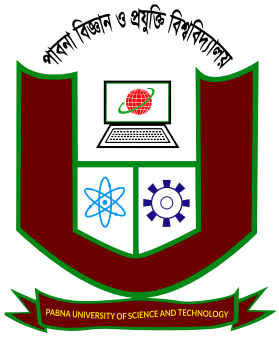

PABNA UNIVERSITY OF SCIENCE AND TECHNOLOGY

Department of Information and Communication Engineering (ICE)

Assignment

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Submitted by:	Submitted to:
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Ans to the question No:05

(a)Give the defects of Rutherford‟s model of atom. What suggestions were given by Bohr to removethese defects?

Answer:

Defects of Rutherford's Model of Atom:

1. **Instability of the Atom (Based on Classical Electromagnetic Theory):**
   According to classical physics, an electron revolving around the nucleus in a circular orbit should continuously emit electromagnetic radiation. This would cause the electron to lose energy, spiral inward, and eventually collapse into the nucleus — making the atom unstable. But atoms are actually stable.
2. **No Explanation for Atomic Spectra:**
   Rutherford's model couldn't explain the line spectra (discrete wavelengths) observed in the emission or absorption spectra of atoms, especially hydrogen. According to his model, the electron could emit radiation of any frequency, which would produce a continuous spectrum, not the observed discrete lines.

Bohr's Suggestions to Remove These Defects:

1. **Quantized Energy Levels:**
   Bohr proposed that electrons revolve in specific, fixed orbits (called energy levels or shells) around the nucleus without emitting energy. These orbits are called *stationary states*.
2. **Energy Emission/Absorption in Quantized Form:**
   An electron can move from one energy level to another by absorbing or emitting a quantum of energy ($E = h\nu$), which corresponds to the difference in energy between the two levels. This explained the discrete lines in the atomic spectra.
3. **Angular Momentum Quantization:**
   Bohr stated that the angular momentum of an electron in orbit is quantized and given by:

$$mvr=n\hbar \quad \text{where } n=1,2,3,\ldots$$

(m = mass of electron, v = velocity, r = radius of orbit, $\hbar$=h\2π)

(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:

The term "quantum number" refers to a set of numerical values that describe the unique quantum state of an electron in an atom. These numbers are used to determine the location, energy, orientation, and **spin** of an electron within an atom.

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

1. Principal Quantum Number (n)

- **Symbol**: n
- **Values**: Positive integers (1, 2, 3, …)
- **Significance**: Indicates the main energy level or shell of the electron. It also relates to the average distance of the electron from the nucleus—the higher the n, the farther the electron is.
- **Example**: n=1 is the first shell (closest to the nucleus), n=2 is the second shell, and so on.

2. Azimuthal or Angular Momentum Quantum Number (l)

- **Symbol**: l
- **Values**: Integers from 0 to n−1
- **Significance**: Defines the **shape of the orbital** (subshell). Each value corresponds to a type of orbital:
  - l=0l → **s-orbital**
  - l=1l → **p-orbital**
  - l=2l → **d-orbital**

- l=3l → **f-orbital**

3. Magnetic Quantum Number (m_l)

- **Symbol**: m
- **Values**: Integers from −l to +l including zero
- **Significance**: Specifies the orientation of the orbital in space. For example, if l=1, then m can be -1, 0, or +1, indicating three different p-orbital orientations (px, py, pz).

4. Spin Quantum Number (s)

- **Symbol**: s
- **Values**: -1/2 or + 1/2
- **Significance**: Refers to the spin direction of the electron, a quantum property. Since each orbital can hold two electrons, they must have opposite spins to obey the Pauli Exclusion Principle.

Ans to the question No:06

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

Answer:

Ionic Compounds

- Formation: Formed through the transfer of electrons from a metal to a non-metal.
- One atom loses electrons to become a positive ion (cation), and the other gains electrons to become a negative ion (anion).
- The oppositely charged ions are held together by strong electrostatic forces (ionic bonds).

Example:

- **NaCl (Sodium chloride)**
  $Na \rightarrow Na^+ + e^-$
  $Cl + e^- \rightarrow Cl^-$
  $\rightarrow Na^+$ and Cl^- attract to form NaCl.

Covalent Compounds

- Formation: Formed by the sharing of electrons between two non-metals.
- Atoms share electrons to achieve a stable electron configuration (usually octet).
- The shared pair of electrons creates a covalent bond.

Example:

- H_2O (Water)
  Each hydrogen shares one electron with oxygen → 2 single covalent bonds.

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

Answer:

A coordinate covalent bond (also called a dative bond) is a type of covalent bond in which both electrons in the shared pair come from the same atom.In a regular covalent bond, each atom donates one electron to the bond. But in a coordinate covalent bond, one atom donates both electrons to form the bond, while the other atom accepts them.

Example:

In the ammonium ion (NH_4^+), an ammonia molecule (NH_3) has a lone pair on nitrogen. This lone pair can be donated to a hydrogen ion (H^+, which has no electrons), forming a coordinate bond:

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

The bond between N and the extra H is a coordinate covalent bond.

In contrast, a normal covalent bond is formed when each atom contributes one electron to the shared pair.

Example to make it clearer:

In a normal covalent bond, like in H_2 (hydrogen gas):

- Each hydrogen atom brings one electron.
- They share these to form a bond.

In a coordinate covalent bond, like in ammonium ion (NH_4^+):

- Ammonia (NH_3) has a lone pair on nitrogen.
- A hydrogen ion (H^+) has no electrons.
- Nitrogen donates both electrons to bond with H^+.
- This forms a coordinate bond.

Even though the source of electrons is different, once the bond is formed, it behaves just like a regular covalent bond.

Ans to the question No:07

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

Answer:

A hydrogen bond is a weak electrostatic attraction that occurs when a hydrogen atom, which is covalently bonded to a highly electronegative atom (like O, N, or F), is attracted to another electronegative atom nearby.

It's not a real bond like covalent or ionic, but more like a strong intermolecular force.

Types of Hydrogen Bonds:

1. Intermolecular Hydrogen Bond
    - Occurs between different molecules.
    - Example: Water (H_2O)
        - Hydrogen from one molecule forms a hydrogen bond with oxygen of another.
        - This causes water molecules to stick together (cohesion).
2. Intramolecular Hydrogen Bond
    - Occurs within the same molecule, between two parts of the molecule.
    - Example: Ortho-nitrophenol
        - Hydrogen bonds form between the –OH group and the nearby $–NO_2$ group in the same molecule.
        - This affects boiling point and solubility.

Why Water Has an Abnormally High Boiling Point:

Water has a high boiling point compared to other group 16 hydrides (like H_2S, H_2Se), mainly because of extensive hydrogen bonding.

Here's why:

- Each water molecule can form up to 4 hydrogen bonds (2 through its lone pairs, 2 via its hydrogen atoms).
- These bonds hold water molecules strongly together.
- A lot of energy is needed to break these intermolecular forces during boiling.
- So, water boils at 100°C, much higher than similar-sized molecules.

(b)Why bond angles of H2O and NH3 are 104.5° and 107° respectively although central atoms are sp3 hybridized?

Answer:

The bond angles in molecules like H_2O and NH_3 are affected by electron pair repulsion as described by VSEPR (Valence Shell Electron Pair Repulsion) theory.

- sp^3 hybridization leads to a tetrahedral arrangement with ideal bond angles of 109.5° when there are 4 bonding pairs.
- However, lone pairs of electrons occupy more space than bonding pairs because they are localized closer to the nucleus. This results in greater repulsion from the lone pairs, which pushes the bonding pairs closer together, reducing the bond angles.
- NH_3 has 1 lone pair and 3 bonding pairs, so the bond angle is 107° (slightly less than 109.5°).
- H_2O has 2 lone pairs and 2 bonding pairs, so the bond angle is 104.5° (even smaller due to stronger lone pair repulsion).

Thus, more lone pairs = smaller bond angles.

Ans to the question No:08

(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 of an Element:

The ionization potential (or ionization energy) of an element refers to the amount of energy required to remove one mole of electrons from one mole of atoms or ions in the gas phase. Specifically, it is the energy needed to remove the most loosely bound electron from a neutral atom, converting it into a positively charged ion.

For example, the first ionization potential refers to the energy required to remove the first electron from a neutral atom ($X \rightarrow X^+ + e^-$), while the second ionization potential refers to the energy needed to remove a second electron from the positively charged ion ($X^+ \rightarrow X^{2+} + e^-$).

Why the First Ionization Potential is Less Than the Second Ionization Potential:

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

1. Electron Configuration Changes: After the first electron is removed, the resulting ion has a positive charge. The remaining electrons are now more strongly attracted to the nucleus because of the reduced electron-electron repulsion. This stronger attraction makes it harder to remove another electron, thus requiring more energy for the second ionization.
2. Effective Nuclear Charge: After the first ionization, the remaining electrons experience a stronger effective nuclear charge (Z_eff). Since the effective nuclear charge is higher after the removal of one electron, it is more difficult to remove the next electron, leading to an increase in ionization energy.
3. Electron-Electron Repulsion: In the neutral atom, electrons repel each other. After the first ionization, the repulsion is less because there are

fewer electrons, leading to a greater pull from the nucleus on the remaining electrons, which increases the ionization energy for subsequent electrons.

Ionization Potential and Atomic Volume:

The ionization potential of an element is inversely related to its atomic size or atomic volume.

- Larger Atomic Volume (Lower Ionization Potential): In larger atoms, the outermost electrons are farther from the nucleus, and the attractive force between the nucleus and these electrons is weaker. As a result, it takes less energy to remove an electron, resulting in a lower ionization potential. Hence, elements with larger atomic radii (like those in lower periods of the periodic table) tend to have lower ionization energies.
- Smaller Atomic Volume (Higher Ionization Potential): For smaller atoms (typically in the upper-right of the periodic table), the outer electrons are closer to the nucleus, and the effective nuclear charge is stronger, making the electrons harder to remove. This results in higher ionization potentials for these elements.

So, as atomic volume increases, the ionization potential tends to decrease, and as atomic volume decreases, the ionization potential tends to increase.

Summary:

- Ionization potential refers to the energy required to remove an electron from an atom.
- The first ionization potential is lower than the second because after the first electron is removed, the ion experiences a stronger attraction on the remaining electrons, making it harder to remove another electron.
- Ionization potential generally decreases as atomic volume increases, and increases as atomic volume decreases, due to the relationship between the distance of the outermost electrons from the nucleus and the effective nuclear charge experienced by those electrons.

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

Answer:

F-Block Elements:

The f-block elements are elements in the periodic table that are characterized by the filling of f-orbitals with electrons. These elements are placed in the two rows at the bottom of the periodic table:

1. Lanthanides (elements 57 to 71, from Lanthanum to Lutetium).
2. Actinides (elements 89 to 103, from Actinium to Lawrencium).

In these elements, the 4f and 5f orbitals are being filled, respectively.

Why Are F-Block Elements Called Inner Transition Elements?

The term "inner transition elements" refers to the fact that these elements transition between their inner electron shells (those that are not directly involved in the main body of the periodic table, which is the d-block). Here's why:

1. Electron Configuration: The f-block elements are filling f-orbitals, which are inner orbitals, unlike the d-block elements (transition metals) that fill d-orbitals. The f-orbitals are at a lower energy level than the d-orbitals, which is why the f-block elements are placed in two separate rows below the main periodic table.
2. Position on the Periodic Table: The f-block elements are inserted into the periodic table between the d-block elements (transition metals) to avoid making the table too wide. These elements are part of the transition series, but their filling of the f-orbitals places them in a separate block. Hence, they are called the "inner" transition elements because their electronic configuration involves the filling of inner (f) orbitals.

Key Characteristics of F-Block Elements:

- They exhibit unique chemical and physical properties due to the filling of f-orbitals.
- Many of them are lanthanides and actinides, with notable examples like Cerium (Ce), Uranium (U), and Plutonium (Pu).

- F-block elements tend to have high melting points, are often radioactive (especially the actinides), and they usually form complex compounds due to their ability to use f-electrons in bonding.

Summary:

- F-block elements are those in which the f-orbitals are being filled.
- They are called inner transition elements because they fill inner orbitals (the f-orbitals) and are placed in a separate section of the periodic table below the d-block elements.