



**Faculty of Engineering & Technology**  
***Department of Information and Communication Engineering***

**Assignment**

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**Submitted By:**

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## Question No: 1

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

**Answer:** Rutherford's model of the atom, proposed in 1911, was a major step forward in understanding atomic structure. However, it had several defects that Bohr later addressed. Below is an explanation of the defects of Rutherford's model and the suggestions given by Niels Bohr to resolve them:

### Defects of Rutherford's Model:

#### 1. Instability of the Atom:

- According to Rutherford's model, the electrons revolve around the nucleus in circular orbits.
- This should cause the electrons to constantly accelerate and emit electromagnetic radiation (energy) due to centripetal acceleration.
- As a result, the electron would lose energy and spiral inward, eventually falling into the nucleus. This would lead to the atom collapsing, which contradicts the stability of atoms observed in reality.

#### 2. Continuous Spectrum of Radiation:

- Rutherford's model predicted that the electrons would emit a continuous spectrum of radiation as they lost energy while orbiting the nucleus.
- However, experimental observations (such as the emission spectra of elements) showed that atoms emitted discrete spectral lines (specific frequencies), not a continuous spectrum. This suggested that electrons could only occupy certain energy levels, which Rutherford's model could not explain.

#### 3. Lack of Quantization:

- Rutherford's model did not include any quantization of energy levels for the electrons. It failed to explain why electrons in atoms only absorb or emit specific wavelengths of light, which is characteristic of quantized energy states in atoms.

### Bohr's Suggestions to Remove These Defects:

Niels Bohr, in 1913, proposed a new model of the atom to address these issues. His suggestions were:

#### 1. Quantized Orbits (Energy Levels):

- Bohr suggested that electrons revolve in fixed, stable orbits around the nucleus without radiating energy. These orbits correspond to discrete energy levels.
- These orbits were quantized, meaning electrons could only exist in certain allowed orbits. The electron does not lose energy while moving in these orbits.

#### 2. Energy Emission/Absorption in Discrete Quantities:

- According to Bohr, an electron could absorb or emit energy only when it jumps from one orbit (energy level) to another. The energy of the emitted or absorbed radiation was equal to the difference between the two energy levels.
- This resolved the problem of continuous radiation, explaining the discrete spectral lines observed in atomic spectra.

### 3. **Angular Momentum Quantization:**

- Bohr introduced the idea that the angular momentum of an electron in an allowed orbit is quantized. It is given by:

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

where  $\hbar$  is the reduced Planck's constant, and  $n$  is a positive integer (called the principal quantum number).

- This quantization of angular momentum was a key aspect that allowed only certain orbits to be stable.

### 4. **Stable Orbits:**

- Bohr's model proposed that electrons in stable orbits do not radiate energy. This removed the issue of the atom collapsing due to the continuous radiation predicted by Rutherford's model.

Although Bohr's model was a breakthrough, it was later refined and superseded by quantum mechanics, which offered a more complete understanding of atomic behavior. Nonetheless, Bohr's ideas were a major step in the development of atomic theory.

**(b)** What do you understand by the term, “Quantum numbers”. How many quantum numbers has an electron in an orbital? Explain the significance of each quantum number.

**Answer:** Quantum numbers are a set of numbers used to describe the properties of electrons in an atom. These numbers define the energy levels, shapes, and orientations of electron orbitals. They help in determining the exact position and energy of an electron in an atom.

An electron in an orbital is described by **four quantum numbers**. Each of these quantum numbers provides specific information about the electron's location and energy state. The four quantum numbers are:

1. **Principal Quantum Number (n)**
2. **Azimuthal Quantum Number (l)**
3. **Magnetic Quantum Number (m<sub>l</sub>)**
4. **Spin Quantum Number (m<sub>s</sub>)**

## Significance of Each Quantum Number:

### 1. Principal Quantum Number (n):

- **Range:**  $n=1,2,3,\dots n = 1, 2, 3$
- **Significance:** It defines the **energy level** or shell of the electron and determines how far the electron is from the nucleus. Higher values of  $n$  correspond to electrons that are farther from the nucleus and have higher energy.
- **Example:** For  $n=1$  the electron is in the first energy level (closest to the nucleus).

### 2. Azimuthal Quantum Number (l):

- **Range:**  $l=0$  to  $n-1$
- **Significance:** It defines the **shape** of the orbital (sublevel). The value of  $l$  determines the type of orbital
  - $l=0$  corresponds to an **s** orbital (spherical).
  - $l=1$  corresponds to a **p** orbital (dumbbell-shaped).
  - $l=2$  corresponds to a **d** orbital (double- dumbbell-shaped).
  - $l=3$  corresponds to an **f** orbital (complex shape).

### 3. Magnetic Quantum Number ( $m_l$ ):

- **Range:**  $m_l=-1$  to  $+1$  (including 0)
- **Significance:** It defines the **orientation** of the orbital in space relative to an external magnetic field. For example, for  $l=1$  (p orbitals),  $m_l$  can be  $-1$ ,  $0$ , or  $+1$ , corresponding to the three different p orbitals ( $p_x$ ,  $p_y$ ,  $p_z$ ).

### 4. Spin Quantum Number ( $m_s$ ):

- **Range:**  $m_s = +1/2$  or  $m_s=-1/2$
- **Significance:** It describes the **spin** of the electron, which can either be clockwise ( $+1/2$ ) or counterclockwise ( $-1/2$ ). This quantum number accounts for the two possible spin states of an electron.

## Question No: 2

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

**Answer:** Comparison of Properties of Ionic and Covalent Compounds

Property	Ionic Compounds	Covalent Compounds
<b>Bond Formation</b>	Formed by the transfer of electrons from one atom to another (between a metal and a non-metal).	Formed by the sharing of electrons between two non-metals.
<b>Bond Type</b>	Ionic bonds (electrostatic attraction between oppositely charged ions).	Covalent bonds (shared pair of electrons between atoms).
<b>Melting and Boiling Points</b>	Generally high due to strong ionic bonds.	Lower compared to ionic compounds due to weaker intermolecular forces.
<b>Solubility in Water</b>	Typically soluble in water (due to the formation of ions in solution).	Solubility varies, but many are not soluble in water.
<b>Electrical Conductivity</b>	Conducts electricity in molten state or when dissolved in water (due to free ions).	Does not conduct electricity (no free ions or electrons).
<b>State at Room Temperature</b>	Usually solid.	Can be solid, liquid, or gas.
<b>Hardness</b>	Generally hard and brittle.	Soft and flexible (in most cases).
<b>Examples</b>	Sodium chloride (NaCl), Magnesium oxide (MgO).	Water (H <sub>2</sub> O), Carbon dioxide (CO <sub>2</sub> ).

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

**Answer:**

**Co-ordinate Covalent Bond:**

A **co-ordinate covalent bond** (also known as a **dative bond**) is a type of covalent bond in which **both electrons** in the bond come from the **same atom**. This occurs when one atom donates a lone pair of electrons to another atom that has an empty orbital, forming a bond between the two.

- In a **normal covalent bond**, each of the two atoms contributes **one electron** to the bond, so both atoms share the electrons equally or unequally (in case of polar covalent bonds).
- In a **coordinate covalent bond**, the atom donating the lone pair provides **both electrons** to the bond, while the atom accepting the electrons has an empty orbital to accept the lone pair.

**Differences Between a Co-ordinate Covalent Bond and a Normal Covalent Bond:**

Property	Co-ordinate Covalent Bond	Normal Covalent Bond
<b>Electron Donation</b>	Both electrons in the bond come from the same atom (donor).	Each atom contributes one electron to the bond.
<b>Formation</b>	Formed when one atom donates a lone pair to an atom with an empty orbital.	Formed by the sharing of one electron from each atom.
<b>Nature of Bond</b>	The bond is still a covalent bond, but the electron pair comes from only one atom.	The bond involves the sharing of electron pairs between atoms.
<b>Examples</b>	Ammonium ion ( $\text{NH}_4^+$ ), Carbon monoxide (CO).	Hydrogen molecule ( $\text{H}_2$ ), Oxygen molecule ( $\text{O}_2$ ).

### Question No: 3

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

Answer:

#### Hydrogen Bond:

A **hydrogen bond** is a type of **weak attractive force** that exists between a **hydrogen atom** which is covalently bonded to a highly electronegative atom (such as **oxygen, nitrogen, or fluorine**) and another electronegative atom with a lone pair of electrons.

#### Types of Hydrogen Bonds:

1. **Intermolecular Hydrogen Bond (between molecules):**
  - **Example:** In **water (H<sub>2</sub>O)**, hydrogen atoms of one molecule are attracted to the oxygen atoms of another.
2. **Intramolecular Hydrogen Bond (within a molecule):**
  - **Example:** In **DNA**, hydrogen bonds hold the two strands together.

#### Why Water Has a High Boiling Point:

Water has a **high boiling point** because of its **strong hydrogen bonds**. These bonds require a lot of energy to break, so water doesn't easily vaporize, leading to its relatively high boiling point (100°C). This is much higher than similar-sized molecules like **methane (CH<sub>4</sub>)**, which don't form hydrogen bonds.

- (b) Why bond angles of H<sub>2</sub>O and NH<sub>3</sub> are 104.5° and 107° respectively although central atoms are sp<sup>3</sup> hybridized.

Answer:

The bond angles in **H<sub>2</sub>O (water)** and **NH<sub>3</sub> (ammonia)** are slightly less than the ideal bond angle of 109.5° expected for **sp<sup>3</sup> hybridized** atoms due to **lone pairs** of electrons on the central atoms. Here's why the bond angles are different for both molecules:

#### H<sub>2</sub>O (Water) - Bond Angle = 104.5°:

- **Hybridization:** The oxygen atom in H<sub>2</sub>O is **sp<sup>3</sup> hybridized**, meaning it has four electron pairs (two bonding pairs from hydrogen and two lone pairs).
- **Effect of Lone Pairs:** Lone pairs of electrons repel more strongly than bonding pairs because they are localized closer to the nucleus. The repulsion between the lone pairs pushes the

bonding pairs of electrons closer together, reducing the bond angle from the ideal  $109.5^\circ$  to  $104.5^\circ$ .

### **NH<sub>3</sub> (Ammonia) - Bond Angle = $107^\circ$ :**

- **Hybridization:** The nitrogen atom in NH<sub>3</sub> is also **sp<sup>3</sup> hybridized**, meaning it has four electron pairs (three bonding pairs from hydrogen and one lone pair).
- **Effect of Lone Pair:** While the lone pair in NH<sub>3</sub> also causes repulsion, it is not as significant as in H<sub>2</sub>O. The lone pair on nitrogen exerts some repulsion, but since there are fewer lone pairs compared to H<sub>2</sub>O, the bond angle is reduced slightly to  **$107^\circ$**  from the ideal  $109.5^\circ$ .

### **Question No: 4**

- (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:**

The **ionization potential** (also called **ionization energy**) of an element is the amount of energy required to **remove an electron** from an atom or ion in the **gaseous state**. This is typically measured in electron volts (eV) or kilojoules per mole (kJ/mol).

For example, the first ionization potential refers to the energy needed to remove the first electron from a neutral atom:



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

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

1. **After the First Electron is Removed:**
  - The removal of the first electron results in a **positively charged ion** ( $M^+$ ). The nucleus now has a stronger effective nuclear charge acting on the remaining electrons because the electron-electron repulsion is reduced.
2. **Increased Nuclear Attraction:**
  - Once the first electron is removed, the remaining electrons experience a stronger pull from the nucleus, making it harder to remove the second electron. This increased



attraction requires more energy to overcome the attractive force between the nucleus and the remaining electron, thus the **second ionization potential** is higher than the first.

## Ionization Potential and Atomic Volume:

The **ionization potential** and **atomic volume** are related as follows:

- **Inverse Relationship:**
  - As the **atomic volume** increases (meaning the atom is larger), the **ionization potential** generally **decreases**. This is because, in larger atoms, the outermost electrons are farther from the nucleus and experience less effective nuclear charge, making them easier to remove.
- **Smaller Atoms:** In smaller atoms (with smaller atomic volume), the electrons are closer to the nucleus and more tightly bound, so it takes more energy (higher ionization potential) to remove an electron.

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

**Answer:**

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

1. **Lanthanides (Rare Earth Elements):** The first row of f-block elements, with atomic numbers 57 to 71, where electrons fill the 4f orbitals.
2. **Actinides:** The second row of f-block elements, with atomic numbers 89 to 103, where electrons fill the 5f orbitals.

## Why f-block Elements Are Called Inner Transition Elements:

The f-block elements are also known as **inner transition elements** because they involve the **transition** of electrons into the **inner f-orbitals** (as opposed to the outer d-orbitals in the d-block). Here's why:

1. **Position in the Periodic Table:** These elements are placed in the **inner part** of the periodic table, below the main body of transition elements (the d-block). This placement is because their electron configurations involve the filling of **f-orbitals** that are inner shells compared to the outer d-orbitals of transition metals.
2. **Electron Configuration:** In these elements, the electrons are being added to the **f-orbitals**, which are further away from the nucleus than the **d-orbitals**, hence referred to as "inner" transition elements.