meaning of frequency, wavelength, amplitude, velocity. Interference and diffraction patterns: ripple tank. Interference and diffraction of light waves. Nature of light. Electromagnetic spectrum.

   Measurement of wavelength by any interference method. Reference to wave/particle nature of light. Reference to photoelectric effect. Dispersion of light by prism, visible spectrum, infrared and ultra violet, electromagnetic spectrum.

5. The structure of the atom: electrons, protons, neutrons, atomic number, mass number, atomic weight. Radioactivity and isotopes. Emission of alpha, beta and gamma rays. Nuclear reactions, fission and fusion. Mass energy conservation. Energy levels. Electron probability distribution (qualitative treatment).

   Orbitals. Ionisation energy.
   Reference to charge and location of sub-atomic particles. Atomic nuclei as protons plus neutrons. Blackening of photographic plates by ionisation produced by radiations from radioactive substances. Isotopes as having the same atomic (proton) number, different neutron number. Properties of alpha, beta and gamma radiation, e.g. relative ionizing ability and absorption of emissions. Concept of half-life. Simple examples of nuclear reactions; principles of fission and fusion.
   Emphasis on energy levels rather than Bohr orbits; reference to spectra, giving relationship $E_1 - E_2 = h\nu$, otherwise qualitative treatment. Although not strictly exact, the charged cloud interpretation of probability density may be used. Brief treatment of quantum numbers. Measurement of ionization energy not required.

6. The Periodic Table in relation to atomic structure: the structure and valence of elements as exemplified by the first thirty-six elements of the Periodic Table.
   Reference to variable valence in Transition Elements is sufficient.

7. Ionic, covalent, metallic bonds. Electronegativity. Polar covalent bonds. Simple examples of ionic, molecular, metallic and covalent crystals, and general properties associated with these types of crystals.
   Hybrid orbitals need not be considered.
   Electronegativity to be considered as the relative tendency of an atom in forming a bond to attract a shared electron pair. Simple treatment of dipole moment but not its measurement.

8. Structure and shape of simple molecules of both organic and inorganic compounds (linear, planar and tetrahedral).
   Use electron pair repulsion theory. Suggested examples BeH_2, BF_3, CH_4, NH_3, H_2S, H_2O.
   Presence or absence of overall dipole moment in deciding molecular shape.

9. Electric current as a flow of electrons. Magnetic effect of current. Definition of unit current. Potential difference. Relations between current and potential difference for various kinds of conduction. Ohm's law. Resistance. Heating effect of current.

   Magnetic field patterns due to current flowing in straight wire, loop and solenoid. Simple electrical calculations.

10. Methods of producing magnetic fields. Force on current-carrying conductor in magnetic field. Moving-coil meters and galvanometers. Electromagnetic induction: elementary ideas on alternating current.
    Electromagetic induction; examples dynamo ,transformer.

    Use of high voltage in transmission of electric current, with reasons
    Use of alternating current.

11. Relative atomic mass (atomic weight), relative molecular mass (molecular weight), Formula weight. Avogadro constant and the mole. Chemical equations.

    The mole should be considered both as the formula weight and as the Avogadro number of units. Calculation of formulae, from composition by weight and of yields, etc., from chemical equations.

12. First Law of Thermodynamics. Hess's law. Heats of combustion, formation, solution, reaction.
Simple calculation based on Hess's law. The delta H convention should be used.

13. The Brønsted-Lowry theory of acid and basis. Strong acids, pH and indicators. Acid-base titrations.
Conjugate acid-base pairs. Strenght of acids and bases – the equilibrium lies on the side of the weaker pair; pH of strong acids and bases. Use of indicators in titrations.

14. Oxidation and reduction as electron transfer; voltaic cells; electrochemical series. Qualitative description of electrical conduction in solution. Faraday's laws: electrolysis of water and simple melts.
Simple electrochemical cells should be considered in connection with the electrochemical series but oxidation potential need not be covered. The oxidation and reduction of water itself should be considered in electrolysis.

15. Hydrogen, oxygen, water, hydrogen peroxide. Comparison of the single binary compounds of the following with hydrogen: Na, K, Ca, Mg, C, N, S, halogens.
Comparisons of the oxides of Na, K, Mg, Ca, Al, C, N, S, Fe, Cu
Comparisons of the chlorides of Na, K, Ca, Mg, C, P, Fe, Cu
Examples of preparation of typical binary compounds of the elements with H_2 and with O_2. This need not be detailed or extensive.
The periodic Table should be the basis for the study of this section of the syllabus. It is the intention that trends in properties should be emphasised – a detailed study of all the compounds is not required. It is intended that this section should be further used to illustrate and develop the ideas of bonding, the electrochemical series and acid-base character that have already been advanced.
Binary hydrogen compounds of the elements may be considered under the headings (a) general appearance (solid, liquid, gas, colour), (b) bond type leading to (c) reaction (if any) of compound with water giving a solution which may be considered from the point of view of (d) acid base character and that of salts formed – solubility and appearance.
Binary compounds with oxygen. (a) as above; (b) Bond type. Basic, acidic, neutral, amphoteric oxides. (c) as above; (d) as above plus salts formed by action of acids on oxides.
Chlorides: bond type, appearance, behaviour in water.

16. (i) Hydrocarbons. Characteristic properties of
    (a) Aliphatics: -
        Alkanes: methane and its homologues; (as far as butane).
        Alkenes: ethane (ethylene) and its homologues (as far as butene).
        Alkynes: acetylene
    (b) Aromatics: benzene; mono-substitution reactions (bromination and nitration).

    (ii) The hydroxyl group in alcohols and phenols with special reference to methanol, ethanol and phenol.

    (iii) The hydroxyl group in acids with special reference to formic, acetic and benzoic acids. Acid derivatives: esters.

    (iv) The carbonyl group in aldehydes, with special reference to acetaldehyde. The carbonyl group in ketones, acetone. Detailed preparations are *not* required.

Reactions considered should include:

(i) (a) Chlorination of methane and homologues; addition of bromine and chlorine to ethene and its homologues; Baeyer test with KMnO; hydration

of ethene; addition of bromine and chlorine to acetylene; reaction to acetylene with sodium in liquid ammonia; hydration of acetylene.

(b) as syllabus

(ii) Methanol and ethanol – oxidation and dehydrogenation; reaction with Na; reaction with HCl; reaction of ethanol with H_2SO, giving (a) ethane, (b) ether.
Phenol – preparation; acidic properties; reaction with bromine water.

(iii) Acidic properties of carboxyl group; preparation of benzoic acid from toluene. Solubility of acids. Preparation of esters of acids given. Acid and base catalysed hydrolysis of derivatives.

(iv) Reaction (or lack of it) of the given compounds with oxidising agents. Reaction with phenylhydrazine.