Chemistry Assignment Questions and Answers

5. (a) Give the defects of Rutherford's model of atom. What suggestions were given by Bohr to remove these defects?

Defects of Rutherford's model of atom:

1. Instability of Atom: According to classical electromagnetic theory, an electron moving in a circular orbit around the nucleus should continuously lose energy in the form of radiation. As a result, the electron should spiral inward and eventually fall into the nucleus, making the atom unstable.

   However, atoms are stable in reality.
2. No Explanation of Atomic Spectra: Rutherford's model could not explain the discrete line spectra (specific wavelengths) observed in the emission or absorption spectra of atoms. If electrons could revolve at any distance, the energy changes should be continuous, not discrete.

Bohr's Suggestions to Remove These Defects:

1. Quantized Orbits: Bohr proposed that electrons revolve around the nucleus only in certain permitted, discrete orbits (called energy levels) without radiating energy.
2. Energy Emission or Absorption: Electrons can move between these orbits by absorbing or emitting a photon of energy equal to the difference between the two energy levels: $\Delta E = E_2 - E_1 = h\nu$ where h is Planck's constant and ν is the frequency of radiation.
3. Stability of Orbits: As long as an electron stays in a particular orbit (energy level), it does not lose energy, thus explaining the stability of the atom.

5. (b) What do you understand by the term, "Quantum number". How many quantum numbers has an electron in an orbital? Explain the significance of each quantum number?

Quantum Numbers:

Quantum numbers are a set of numbers used to describe the position, energy, and orientation of an electron in an atom. They give a complete address of an electron --- like its energy level, shape of orbital, orientation, and spin.

An electron in an orbital is described by four quantum numbers:

1. **Principal Quantum Number (n):**
   - Indicates the main energy level or shell of the electron.
   - It determines the size and energy of the orbital.
   - Values: n = 1, 2, 3, 4, ... (called K, L, M, N shells, etc.)
   - Higher n → higher energy and larger size of orbital.

2. **Azimuthal Quantum Number (l)**:
   - Indicates the shape of the orbital (also called the angular momentum
   - quantum number).
   - Values: l = 0 to (n-1)
     - Each value of l corresponds to a type of orbital:
     - l = 0 → s-orbital (spherical) l
     - = 1 → p-orbital (dumbbell-
     - shaped) l = 2 → d-orbital l =

   3 → f-orbital

3. **Magnetic Quantum Number (m or m_1):**
   - Indicates the orientation of the orbital in space.
   - Values: m = -l to +l (including 0)
   - For example, if l = 1 (p-orbital), then m = -1, 0, +1, meaning three different orientations.

4. **Spin Quantum Number (s or m_1):**
   - Indicates the direction of spin of the electron.
   - Values: m_1 = +½ (spin up) or m_1 = -½ (spin down)
   - Significance: Each orbital can hold two electrons with opposite spins.

6.(a) Compare the properties of ionic and covalent compounds. Give two examples of each type of compounds?

Comparison of Properties:

Property	Ionic Compounds	Covalent Compounds
Nature of Bond	Formed by transfer of electrons (metal + non-metal)	Formed by sharing of electrons (non-metal + non-metal)
Physical State	Usually crystalline solids	Usually gases, liquids, or soft solids
Melting and Boiling Points	High melting and boiling points	Generally low melting and boiling points
Electrical Conductivity	Conduct electricity in molten state or in solution (due to free ions)	Do not conduct electricity (no free ions), except some polar covalent compounds
Solubility	Soluble in water, but insoluble in organic solvents	Soluble in organic solvents, but insoluble in water
Strength	Hard and brittle	Generally soft and flexible (except diamond, etc.)

Examples:

- Ionic Compounds:
- Sodium chloride ($NaCl$)
- Magnesium oxide (MgO)
- Covalent Compounds:
- Water (H_2O)
- Carbon dioxide (CO_2)

6.(b) What is a co-ordinate covalent bond? How does it differ from a normal covalent bond?

Co-ordinate Covalent Bond:

- A co-ordinate covalent bond (also called a dative bond) is a type of covalent bond where both electrons shared in the bond come from the same atom.
- It usually forms when one atom with a lone pair of electrons donates both electrons to an atom that has an empty orbital.

Example:

In the formation of the ammonium ion (NH_4^+), the nitrogen atom donates a lone pair to bond with a hydrogen ion (H^+).

$NH_3 + H^+ \rightarrow NH_4^+$

Difference between Co-ordinate Covalent Bond and Normal Covalent Bond:

Feature	Normal Covalent Bond	Co-ordinate Covalent Bond
Electron Contribution	Each atom contributes one electron to the shared pair.	One atom donates both electrons to the bond.
Formation	Between atoms having incomplete octets.	One atom must have a lone pair, the other an empty orbital.
Representation	Simply a line (e.g., H--H)	Often shown by an arrow (→) from donor to acceptor atom.

7.(a) What do you understand by hydrogen bonds? Classify them with examples. Explain why water has abnormally high boiling point?

Hydrogen Bond:

- A hydrogen bond is a special type of attractive force that occurs between a hydrogen atom (attached to a highly electronegative atom like F, O, or N) and another electronegative atom.
- It is weaker than a covalent bond but stronger than van der Waals forces.

Classification of Hydrogen Bonds:

1. **Intermolecular Hydrogen Bond:**

- Occurs between molecules.
- Example: In water (H_2O), hydrogen bonds form between different water
- molecules.

$H_2O...H_2O$

2. **Intramolecular Hydrogen Bond:**
   - Occurs within the same molecule.
   - Example: In ortho-nitrophenol, a hydrogen bond forms between the --OH group and the --NO_2 group inside the same molecule.

Why Water Has an Abnormally High Boiling Point:

- In water, strong intermolecular hydrogen bonds hold the molecules tightly
- together.

A large amount of heat energy is needed to break these hydrogen bonds to convert water from liquid to gas.

- Therefore, water has a much higher boiling point compared to other similar-sized molecules (like H_2S, which does not form hydrogen bonds).

7.(b) Why bond angles of H_2O and NH_3 are 104.5° and 107° respectively although central atoms are sp^3 hybridized?

Bond Angles in H_2O and NH_3:

- In both H_2O and NH_3, the central atoms (Oxygen in H_2O and Nitrogen in NH_3) are sp^3 hybridized.
- Ideal sp^3 hybridization would give a tetrahedral angle of 109.5°.
- However, in H_2O and NH_3, the bond angles are less than 109.5° because of lone pair--bond pair repulsion.

Detailed Reason:

NH_3 (Ammonia):

✓ Nitrogen has one lone pair and three bond pairs.

✓ The lone pair exerts more repulsion on the bond pairs, slightly reducing the bond angle to 107° from 109.5°.

H_2O (Water):

✓ Oxygen has two lone pairs and two bond pairs.

✓ Two lone pairs exert even greater repulsion than in NH_3.

✓ Thus, the bond angle reduces even more to 104.5°.

8.(a) What do you mean by the "ionization potential" of an element? Why the first ionization potential of an element is less than the second ionization potential? How does the ionization potential of an element vary with atomic volume?

Ionization Potential (Ionization Energy):

- The ionization potential (or ionization energy) of an element is the amount of energy required to remove the most loosely bound electron from an isolated gaseous atom to form a cation (positively charged ion).

Atom (g) + Energy → Ion^+ (g) + e^-

Why First Ionization Potential is Less than Second Ionization Potential:

- First ionization potential refers to the energy needed to remove the first
- electron.
- After removing the first electron, the atom becomes positively charged.

In the positive ion, the remaining electrons are more tightly held due to the greater effective nuclear charge (more protons than electrons).

- Thus, removing a second electron requires more energy, making the second ionization potential greater than the first.

Variation of Ionization Potential with Atomic Volume:

- As atomic volume increases (i.e., as atoms become larger), outer electrons are farther from the nucleus and are less tightly held.
- Result: Ionization potential decreases with increasing atomic size.

In the Periodic Table:

- Across a period (left to right): Atomic size decreases → Ionization potential
- increases.

Down a group (top to bottom): Atomic size increases → Ionization potential decreases.

8.(b)What do you mean by f-block elements? Why f-block elements are called inner transition elements?

f-block Elements:

- f-block elements are the elements in which the last electron enters the f-orbital of the atom.
- They are placed separately at the bottom of the Periodic Table, in two rows:
- Lanthanides (elements 58 to 71)
- Actinides (elements 90 to 103)

Why f-block Elements are Called Inner Transition Elements:

They are called inner transition elements because:

- In these elements, the f-orbitals (which are inner orbitals) are being filled.
- The filling of f-orbitals occurs inside the (n-1)th shell, below the outermost energy level.
- Hence, the transition of electrons happens in an inner energy level, unlike d-block (transition) elements where d-orbitals are involved.