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

Ans: Defects of Rutherford’s Model:

1. Stability Issue: According to classical electromagnetic theory, electrons revolving around the nucleus should continuously emit energy due to acceleration and eventually spiral into the nucleus. This contradicts the stability of atoms.
2. No Explanation of Atomic Spectra: Rutherford's model could not explain the discrete lines observed in the atomic spectra, especially the hydrogen spectrum.
3. No Energy Levels: There was no concept of quantized energy levels for electrons.

Bohr’s Suggestions to Remove These Defects:

1. Quantized Orbits: Bohr proposed that electrons revolve in certain stable orbits (called energy levels or shells) without emitting energy.
2. Energy Emission/Absorption: Electrons emit or absorb energy only when they jump from one orbit to another.
3. Angular Momentum Quantization: He introduced the idea that the angular momentum of an electron is quantized and given by $L=n\hbar L$, where n is an integer and $\hbar$ is the reduced Planck constant.

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.

Ans: Quantum Numbers and Their Significance

Quantum Numbers: Quantum numbers are values that describe the state and position of an electron in an atom. Each electron in an atom is described by a unique set of four quantum numbers:

1. **Principal Quantum Number (n):**
   - Represents the main energy level or shell.
   - Determines the size and energy of the orbital.
   - Values: $n = 1, 2, 3, \ldots$
2. **Azimuthal Quantum Number (l):**
   - Describes the shape of the orbital.
   - Also called the orbital angular momentum quantum number.
   - Values: $l = 0$ to $(n - 1)$.
     - $l = 0$ (s), $l = 1$ (p), $l = 2$ (d), $l = 3$ (f)
3. **Magnetic Quantum Number (m):**
   - Represents the orientation of the orbital in space.
   - Values: $m = -l$ to $+l$, including zero.
4. **Spin Quantum Number (s):**
   - Describes the spin of the electron.
   - Values: $+\frac{1}{2}$ or $-\frac{1}{2}$.

Significance: These quantum numbers define the unique position, energy, orientation, and spin of every electron in an atom.

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

Ans: Comparison of Ionic and Covalent Compounds

Property	Ionic Compounds	Covalent Compoun
Formation	Transfer of electrons	Sharing of electrons
Bond Type	Electrostatic attraction	Shared pair electrons
Melting/ Boiling Points	High	Generally low
Solubility	Soluble in water	Soluble in orga solvents
Electrical Conductivity	Conduct in molten solution form	Generally p conductors

Examples:

- **Ionic: NaCl (Sodium chloride), $CaCl_2$ (Calcium chloride)**
- **Covalent: CH_4 (Methane), H_2O (Water)**

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

Ans: Coordinate Covalent Bond:

- A covalent bond where both electrons in the shared pair come from the same atom.
- Example: Formation of ammonium ion (NH_4^+), where the lone pair on nitrogen forms a bond with H^+.

Difference:

- In a normal covalent bond, each atom contributes one electron to the bond.
- In a coordinate bond, both bonding electrons are contributed by one atom.

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

Ans: Hydrogen Bond:

- A weak bond between a hydrogen atom (attached to electronegative atom like O, N, F) and another electronegative atom.

Types:

- Intermolecular hydrogen bonding: between molecules (e.g., water).
- Intramolecular hydrogen bonding: within the same molecule (e.g., ortho-nitrophenol).

Water's High Boiling Point:

Water has an abnormally high boiling point due to the presence of strong intermolecular hydrogen bonding between its molecules.

- Each water molecule (H_2O) has a highly polar covalent bond between hydrogen and oxygen.
- The oxygen atom is highly electronegative and pulls the shared electrons closer, giving it a partial negative charge (δ^-) and leaving the hydrogen atoms with a partial positive charge (δ^+).
- As a result, a strong electrostatic attraction (called hydrogen bonding) forms between the hydrogen atom of one water molecule and the oxygen atom of a neighboring water molecule.
- These hydrogen bonds are much stronger than regular van der Waals forces found in most other molecular liquids.

Because of these strong hydrogen bonds:

- A large amount of energy is needed to break these bonds for water molecules to escape into the vapor phase.
- Thus, water boils at a much higher temperature (100°C) compared to other molecules of similar molecular weight.

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

Ans: Bond Angles of H_2O and NH_3

- Both H_2O and NH_3 have central atoms that are sp^3 hybridized.
- Ideal sp^3 hybrid angle is 109.5°.
- However, the presence of lone pairs distorts this angle due to repulsion:

NH_3:

- 1 lone pair and 3 bond pairs.
- Lone pair–bond pair repulsion reduces angle to 107°.

H_2O:

- 2 lone pairs and 2 bond pairs.
- Greater lone pair–lone pair repulsion reduces angle further to 104.5°.

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

Ans: Ionization Potential:

- The energy required to remove the most loosely bound electron from an isolated gaseous atom.

First < Second Ionization Potential:

- After the removal of one electron, the atom becomes a positively charged ion.
- The remaining electrons are more strongly attracted to the nucleus, making the removal of the second electron harder.

Variation with Atomic Volume:

- Ionization potential decreases with increasing atomic volume (larger atoms have weaker attraction to outer electrons).

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

Ans: f-block Elements:

- Elements in which the last electron enters the f-orbital.
- Located at the bottom of the periodic table in two series:
    1. Lanthanides (Z = 57–71)
    2. Actinides (Z = 89–103)

Why Called Inner Transition Elements:

- Their f-electrons are added to inner (n-2) orbitals, not outermost ones.
- These elements lie between s- and d-block elements in the periodic table but are placed separately to maintain structure.