DINAKRUSHNA COLLEGE, JALESWAR

Department of Chemistry

QUESTION BANK

Paper - Core I Major I: Atomic Structure, Periodicity of Elements and Chemical Bonding

Unit / Chapter	Fill in Blanks (1 Mark)	Short Questions (2 Marks)	Medium Questions (5 Marks)	Long Questions (8 Marks)	Total
Unit-I: Atomic Structure	10	10	10	5	35
Unit-II: Periodicity of Elements	10	10	10	5	35
Unit-III: Chemical Bonding (Ionic)	10	10	10	5	35
Unit-IV: Chemical Bonding (Covalent)	10	10	10	5	35
Total Prepared	**40**	**40**	**40**	**20**	**140**

Unit-I: Atomic Structure

Learning Outcome: *The learners are able to solve conceptual questions using the knowledge gained by studying the quantum mechanical model of the atom.*
Detailed Syllabus: *Rutherford's nuclear model of atom, Bohr's theory and the origin of hydrogen spectrum, Sommerfeld's extension of Bohr's theory, de-Broglie equation, Heisenberg's Uncertainty Principle and its significance. Postulates of wave mechanics, Derivation of Schrödinger's wave equation for hydrogen atom, significance of ψ and ψ^2. Radial and angular wave functions, Radial function plots, radial probability distribution plots, angular distribution curves. Shapes of s-, p-, d- and f-orbitals, Relative energies of orbitals. Slater's rule and its limitations, Quantum numbers and their significance. Pauli's Exclusion Principle, Hund's rule of maximum spin multiplicity and Aufbau principle.*

I. Fill in the Blanks (1 Mark each)

1. The mathematical relationship representing the wave-particle duality of matter is given by the _____ equation.

2. The probability density of finding an electron at a specific point in space is represented by the term _____.

3. According to Sommerfeld’s extension, electron orbits in an atom are not only circular but can also be _____.

4. The quantum number that describes the spatial orientation of an orbital is the _____ quantum number.

5. A region in space where the probability of finding an electron is zero is called a _____.

6. The rule states that no two electrons in an atom can have the same set of all four quantum numbers is _____ Exclusion Principle.

7. According to _____ rule, electron pairing in a subshell cannot occur until each orbital is singly occupied.

8. The effective nuclear charge felt by an electron is calculated by subtracting the screening constant from the atomic number using _____ rules.

9. The angular distribution function of a 1s orbital is independent of the angles θ and ϕ, giving it a _____ shape.

10. The principle stating that orbitals are filled in the order of increasing energy is the _____ principle.

II. Short Answer Questions (2 Marks each)

1. State Heisenberg's Uncertainty Principle and write its mathematical form.

2. Differentiate between an orbit (Bohr) and an orbital (Quantum mechanical).

3. What is the physical significance of ψ^2?

4. Define the term 'nodal plane' and find the total number of nodal planes in a 3p orbital.

5. State the limitations of Rutherford’s nuclear model of the atom.

6. Write down the de-Broglie relation and define each term involved.

7. Explain why the electronic configuration of Chromium is [Ar] $3d^5$ $4s^1$ instead of [Ar] $3d^4$ $4s^2$.

8. Define Slater's screening constant (σ).

9. What information is obtained from the Azimuthal Quantum Number (l)?

10. Sketch the shapes of $d_{x^2-y^2}$ and d_{z^2} orbitals showing their axes.

III. Medium Answer Questions (5 Marks each)

1. Derive the de-Broglie relationship and discuss its experimental significance for macroscopic objects.

2. Explain the origins and line series (Lyman, Balmer, Paschen, etc.) of the hydrogen spectrum based on Bohr's theory.

3. Formulate the time-independent Schrödinger wave equation for a hydrogen atom and list its major postulates.

4. Draw and compare the radial probability distribution curves for 2s and 2p orbitals.

5. Explain Slater's rules for calculating the effective nuclear charge (Z*); give one sample calculation for a 3d electron of Iron (Z = 26).

6. Describe Sommerfeld’s extension of Bohr’s atomic model and explain how it addresses fine spectrum lines.

7. Discuss the shapes, orientations, and nodes of the seven f-orbitals.

8. Explain the (n + l) rule with suitable examples to justify the filling order of 4s before 3d.

9. Discuss the concept of exchange energy and its role in the exceptional stability of half-filled and fully-filled electronic configurations.

10. Define radial and angular wave functions. How do they combine to give the total wave function of an electron?

IV. Long Answer Questions (8 Marks each)

1. Setup the Schrödinger wave equation for a hydrogen-like system in spherical polar coordinates and discuss the significance of separating it into radial and angular parts.

2. Discuss Bohr's atomic theory in detail, deriving expressions for the radius and energy of an electron in the n-th orbit of a hydrogen atom. Mention its limitations.

3. State and explain the three fundamental principles (Aufbau Principle, Pauli's Exclusion Principle, and Hund's Rule) that govern the building-up of electronic structures in multi-electron atoms.

4. Write a comprehensive essay on quantum numbers, detailing their origins from the wave equation, physical significance, and permissible values.

5. Provide a detailed comparative account of the boundary surface diagrams, radial nodes, angular nodes, and total nodes of s, p, and d block elements.

Unit-II: Periodicity of elements

Learning Outcome: *The learners understand the various atomic properties of atoms and their variations in the periodic table.*
Detailed Syllabus: *Introduction to long form periodic table, Cause of periodicity, Division of elements into s-, p-, d- and f-blocks. Atomic radius, ionic radius, covalent radius and Van der Waals radius. Periodic trends in ionic and covalent radii. Ionization energy, electron affinity, electronegativity, and their variations in the periodic table. Applications of electronegativities, Pauling's/Mulliken's scale of electronegativity, Sanderson's electron density ratio.*

I. Fill in the Blanks (1 Mark each)

1. The modern periodic law states that the physical and chemical properties of elements are periodic functions of their _____.

2. The distance between two non-bonded touching atoms of adjacent molecules in a solid state is known as the _____ radius.

3. Across a period from left to right, the atomic radius generally _____ due to an increase in effective nuclear charge.

4. The energy released when an electron is added to a neutral gaseous atom is termed _____.

5. Element groups 3 to 12 in the periodic table constitute the _____-block elements.

6. The electronegativity scale based on the arithmetic mean of ionization energy and electron affinity is _____ scale.

7. _____ radius is always smaller than the corresponding atomic radius of a neutral parent atom.

8. The scale of electronegativity that utilizes the concept of electron density ratio is the _____ scale.

9. Nitrogen has a higher first ionization energy than Oxygen due to its stable _____ electronic configuration.

10. On Pauling's scale, Fluorine is assigned a maximum electronegativity value of _____.

II. Short Answer Questions (2 Marks each)

1. Explain the primary cause of periodicity in the properties of elements.

2. Arrange the following in increasing order of size and justify your answer: N^{3-}, O^{2-}, F^{-}, Na^{+}.

3. Why is the second ionization energy of Sodium significantly higher than its first ionization energy?

4. Define covalent radius and show how it differs from Van der Waals radius.

5. State Mulliken’s formula for calculating the electronegativity of an element.

6. Why are the electron affinity values of noble gases positive (or zero)?

7. How does the metallic character change down a group and across a period?

8. Define Sanderson's electron density ratio.

9. Why is the electron affinity of Chlorine greater than that of Fluorine?

10. Mention two important applications of electronegativity values in predicting bond nature.

III. Medium Answer Questions (5 Marks each)

1. Classify the elements into s, p, d, and f blocks based on their valence shell configurations and summarize their characteristic properties.

2. Discuss the variations of ionization energy across a period and down a group, highlighting exceptions in the second period.

3. Explain Pauling's thermodynamic scale of electronegativity and show how it relates to bond dissociation energies.

4. Describe the concept of lanthanide contraction and explain its consequences on the atomic radii of 5d transition elements.

5. Explain how electronegativity difference is utilized to calculate the percentage ionic character of a heteronuclear diatomic bond.

6. Detail the periodic trends of electron affinity and discuss the factors influencing its values.

7. Compare and contrast Pauling's and Mulliken's scales of electronegativity. Show how their values can be interconverted.

8. Discuss the variation of valency and oxidation states across a period and down a group.

9. Explain why the atomic radius of Gallium is slightly smaller than that of Aluminum.

10. Elaborate on how electronegativity can be used to explain the acidic or basic nature of oxides in p-block elements.

IV. Long Answer Questions (8 Marks each)

1. Discuss in detail the periodic trends of atomic, ionic, covalent, and Van der Waals radii. Explain how these trends are affected by effective nuclear charge and principal quantum numbers.

2. Give a comprehensive account of electronegativity. Discuss its derivation using Pauling, Mulliken, and Sanderson approaches, and detail its wide applications in chemistry.

3. Write an essay on ionization energy, outlining experimental determination principles, contributing factors (shielding, penetration, configurations), and its systematic periodic variations.

4. Provide a detailed account of the structure of the long form periodic table. Explain how the electronic configuration of elements underpins their division into periods and blocks.

5. Explain the differences between electron affinity and electronegativity. Discuss why periodic trends in electron affinity exhibit more irregularities than trends observed in ionization energies.

Unit-III: Chemical bonding (Ionic Bond)

Learning Outcome: *The learners gain the idea of different types of bonding and their associated properties (Ionic systems).*
Detailed Syllabus: *Ionic bond-General characteristics, types of ions, size effects, radius ratio rule and its limitations. Packing of ions in crystals, Lattice energy, Born-Haber cycle and its application, Born-Lande equation, Madelung constant, importance of Kapustinskii equation for lattice energy. Solvation energy, Covalent character in ionic compounds, polarizing power and polarizability. Fajan's rules and consequences of polarization.*

I. Fill in the Blanks (1 Mark each)

1. The coordination number of a cation in an ionic crystal is determined by its _____ ratio.

2. The cyclic process used to determine the lattice energy of an ionic crystal experimentally is the _____ cycle.

3. The proportionality constant in the Born-Landé equation that accounts for the geometric arrangement of ions is the _____ constant.

4. According to Fajan's rules, a smaller cation possesses a higher _____ power.

5. The equation that allows estimation of lattice energy without knowing the exact crystal geometry is the _____ equation.

6. The energy released when one mole of gaseous ions interacts with water molecules is called _____ energy.

7. An ionic compound will be more soluble in water if its solvation energy is greater than its _____ energy.

8. Large, highly charged anions exhibit high _____ when placed near a cation.

9. If the radius ratio (r_+/r_-) falls in the range 0.414 - 0.732, the expected coordination number is _____.

10. The phenomenon where a cation distorts the electron cloud of an anion is termed _____.

II. Short Answer Questions (2 Marks each)

1. State the Radius Ratio Rule.

2. Define Lattice Energy of an ionic crystalline solid.

3. What is the physical significance of the Madelung constant (A)?

4. State Fajan's rules regarding the charge of the cation.

5. Why is LiI more soluble in organic solvents like ethanol than LiF?

6. Define the Born exponent (n) used in the Born-Landé equation.

7. Why does the r_+/r_- rule often fail to predict actual crystal structures?

8. Differentiate between polarizing power and polarizability.

9. Write down the mathematical form of the Kapustinskii equation and define its variables.

10. Why are transition metal cations (e.g., Cu^+) more polarizing than alkali metal cations of similar size (e.g., Na^+)?

III. Medium Answer Questions (5 Marks each)

1. Explain the close packing of spheres in crystals, differentiating between hexagonal close packing (hcp) and cubic close packing (ccp).

2. Construct a complete Born-Haber cycle for the formation of one mole of crystalline NaCl from its constituent elements, labeling all energy terms involved.

3. Derive the expression for the radius ratio threshold (0.414) for an octahedral coordination geometry.

4. Discuss the factors that influence the magnitude of the solvation energy of an ionic compound.

5. State the Born-Landé equation for calculating lattice energy. Discuss how the terms for attractive and repulsive forces are accounted for in its derivation.

6. Explain the consequences of polarization in ionic compounds with respect to melting points and solubility.

7. Discuss the importance and advantages of the Kapustinskii equation over the Born-Landé equation.

8. Explain why AgCl is white and insoluble in water, whereas AgF is clear and highly soluble, based on Fajan’s concepts.

9. Describe the types of ions and size effects that favor the formation of a stable, highly exothermic ionic lattice structure.

10. Explain the relationship between lattice energy, solvation energy, and the thermodynamic solubility behavior of alkaline earth metal sulfates.

IV. Long Answer Questions (8 Marks each)

1. Give a detailed account of the Born-Haber cycle. Explain how it can be applied to calculate electron affinity, lattice energy, and to predict the stability of non-existent compounds like NeCl.

2. State and explain Fajan's rules in detail. Provide appropriate examples to illustrate how cation size, anion size, electronic configuration, and ionic charges introduce covalent character into ionic bonds.

3. Describe the structural arrangements of ions in NaCl, CsCl, and Zinc Blende (ZnS) crystals. Discuss how their stability correlates with radius ratio values.

4. Derive the Born-Landé equation from fundamental electrostatic interactions and quantum mechanical repulsions. Outline its basic parameters and limitations.

5. Provide a detailed comprehensive essay on the energetic pathways of ionic compound formation, contrasting the competitive roles of lattice enthalpy versus hydration enthalpy during dissolution.

Unit-IV: Chemical bonding (Covalent Bond)

Learning Outcome: *The learners understand covalent models, molecular geometry rules, and orbital structural frameworks.*
Detailed Syllabus: *Covalent bond-Valence shell electron pair repulsion (VSEPR) theory, shapes of the following simple molecules and ions containing lone pairs and bond pairs of electrons: CH4, H2O, NH3, PCl3, PCl5, SF6, ClF3, $I3^-$, $BrF2^+$, $PCl6^-$, $ICl2^-$, $ICl4^-$, $NH4^+$, $PO4^{3-}$ and $SO4^{2-}$. Valence Bond theory (Heitler-London approach), Hybridization, equivalent and non-equivalent hybrid orbitals, Ionic character in covalent compounds: Dipole moment. Percentage ionic character from dipole moment and electronegativity difference, Molecular orbital diagrams of homo- & hetero-diatomic molecules (N2, O2, C2, B2, F2, CO, NO) and their ions. Calculation of bond order. Concept of bent rule.*

I. Fill in the Blanks (1 Mark each)

1. According to VSEPR theory, the order of electron pair repulsive interactions is lp-lp > _____ > bp-bp.

2. The hybridization state of the central Phosphorus atom in a PCl5 molecule is _____.

3. The molecular geometry of the $I3^-$ ion is _____.

4. According to Bent's rule, more electronegative substituents prefer to occupy hybrid orbitals having _____ s-character.

5. A molecule is non-polar and has a net dipole moment of zero if its spatial structure is perfectly _____.

6. The Heitler-London approach forms the conceptual foundation for the _____ theory of chemical bonding.

7. In a Molecular Orbital diagram, if the number of bonding electrons equals the number of antibonding electrons, the bond order is _____.

8. The molecule O2 is paramagnetic due to the presence of two unpaired electrons in its _____ molecular orbitals.

9. Hybrid orbitals that have identical shapes, sizes, and mixing ratios are classified as _____ hybrid orbitals.

10. The CO molecule is an example of a _____-nuclear diatomic molecule.

II. Short Answer Questions (2 Marks each)

1. Why is the bond angle in H2O (104.5°) smaller than that in NH3 (107.5°)?

2. Predict the shape and hybridization state of the ClF3 molecule.

3. Define dipole moment and state its SI unit.

4. What is meant by non-equivalent hybrid orbitals? Give an example.

5. Calculate the bond order of the $O2^+$ ion.

6. State the fundamental difference between a bonding molecular orbital and an antibonding molecular orbital.

7. Why does the B2 molecule exhibit paramagnetism according to MO theory?

8. Draw the Lewis structure and assign the shape of the $SO4^{2-}$ ion.

9. State the core concept of Bent's rule.

10. Write down the equation used to calculate percentage ionic character from experimental dipole moments.

III. Medium Answer Questions (5 Marks each)

1. Discuss the key postulates of the Valence Shell Electron Pair Repulsion (VSEPR) theory.

2. Determine the hybridization, number of lone pairs, and precise spatial geometry of SF_6 and ICl_4^-.

3. Explain the Heitler-London approach of Valence Bond theory and how it explains stable covalent molecule formation using a potential energy curve.

4. Using Bent's rule, explain why the F-P-F axial bond angle in PCl_3F_2 is different from the equatorial angles.

5. Draw the Molecular Orbital diagram for the C_2 molecule and account for its bond order and magnetic behavior.

6. Explain the concept of sp^3d hybridization. Differentiate between its axial and equatorial bonds with reference to length and stability.

7. Calculate the percentage ionic character of an HF molecule if its observed dipole moment is 1.91 D and the bond length is 0.92 Å.

8. Describe the shapes of I_3^- and NH_4^+ ions using VSEPR guidelines.

9. Discuss the concept of orbital mixing (s-p mixing) and explain why the N_2 MO energy ordering differs from that of O_2.

10. Predict the molecular shapes and central atom configurations for BrF_2^+ and PCl_6^-.

IV. Long Answer Questions (8 Marks each)

1. Draw complete, neatly labeled Molecular Orbital energy level diagrams for the heteronuclear diatomic molecules CO and NO. Calculate their bond orders, discuss their magnetic properties, and explain their frontier molecular orbitals.

2. Provide a detailed structural account using VSEPR theory to deduce the shapes, ideal angles, and lone pair locations for the following set: H_2O, PCl_5, ClF_3, and ICl_2^-.

3. Discuss Valence Bond Theory comprehensively. Explain the concept of orbital hybridization, deriving mathematical expressions for equivalent sp^2 hybrid orbital coefficients.

4. Write a comprehensive essay on dipole moments. Detail how they are utilized to differentiate cis/trans isomers, determine ortho/meta/para substitution patterns, and evaluate the percentage ionic character of highly polar covalent links.

5. Discuss the structural features of PCl_5 and SF_6. Use these molecules to compare equivalent versus non-equivalent hybrid orbitals, and explain how Bent's rule governs substituent placement in trigonal bipyramidal systems.