PABNA UNIVERSITY OF SCIENCE AND TECHNOLOGY

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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?

Rutherford's atomic model described the atom as having a dense, positively charged nucleus at the center, with electrons revolving around it in circular orbits, much like planets around the sun.

Defects in Rutherford's model:

1. **Instability of electrons:** According to classical electromagnetic theory, a charged particle (like an electron) moving in a circular path should emit energy in the form of electromagnetic waves. As a result, the electron would lose energy, spiral inward, and eventually fall into the nucleus. But in reality, atoms are stable.
2. **No explanation for atomic spectra:** Rutherford's model couldn’t explain the line spectra of hydrogen and other elements. If electrons could move freely in orbits, they should emit a continuous spectrum, but atoms give line spectra instead.
3. **No energy levels:** It did not describe specific energy levels or shells for the electrons, which are observed in atomic behavior.

Bohr's suggestions to overcome these defects:

1. **Electrons revolve in fixed energy levels (orbits):** Bohr proposed that electrons move in certain allowed circular paths called energy levels or shells without losing energy.
2. **Energy quantization:** An electron in a particular orbit has a fixed energy and does not radiate energy.
3. **Energy emission or absorption:** Electrons can jump from one orbit to another by absorbing or emitting a definite amount of energy (a quantum of energy).

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 are a set of four numbers that describe the position and energy of an electron in an atom. These help in locating an electron and understanding its behavior.

An electron in an atom is described by **four quantum numbers**:

1. **Principal Quantum Number (n):**
   - Indicates the main energy level or shell of the electron.
   - Higher value of n means the electron is farther from the nucleus.
   - It can be 1, 2, 3, etc.
   - Example: For n = 2, the electron is in the second shell (L-shell).
2. **Azimuthal or Angular Momentum Quantum Number (l):**
   - Determines the shape of the orbital.
   - It can have values from 0 to (n – 1).
   - For example: If n = 2, then l = 0 or 1 → s and p orbitals.
   - l = 0 (s orbital), l = 1 (p orbital), l = 2 (d orbital), l = 3 (f orbital).
3. **Magnetic Quantum Number (m):**
   - Shows the orientation of the orbital in space.
   - Its values range from –l to +l.
   - For example: If l = 1, m can be –1, 0, +1 (three orientations for p orbitals).

4. **Spin Quantum Number (s):**
   - Describes the direction of the electron's spin.
   - Can be +½ (clockwise) or –½ (anticlockwise).
   - Helps to distinguish between two electrons in the same orbital.

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

Property	**Ionic Compounds**	**Covalent Compounds**
Bond formation	Transfer of electrons	Sharing of electrons
Types of elements	Between metals and non-metals	Between non-metals only
Physical state	Generally solid and crystalline	Can be solid, liquid, or gas
Melting/Boiling points	High	Low to moderate
Electrical conductivity	Conducts in molten or aqueous solution	Usually do not conduct electricity
Solubility	Soluble in water	Soluble in organic solvents

Examples:

- **Ionic compounds:** Sodium chloride (NaCl), Magnesium oxide (MgO)
- **Covalent compounds:** Water (H_2O), Methane (CH_4).

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

A **co-ordinate covalent bond** is a type of covalent bond in which **both electrons** in the shared pair come from the **same atom**.

In a co-ordinate bond:

- One atom (the **donor**) has a **lone pair of electrons**.
- Another atom (the **acceptor**) has an **empty orbital** and needs electrons.
- The donor donates its **entire lone pair** to form a shared bond.

Difference between a normal covalent bond and a coordinate covalent bond:

Feature	Normal Covalent Bond	Coordinate Covalent Bond
Electron source	Each atom contributes one electron to the bond pair	One atom donates **both electrons** to the bond pair

Feature	Normal Covalent Bond	Coordinate Covalent Bond
Formation	Formed by mutual sharing of electrons	Formed by **donation** of a lone pair from donor to acceptor
Example	H_2, O_2, H_2O (oxygen and hydrogen share electrons)	NH_4^+, CO, SO_2 (one atom donates lone pair)
Notation	Shown by a simple dash (–)	Often shown by an **arrow (→)** from donor to acceptor
Bond strength and properties	Generally stronger and more common	Slightly weaker and less common

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

A **hydrogen bond** is a weak electrostatic attraction between a hydrogen atom, which is attached to a highly electronegative atom (like N, O, or F), and another electronegative atom.

Types of Hydrogen Bonds:

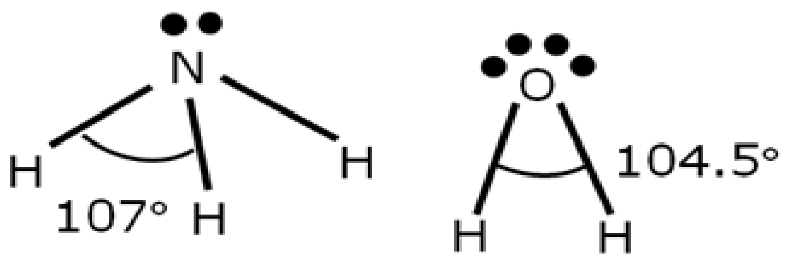
1. **Intermolecular hydrogen bonding:** Between different molecules.
   - Example: Between water molecules ($H_2O–H\cdots O–H_2O$)
2. **Intramolecular hydrogen bonding:** Within the same molecule.
   - Example: In o-nitrophenol, hydrogen bonding occurs between –OH and $–NO_2$ groups within the same molecule.

Why water has high boiling point:

Water molecules form strong intermolecular hydrogen bonds. These bonds require a large amount of energy to break, so water boils at a higher temperature compared to other molecules of similar size.

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

Both water (H_2O) and ammonia (NH_3) molecules have **sp^3 hybridization**, meaning they form four hybrid orbitals.



- In **NH_3**, one of these orbitals has a lone pair, and the remaining three are used to bond with hydrogen atoms. The lone pair causes repulsion, slightly reducing the bond angle from 109.5° to **107°**.
- In **H_2O**, there are **two lone pairs** on the oxygen atom, causing even greater repulsion. So, the bond angle reduces further to **104.5°**.

 Both H_2O and NH_3 are sp^3 hybridized and ideally tetrahedral (109.5°).

- Due to **lone pair repulsion**, the bond angles become smaller.
- The more lone pairs on the central atom, the **greater the repulsion**, and hence **smaller the bond angle**.

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 (or ionization energy) is the energy required to remove the outermost electron from an isolated gaseous atom.

- **First ionization potential:** Energy required to remove the first electron.
- **Second ionization potential:** Energy required to remove the second electron.

Why is the second greater than the first?

After removing one electron, the atom becomes a positively charged ion. The attraction between the nucleus and remaining electrons increases, so more energy is needed to remove the next electron.

Variation with atomic volume:

- **Across a period:** Ionization potential increases due to decreasing atomic size and stronger nuclear attraction.
- **Down a group:** Ionization potential decreases due to increasing atomic size and shielding effect.

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

f-Block elements are elements in which the last electron enters the **f-orbital**.

- These elements are placed separately at the bottom of the periodic table.
- They include two series:
  - **Lanthanides (4f series)**: Elements 58 to 71
  - **Actinides (5f series)**: Elements 90 to 103

f-block elements are called **inner transition elements** because:

- The **f-orbitals being filled are not the outermost shells**, but **inner energy levels**.
- For example, in lanthanides, electrons enter the **(n–2)f orbital**, which is an inner orbital.
- This is **different from main group elements**, where electrons enter the outermost shells.

So, the transition (change) of electrons is happening in **inner orbitals**, hence the name **"inner transition"**.