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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?

Answer:

Rutherford's atomic model had some significant drawbacks:

(1) Rutherford’s model was unable to explain the stability of an atom. According to Rutherford’s postulate, electrons revolve at a very high speed around a nucleus of an atom in a fixed orbit. However, Maxwell explained accelerated charged particles release electromagnetic radiations. Therefore, electrons revolving around the nucleus will release electromagnetic radiation.

(2) The electromagnetic radiation will have energy from the electronic motion as a result of which the orbits will gradually shrink. Finally, the orbits will shrink and collapse in the nucleus of an atom. According to the calculations, if Maxwell’s explanation is followed Rutherford’s model will collapse with 10-8 seconds. Therefore, Rutherford atomic model was not following Maxwell’s theory and it was unable to explain an atom’s stability.

(3) Rutherford’s theory was incomplete because it did not mention anything about the arrangement of electrons in the orbit. This was one of the major drawbacks of Rutherford atomic model.

To address these issues, Niels Bohr proposed modifications:

Postulate 1: When an electron revolves in any selected orbits, it neither emits nor absorbs energy. The energy of an electron in a particular orbit is constant. These orbits are therefore, called stationary orbits. Depending on the distance of from the nucleus, these orbits are divided into energy levels such as K, L, M, N … etc. and these are designated respectively by the numbers 1, 2, 3, 4…etc.

Postulate 2: The electron in the hydrogen atom revolves around the nucleus only in certain selected circular paths (called orbit) which are associated with definite energies. Only those orbits are permitted for the revolving electron in which the angular momentum of the electron is a whole number multiple of h/2π angular momentum of electron,

$$\mathbf{mvr} = \boldsymbol{n} \times \boldsymbol{h} / 2\boldsymbol{\pi}$$

Where, n = 1, 2, 3, 4…etc.

m = mass of the electron

v = velocity of the electron

r = radius of the orbit

h = Plank’s constant

Thus the angular momentum of electrons in an atom is quantized.

Postulate 3: As long as an electron remains in a particular orbit, it neither emits (i.e., radiates or losses) nor absorbs energy. But when the electron is excited from a lower energy level to a higher energy level, it absorbs energy. On the other hand when it comes back from a higher energy level to a lower energy level, it emits energy. Now if ν the frequency of the radiation absorbed by the electron, then, according to Plank's quantum theory of radiation, energy absorbed by the electron will be

$$\Delta E = E2 - E1 = h\nu$$

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.

Answer:

Quantum numbers are a set of values that describe the unique quantum state of an electron in an atom. They define the electron's position, energy, and behavior within an orbital.

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

1. **Principal Quantum Number (n):** Determines the energy level and size of the orbital. Higher values of *n* indicate larger orbitals and higher energy levels.
2. **Azimuthal Quantum Number (l):** Defines the shape of the orbital. It ranges from 0 to *(n-1)* and corresponds to different subshells (s, p, d, f).
3. **Magnetic Quantum Number (m_l):** Specifies the orientation of the orbital in space. It can take values from *-l* to *+l*.
4. **Spin Quantum Number (m_s):** Represents the spin direction of the electron, either +½ or -½, indicating the two possible spin states.

Together, these quantum numbers provide a complete description of an electron's behavior in an atom.

6.(a)

Compare the properties of ionic and covalent compounds. Give two examples of each type of compounds.

Answer:

Here's a comparison of **ionic** and **covalent** compounds:

Properties of Ionic Compounds:

- High melting and boiling points due to strong electrostatic forces.
- Hard and brittle in solid form.
- Soluble in water but insoluble in nonpolar solvents.
- Conduct electricity when dissolved in water or molten.

Properties of Covalent Compounds:

- Lower melting and boiling points compared to ionic compounds.
- Soft or flexible in solid form.
- Insoluble in water but soluble in nonpolar solvents.
- Do not conduct electricity in water.

Examples:

- **Ionic Compounds**: Sodium chloride (**$NaCl$**), Calcium oxide (**CaO**).
- **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?

Answer:

A coordinate covalent bond, also known as a dative bond, is a type of covalent bond in which one atom donates both electrons to be shared with another atom. This occurs when an atom with a lone pair of electrons shares them with an electron-deficient atom, forming a stable bond.

Difference from a Normal Covalent Bond:

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

Feature	Normal Covalent Bond	Coordinate Covalent Bond
Formation	Occurs when two atoms share electrons equally or unequally	Forms when an atom with a lone pair donates electrons to an electron-deficient atom
Examples	Water (H_2O), Carbon dioxide (CO_2)	Ammonium ion (NH_4^+), Sulfur dioxide (SO_2)

In essence, while both bonds involve electron sharing, a coordinate covalent bond is unique because only one atom provides the shared electron pair.

7. (a)

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

Answer:

Hydrogen Bonds and Their Classification

A hydrogen bond is a type of weak electrostatic attraction between a hydrogen atom (bonded to a highly electronegative atom like oxygen, nitrogen, or fluorine) and another electronegative atom. This interaction plays a crucial role in determining the properties of many substances, including water.

Types of Hydrogen Bonds

1. **Intermolecular Hydrogen Bonding**: Occurs between molecules.
   - Example: Water (H_2O) molecules form hydrogen bonds with neighboring water molecules, leading to high cohesion.
   - Other examples: Ammonia (NH_3), Ethanol (C_2H_5OH).
2. **Intramolecular Hydrogen Bonding**: Occurs within a single molecule when hydrogen bonds form between atoms in the same molecule.
   - Example: Salicylaldehyde ($C_7H_6O_2$) has internal hydrogen bonding, affecting its chemical behavior.
   - Other examples: Ethylene glycol ($C_2H_4(OH)_2$).

Why Water Has an Abnormally High Boiling Point

Water's high boiling point is due to strong hydrogen bonding between its molecules. These bonds require significant energy to break, leading to a much higher boiling point than expected for a molecule of its size. This property is essential for life, as it allows water to remain liquid over a wide range of temperatures.

7.(b)

Why bond angles of H2O and NH3 are 104.5° and 107° respectively although central atoms are sp3 hybridized.

Answer:

The bond angles of H_2O (104.5°) and NH_3 (107°) are slightly less than the ideal 109.5° expected for sp^3 hybridization due to the influence of lone pair repulsion.

Explanation:

1. **Ammonia (NH_3) – 107° Bond Angle**:
    - The nitrogen atom in NH_3 has one lone pair and three bonded pairs.
    - Lone pairs exert greater repulsion than bonded pairs, pushing the hydrogen atoms closer together.
    - This repulsion reduces the bond angle from 109.5° to 107°.
2. **Water (H_2O) – 104.5° Bond Angle:**
    - The oxygen atom in H_2O has two lone pairs and two bonded pairs.
    - The two lone pairs exert even stronger repulsion, compressing the bond angle further.
    - This results in a bond angle of 104.5°, smaller than NH_3.

Thus, the presence and number of lone pairs significantly affect bond angles, making them deviate from the ideal tetrahedral 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?

Answer:

Ionization Potential and Its Trends

Definition

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, converting it into a positively charged ion. It is measured in electron volts (eV) or kilojoules per mole (kJ/mol).

Why the First Ionization Potential Is Less Than the Second

- The first ionization energy refers to the energy needed to remove the first electron from a neutral atom.
- The second ionization energy is the energy required to remove another electron from the resulting positively charged ion.
- Since the effective nuclear charge increases after the removal of the first electron, the attraction between the nucleus and the remaining electrons becomes stronger.
- As a result, more energy is required to remove the second electron, making the second ionization potential higher than the first.

Variation of Ionization Potential with Atomic Volume

- **Across a Period (Left to Right):** Ionization potential increases because atomic size decreases, leading to stronger nuclear attraction on valence electrons.
- **Down a Group (Top to Bottom):** Ionization potential decreases as atomic size increases, reducing the nuclear attraction on outer electrons.
- **Influence of Shielding Effect:** Inner electrons shield the outermost electrons from the nucleus, lowering ionization energy.

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

Answer:

Understanding f-Block Elements

f-block elements are those in which the last electron enters the f-orbital. They are located at the bottom of the periodic table and consist of lanthanides (4f series) and actinides (5f series).

f-block elements are called inner transition elements because:

1. Their electrons fill the (n-2)f orbitals, which are inner orbitals compared to d-block elements.
2. They exhibit gradual transition in properties across the series, similar to d-block elements but occurring within the inner orbitals.
3. They show variable oxidation states, colored ions, and unique magnetic properties due to their f-electrons.