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Part B

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:

1. Electron Instability:

In Rutherford's model, electrons were thought to revolve in circular orbits around the nucleus. According to classical electromagnetic theory, an accelerating charged particle (like an electron) should emit electromagnetic radiation. This would cause the electron to lose energy and spiral inward towards the nucleus. In practice, however, atoms are stable and don't collapse, meaning Rutherford's model could not explain this stability.

2. Absence of Atomic Spectra:

According to Rutherford's model, electrons should continuously radiate energy in the form of electromagnetic waves as they spiral toward the nucleus. This would lead to a continuous spectrum of radiation. However, atoms emit a series of discrete lines (spectral lines) in their emission spectra. Rutherford's model couldn't account for these specific wavelengths.

3. No Explanation for Atomic Line Spectra:

The atomic line spectra observed in elements like hydrogen cannot be explained by Rutherford's model. Instead of emitting a continuous spectrum, elements emit only certain discrete lines (like the Balmer series for hydrogen), which Rutherford's model did not predict.

Bohr's Suggestions to Remove These Defects:

1. Quantized Orbits and Energy Levels:

Bohr proposed that electrons orbit the nucleus only in certain fixed orbits or energy levels, where they do not radiate energy. These orbits are quantized, meaning only specific, discrete orbits are allowed. This explained why atoms emit only specific wavelengths of light, corresponding to transitions between these allowed orbits.

2. Energy Emission and Absorption:

According to Bohr's model, electrons could absorb or emit energy only when they jump from one orbit to another. This explained the discrete spectral lines, as the energy difference between these levels corresponded to the frequency of light

emitted or absorbed.

3. Electron Stability in Specific Orbits:

Electrons in these stable orbits do not radiate energy, which resolves the issue of electron instability in Rutherford's model. Only when an electron jumps to a higher or lower orbit does it emit or absorb 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.

What is a Quantum Number?

A quantum number is a number that describes the properties of an electron in an atom. They are used to describe the position, energy, and behavior of an electron within an atom. Each electron in an atom is described by four quantum numbers.

The Four Quantum Numbers:

1. Principal Quantum Number (n):

- Definition: This number describes the main energy level or shell of an electron. It defines the electron's distance from the nucleus.
- Values: It can take any positive integer value (n = 1, 2, 3, ...).
- Significance: Higher values of n indicate that the electron is farther from the nucleus and the atom has higher energy.

2. Azimuthal Quantum Number (l):

- Definition: This number describes the shape of the orbital or sublevel. It is related to the angular momentum of the electron.
- Values: l can range from 0 to n-1, where n is the principal quantum number.
  - For n = 1, l = 0 (s orbital).
  - For n = 2, l = 0 or 1 (s or p orbitals).
  - For n = 3, l = 0, 1, or 2 (s, p, or d orbitals).
- Significance: The value of l determines the shape of the orbital (e.g., s, p, d, f).

3. Magnetic Quantum Number (m_l):

- Definition: This number specifies the orientation of the orbital in space relative to the other orbitals.
- Values: m_l can take integer values between -l and +l, including 0.
- Significance: For example, in the p orbital (l = 1), m_l can be -1, 0, or +1, which corresponds to the three possible orientations of a p orbital (px, py, pz).

4. Spin Quantum Number (m_s):

- Definition: This number describes the intrinsic angular momentum (spin) of the electron.
- Values: m_s can be either +1/2 or -1/2, representing the two possible spin orientations of an electron.
- Significance: The spin quantum number indicates the direction of the electron's spin (clockwise or counterclockwise).

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

Property	Ionic Compounds	Covalent Compounds
Bond Formation	Formed by the transfer of electrons from one atom to another, resulting in the formation of positively and negatively charged ions.	Formed by the sharing of electron pairs between atoms.
Nature of Bond	Electrostatic attraction between oppositely charged ions.	Electrons are shared between atoms.
Melting/Boiling Point	Typically high due to strong ionic bonds.	Generally lower than ionic compounds due to weaker intermolecular forces.
Solubility in Water	Soluble in water, as water molecules can surround and separate the ions.	Generally not soluble in water, but soluble in non-polar solvents.
Electrical Conductivity	Conduct electricity when molten or dissolved in water due to the presence of free ions.	Poor conductors of electricity as they do not contain free ions.
State at Room Temperature	Typically solid (e.g., NaCl, K_2SO_4).	Can be solid, liquid, or gas (e.g., H_2O, CO_2).
Hardness	Hard and brittle due to the strong	Soft and flexible (except for network covalent

Property	Ionic Compounds	Covalent Compounds
	ionic bonds.	compounds like diamonds).
Examples	Sodium chloride (NaCl), Potassium sulfate (K_2SO_4).	Water (H_2O), Carbon dioxide (CO_2).

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

Definition of Co-ordinate Covalent Bond:

A co-ordinate covalent bond (also called a dative bond) is a type of covalent bond in which both electrons shared in the bond come from the same atom. This occurs when one atom donates a pair of electrons to another atom that does not have any electrons to share in the bond.

Here is a table comparing a co-ordinate covalent bond and a normal covalent bond:

Aspect	Co-ordinate Covalent Bond	Normal Covalent Bond
Electron Donation	Both electrons in the bond are donated by the same atom.	Each atom contributes one electron to the bond.
Electron Pair Source	One atom provides both electrons for the bond.	Each atom provides one electron to form the bond.
Formation	Formed when one atom donates a lone pair to another atom with an empty orbital.	Formed when two atoms share electrons equally or unequally.
Example	NH_4^+ (Ammonium ion): NH_3 donates a lone pair to H^+.	H_2 (Hydrogen molecule): Both H atoms share one electron each.
Bonding Atom's Role	One atom acts as the donor, while the other atom is the acceptor.	Both atoms share the electrons equally or unequally.

Aspect	Co-ordinate Covalent Bond	Normal Covalent Bond
Nature of Bond	Can be formed by atoms or ions with lone pairs and empty orbitals.	Typically formed between atoms with similar electronegativity.
Bond Strength	The strength is generally similar to a normal covalent bond.	Depends on the electronegativity difference and atomic sizes.

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

What are Hydrogen Bonds?

A **hydrogen bond** is a weak to moderate attractive force that exists between a hydrogen atom covalently bonded to a highly electronegative atom (such as nitrogen, oxygen, or fluorine) and another electronegative atom with a lone pair of electrons. Although hydrogen bonds are weaker than covalent and ionic bonds, they play a significant role in determining the physical properties of compounds, especially in water and organic molecules.

Classification of Hydrogen Bonds:

1. **Intermolecular Hydrogen Bonds:**
    - These occur between molecules, for example:
        - **Water (H_2O):** Hydrogen bonding occurs between the hydrogen atom of one water molecule and the oxygen atom of another.
        - **Ammonia (NH_3):** Hydrogen bonding exists between the hydrogen atom of one ammonia molecule and the nitrogen atom of another.
2. **Intramolecular Hydrogen Bonds:**
    - These occur within a single molecule, where a hydrogen atom in one part of the molecule is attracted to an electronegative atom in another part of the same molecule.

- **Ethanol (C_2H_5OH):** The hydrogen of the hydroxyl group may form an intramolecular bond with the oxygen atom within the same molecule.
- **DNA:** Hydrogen bonds between nitrogenous bases hold the two strands of the DNA double helix together.

Explanation of Water's High Boiling Point:

Water has an **abnormally high boiling point** due to the presence of **hydrogen bonds**. Each water molecule can form up to four hydrogen bonds with surrounding molecules, creating a strong intermolecular force that requires a significant amount of energy to break. The energy required to break these hydrogen bonds is what causes water to have a high boiling point compared to other molecules of similar size. For example, while methane (CH_4), with a similar molecular weight, boils at -161.5°C, water boils at 100°C due to its hydrogen bonds.

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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 **H_2O** (water) and **NH_3** (ammonia) have **sp^3 hybridization** at their central atoms (oxygen in H_2O and nitrogen in NH_3). However, the bond angles deviate from the expected 109.5° for a perfect tetrahedral geometry. The main reason for this deviation is **lone pair repulsion**.

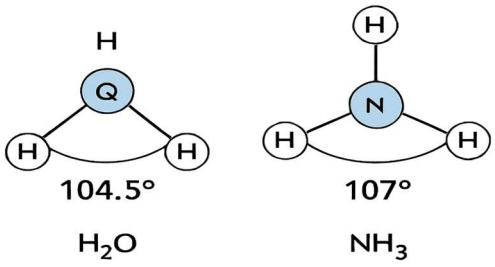
1. **H_2O (Water):**
    - **Hybridization:** Oxygen in H_2O undergoes sp^3 hybridization, resulting in four electron pairs around it — two bonding pairs (with hydrogen atoms) and two lone pairs.
    - **Bond Angle:** The ideal bond angle for sp^3 hybridization is 109.5°, but in water, the bond angle is reduced to **104.5°** due to the presence of the lone pairs. Lone pairs occupy more space than bonding pairs, exerting greater repulsive force, which compresses the bond angle between the hydrogen atoms.
2. **NH_3 (Ammonia):**
    - **Hybridization:** Nitrogen in NH_3 is also sp^3 hybridized, with three bonding pairs (with hydrogen atoms) and one lone pair.

- **Bond Angle:** The ideal bond angle for sp^3 hybridization is 109.5°, but in ammonia, the bond angle is **107°**. Again, the lone pair on nitrogen exerts a repulsive force on the bonding pairs, but since there are only three bonding pairs, the angle is less compressed than in H_2O.

In both cases, the lone pair-bond pair repulsion is greater than bond pair-bond pair repulsion, resulting in bond angles smaller than the ideal tetrahedral angle.

Hybridization Modes:

1. **sp^3 Hybridization:**
    - In both H_2O and NH_3, the central atom (oxygen in H_2O and nitrogen in NH_3) undergoes sp^3 hybridization. This means the s and three p orbitals mix to form four equivalent sp^3 hybrid orbitals, which are arranged in a tetrahedral geometry.
    - **Geometry:** Tetrahedral.
    - **Bond Angles:** Expected to be 109.5°, but deviations occur due to lone pair repulsion.
2. **sp^3 Hybridization Example:**
    - **Water (H_2O):** Oxygen undergoes sp^3 hybridization, forming two bonds with hydrogen atoms and two lone pairs of electrons. The repulsion between lone pairs causes the bond angle to be reduced to 104.5°.
    - **Ammonia (NH_3):** Nitrogen undergoes sp^3 hybridization, forming three bonds with hydrogen atoms and one lone pair of electrons. The lone pair causes the bond angle to be reduced to 107°.



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?

What is Ionization Potential?

Ionization potential (also known as ionization energy) is the amount of energy required to remove an electron from a neutral atom or molecule in the gas phase. It measures the tendency of an atom to lose an electron and form a positive ion (cation). It is usually expressed in electron volts (eV) or kilojoules per mole (kJ/mol).

- **First Ionization Potential:** The energy required to remove the first electron from a neutral atom.
- **Second Ionization Potential:** The energy required to remove a second electron from the ion that was formed by the first ionization.

Why is the First Ionization Potential Less than the Second?

The first ionization potential is always less than the second ionization potential because:

1. **Effect of Electron Removal:**
    - After the first electron is removed, the atom becomes positively charged. This increases the effective nuclear charge (the attraction between the nucleus and the remaining electrons).
    - The remaining electrons are pulled more strongly toward the nucleus, making it harder to remove the second electron.
2. **Electron Shielding:**
    - In the case of the first ionization, there is less shielding by other electrons, so the electron is relatively easier to remove.
    - After the first ionization, the effective nuclear charge increases due to the decrease in the number of electrons, which makes the second electron harder to remove.

How Does Ionization Potential Vary with Atomic Volume?

- As **atomic volume** (or atomic radius) increases, ionization potential tends to decrease. This is because the outer electrons are farther from the nucleus and experience less attraction from the protons in the nucleus, making them easier to remove.
- Conversely, when the atomic volume is smaller (i.e., the atom is smaller and the electrons are closer to the nucleus), the ionization potential increases because the electrons are held more tightly by the nucleus.

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

What are f-Block Elements?

The **f-block elements** are elements found in the periodic table that belong to the two rows at the bottom of the table: the **lanthanide series** and the **actinide series**. These elements are characterized by the filling of the **f-orbitals** with electrons.

- **Lanthanides:** These are the 14 elements following **lanthanum (La)**, from **cerium (Ce)** to **lutetium (Lu)**, filling the 4f orbitals.
- **Actinides:** These are the 14 elements following **actinium (Ac)**, from **thorium (Th)** to **lawrencium (Lr)**, filling the 5f orbitals.

Why are f-Block Elements Called Inner Transition Elements?

The f-block elements are called **inner transition elements** because they are part of the transition metals, but they are situated in the inner part of the periodic table.

1. **Transition Elements:** Transition metals are those elements whose electrons are filling the d-orbitals. These elements typically have a high atomic number.
2. **Position in Periodic Table:** The f-block elements are placed below the main body of the periodic table. This is because their electrons are added to the f-orbitals, which are inner orbitals relative to the d-orbitals. Their position "inside" the main body of the transition metals justifies their name as "inner" transition elements.

The f-block elements have unique properties, including a wide range of oxidation states and the ability to form complex compounds, due to the involvement of the f-

orbitals in bonding. The lanthanides are often associated with high magnetic susceptibility, and the actinides include radioactive elements, such as uranium and thorium.