Atomic and Molecular Structure

Atomic Structure

1. **The Hydrogen Atom**: Description of the probability model of the hydrogen atom.
(IB: Topic 2.1 & Topic 2.3)

1.1 Describe Dalton's atomic theory
1.2 Define the terms element, compound, atom, molecule and isotope
1.3 Describe the development of the Thomson model of the atom
1.4 Describe how Rutherford used alpha particles to probe atomic structure.
1.5 Compare the properties of isotopes of an element.
1.6 Discuss the uses of radioisotopes.
1.7 Describe the nuclear model of the atom.
1.8 Describe the main characteristics of an electromagnetic spectrum.
1.9 Describe the line spectrum of hydrogen.
1.10 Distinguish between a continuous spectrum and a line spectrum.
1.11 Explain how the lines in the emission spectrum of hydrogen are related to electron energy levels.
1.12 Use the Balmer equation to predict the wavelengths of the lines in the visible spectrum of the hydrogen atom.
1.13 Describe the Bohr model of the hydrogen atom.
1.14 Understand the relevance of the Balmer equation in terms of the Bohr model.
1.15 Calculate the energy changes involved for all possible electron transitions in a hydrogen atom.
1.16 Calculate the ionization energy of hydrogen.
1.17 Understand the limitations of the Bohr model.

2. The mass spectrometer
(IB: Topic 2.2)

2.1 Describe and explain the operation of a mass spectrometer.
2.2 Describe how the mass spectrometer may be used to determine the relative atomic mass using the ^{12}C scale.
2.3 Calculate the relative atomic masses and abundance of isotopes from given data.

3. **The Periodic Table**: description of the probability model of the multi-electron atom.
(IB: Topic 12)

3.1 Describe the photo-electric effect and wave particle duality and apply the de Broglie equation.
3.2 Describe the quantum numbers n, l, m_l and m_s.
3.3 Sketch boundary surface diagrams for *s* and *p* orbitals.
3.4 State the relative energies of the *s*, *p*, *d*, and *f* orbitals in a single energy level
3.5 State the maximum number of orbitals in a given energy level.
3.6 Describe the Pauli Exclusion and Hund's rule.
3.7 Use the Aufbau method to predict the ground state electron configuration of atoms and ions. From hydrogen through to xenon (Z = 54). Exceptions for chromium and copper should be known.

Periodicity

4. Periodic properties of the Elements
(IB: Topic 12.1.1 — 12.1.4)

4.1 Describe and predict trends in: atomic radius, ionic radius, ionization energy and electron affinity.
4.2 Explain how evidence from first ionization energies across periods accounts for the existence of main energy levels and sub-levels in atoms.
4.3 Explain how successive ionization energy data is related to the electron configuration of an atom.
4.4 Describe whether an atom or an ion is diamagnetic or paramagnetic.

5. Periodicity: Trends across period 3 (IB: Topic 13.1)

5.1 Explain the physical states and the electrical conductivity (in the molten state) of the chlorides and the oxides of the elements in period 3 in terms of their bonding and structure.
5.2 Describe the reactions of chlorine and the chlorides of period 3 with water.

1. First -row d-block elements
(IB: Topic 13.2)

6.1 List the characteristic properties of transition elements.
6.2 Explain why Sc and Zn are considered to be transition elements.
6.3 Explain the existence of variable oxidation number in ions of transition elements.
6.4 Define the term ligand.
6.5 Describe and explain the formation of complexes of d-block elements.
6.6 Explain why some complexes of d-block elements are coloured.
6.7 State examples of the catalytic action of transition elements and their compounds.
6.8 Outline the economic significance of catalysts in the Contact and Haber processes.

Bonding

2. Ionic Bonding
(IB: Topic 4.1)

7.1 Describe the ionic bond as the electrostatic attraction between oppositely charged ions.
7.2 Describe how ions can be formed as a result of electron transfer.
7.3 predict whether a compound of two elements would be ionic from the position of the elements in the periodic table or from their electronegativity values.
7.4 Describe the lattice structure of ionic compounds.
7.5 Describe the typical physical properties of ionic compounds. Describe the solution process of an ionic compound and to explain the factors affecting the enthalpy of solution.

Formation of Binary Ionic Compounds: Born — Haber cycle
(IB Topic 15.2)

7.5 Define and apply the terms lattice energy and electron affinity.
7.6 Construct a Born – Haber cycle for Group 1 and 2 oxides and chlorides and use it to calculate an enthalpy change.
7.7 Explain how the relative sizes and the charges of ions affect the lattice energies of different ionic compounds.
7.8 Describe the difference between theoretical and experimental lattice energy values of ionic compounds in terms of their covalent character.

3. The Covalent Bond Model

(IB: Topic 4.2)

8.1 Describe the covalent bond as the electrostatic attraction between a pair of electrons and positively charged nuclei.

8.2 Deduce the Lewis structures of molecule and ions (up to five and six charge centers).

8.3 State and explain the relationship between the number of bonds, bond length and bond strength.

8.4 Describe and know the main periodic trends in the electronegativity of atoms.

8.5 Relate bond polarity to electronegativity differences between bonded atoms.

4. Molecular Geometry

9.1 Describe the VSEPR Model and use it to predict electron pair geometry, molecular geometry and bond angles.

(IB: Topic 14.1) Predict the shape and bond angles for species with 5 and 6 negative charge centers using the VSEPR theory.

9.2 **(IB: Topic 14.3)** Describe the concept of delocalization of electrons, the concept of resonance and draw the contributing structures of a resonance hybrid; and explain how this can account for the structures of some species, e.g. NO_3^{-1}, NO_2^{-1}, CO_3^{-2}, O_3, $RCOO^{-1}$.

9.3 Use the VSEPR Model to predict the molecular polarity of a structure.

5. Intermolecular Forces

(IB: Topic 4.3)

10.1 Describe each if the following intermolecular forces: dispersion force, (London dispersion force, van der Waals' force), dipole—dipole force and hydrogen—bonding. Explain how they arise from the structural features of the molecules.

10.2 Describe the effects of increasing dispersion forces on boiling and melting points.

10.3 Describe the effects of dipole–dipole forces on boiling and melting points.

10.4 Describe the effects of hydrogen bonding on the boiling point and melting point.

6. Structure and Types of Solids

(IB: Topic 4.2.9 & Topic 4.2.10)

11.1 Describe and compare the structure and bonding in the three allotropes of carbon (diamond, graphite and C_{60} fullerene).

11.2 Describe the structure of and bonding in silicon and silicon dioxide.

7. Metallic Bonding

(IB: Topic 4.4)

12.1 Describe the metallic bond as the electrostatic attraction between a lattice of positive ions and delocalized electrons.

12.2 Explain the electrical conductivity and malleability of metals.

8. Bonding: Hybridization

(IB: Topic 14.2)

13.1 Use valence bond theory and orbital hybridization to explain σ and π bonds.

13.2 Predict whether any central atom is *sp*–, sp^2 — or sp^3 — hybridized.

13.3 Explain the relationships between Lewis structures, molecular shapes and types of hybridization.