



**PABNA UNIVERSITY OF SCIENCE AND TECHNOLOGY**  
**INFORMATION AND COMMUNICATION ENGINEERING**



**Course Name : Chemistry**

**Course Code : CHE-2201**

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**session : 2021-22**

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**Submission Date: 28.04.25**



**Question 5:**

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

Rutherford's model of the atom, also known as the nuclear model, proposed that an atom consists of a tiny, dense, positively charged nucleus at the center, with electrons orbiting around it like planets around the sun. While it was a significant improvement over earlier models, it had several key defects:

**Defects of Rutherford's Model:**

1. **Failure to Explain the Stability of the Atom:** According to classical electromagnetic theory, an accelerating charged particle (like an electron orbiting a nucleus) should continuously <sup>1</sup> emit electromagnetic radiation (lose energy). As the electron loses energy, its orbit would become smaller and smaller, eventually spiraling into the nucleus. This would make the atom unstable, and matter should collapse, which is not observed. Rutherford's model couldn't explain why atoms are stable.
2. **Failure to Explain the Line Spectra of Atoms:** When elements are heated or subjected to electrical discharge, they emit light at specific wavelengths, forming a line spectrum. Rutherford's model predicted that electrons could orbit at any radius, and therefore, atoms should emit a continuous spectrum of radiation, which contradicts the observed discrete line spectra. It provided no mechanism to explain why only certain wavelengths of light were emitted.

**Suggestions by Bohr to Remove These Defects:**

Niels Bohr built upon Rutherford's nuclear model and introduced quantum concepts to address its shortcomings. His postulates were:

1. **Quantized Energy Levels:** Electrons in an atom can only occupy specific, discrete energy levels or orbits around the nucleus. These orbits are stationary states, and electrons in these orbits do not radiate energy even though they are accelerating. Each orbit is associated with a fixed amount of energy.
2. **Quantized Angular Momentum:** The angular momentum of an electron in a permitted orbit is quantized and is an integral multiple of  $2\pi h$  (where  $h$  is Planck's constant). Mathematically,  $mvr = nh/2\pi$ , where  $m$  is the mass of the

electron,  $v$  is its velocity,  $r$  is the radius of the orbit, and  $n$  is a positive integer (principal quantum number,  $n=1,2,3,\dots$ ).

3. **Energy Emission or Absorption During Transitions:** Electrons can transition between these allowed energy levels by absorbing or emitting photons of energy equal to the difference between the two levels.
  - a. When an electron jumps from a higher energy level ( $E_2$ ) to a lower energy level ( $E_1$ ), it emits a photon of energy  $h\nu = E_2 - E_1$ , where  $\nu$  is the frequency of the emitted radiation. This explained the discrete line spectra.
  - b. Conversely, an electron can absorb a photon of energy  $h\nu = E_2 - E_1$  and jump from a lower energy level ( $E_1$ ) to a higher energy level ( $E_2$ ).

By introducing these postulates, Bohr's model successfully explained the stability of the hydrogen atom (the simplest atom with one electron) and the origin of its line spectrum. The quantization of energy levels prevented the continuous loss of energy by the electron, thus resolving the stability issue. The specific energy transitions accounted for the discrete wavelengths observed in atomic spectra.

**(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.**

A quantum number is a set of numerical values that describe the properties of an electron in an atom, such as its energy, shape of its orbital, spatial orientation of the orbital, and the intrinsic angular momentum (spin) of the electron. Each electron in an atom has a unique set of four quantum numbers, as stated by the Pauli Exclusion Principle.

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

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

- a. Symbol:  $n$
- b. Possible values: Positive integers ( $n=1,2,3,\dots$ )
- c. Significance:

- i. Determines the energy level of the electron. Higher values of  $n$  correspond to higher energy levels and greater distance from the nucleus.
- ii. Also indicates the size of the orbital. Larger  $n$  values correspond to larger orbitals.
- iii. These energy levels are also referred to as shells ( $n=1$  is the K shell,  $n=2$  is the L shell,  $n=3$  is the M shell, and so on).

## **2. Azimuthal Quantum Number ( $l$ ) or Angular Momentum Quantum Number:**

- a. Symbol:  $l$
- b. Possible values: Integers ranging from 0 to  $n-1$  ( $l=0,1,2,\dots,n-1$ )
- c. Significance:
  - i. Determines the shape of the atomic orbital. Different values of  $l$  correspond to different subshells, which have characteristic shapes:
    - 1.  $l=0$ : s orbital (spherical shape)
    - 2.  $l=1$ : p orbital (dumbbell shape)
    - 3.  $l=2$ : d orbital (more complex shapes, often with four lobes)
    - 4.  $l=3$ : f orbital (even more complex shapes with multiple lobes)
  - ii. Also contributes to the angular momentum of the electron.

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

- a. Symbol:  $m_l$
- b. Possible values: Integers ranging from  $-l$  to  $+l$ , including 0 ( $m_l = -l, -l+1, \dots, 0, \dots, +l-1, +l$ )
- c. Significance:
  - i. Determines the spatial orientation of an atomic orbital in three-dimensional space relative to an applied magnetic field.
  - ii. For a given value of  $l$ , there are  $2l+1$  possible values of  $m_l$ , indicating the number of orbitals in that subshell (e.g.,  $l=1$  for p subshell has  $2(1)+1=3$  orbitals:  $p_x, p_y, p_z$ ).

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

- a. Symbol:  $m_s$
- b. Possible values:  $+1/2$  or  $-1/2$
- c. Significance:

- i. Describes the intrinsic angular momentum of the electron, which is also quantized and is called spin angular momentum.
- ii. Electrons behave as if they are spinning, creating a magnetic dipole moment. The two possible spin states are often referred to as "spin up" (+1/2) and "spin down" (-1/2).

### Question 6:

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

Ionic and covalent compounds exhibit distinct properties due to the different types of chemical bonds holding their constituent atoms together. Here's a comparison:

Property	Ionic Compounds	Covalent Compounds
<b>Bonding</b>	Electrostatic attraction between oppositely charged ions (cations and anions).	Sharing of electron pairs between atoms.
<b>Formation</b>	Typically between metals and nonmetals.	Typically between nonmetals.
<b>Physical State at Room Temperature</b>	Usually solids with high melting and boiling points.	Can be solids, liquids, or gases with generally lower melting and boiling points.
<b>Solubility in Water</b>	Many are soluble in polar solvents like water, dissociating into ions.	Many are insoluble in water; some polar covalent compounds can dissolve.
<b>Electrical Conductivity</b>	Conduct electricity when molten or dissolved in water (due to mobile ions). Do not conduct in the solid state (ions are fixed in the lattice).	Generally poor conductors of electricity (no mobile ions or free electrons). Some exceptions exist (e.g., graphite).
<b>Hardness and Brittleness</b>	Generally hard and brittle crystalline solids.	Can range from soft to hard; solids are often brittle or waxy. Liquids and gases are not applicable.

<b>Volatility</b>	Generally non-volatile (due to strong electrostatic forces).	Many are volatile (especially liquids and gases) due to weaker intermolecular forces.
<b>Nature of Bonds</b>	Strong electrostatic forces acting in all directions.	Directional bonds, leading to specific molecular shapes.
<b>Reaction Rates</b>	Reactions in aqueous solutions are typically fast (ionic reactions).	Reactions are often slower as they involve bond breaking and formation between molecules.

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### Examples:

- **Ionic Compounds:**
  - Sodium Chloride (NaCl) - Common table salt.
  - Magnesium Oxide (MgO) - Used in antacids and refractory materials.
- **Covalent Compounds:**
  - Water (H<sub>2</sub>O) - Essential for life.
  - Carbon Dioxide (CO<sub>2</sub>) - A gas produced during respiration and combustion.

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

A co-ordinate covalent bond (also known as a dative bond or Lewis acid-base adduct) is a type of covalent bond in which one atom provides both electrons of the shared pair. The other atom in the bond contributes an empty orbital to accommodate these electrons.

### Difference from a Normal Covalent Bond:

In a normal covalent bond, each of the two atoms participating in the bond contributes one electron to the shared pair. The shared electrons are then attracted to the nuclei of both atoms, resulting in the bond.

### Key Differences Summarized:

Feature	Normal Covalent Bond	Co-ordinate Covalent Bond
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<b>Electron Contribution</b>	Each atom contributes one electron to the shared pair.	One atom contributes both electrons to the shared pair; the other contributes an empty orbital.
<b>Formation</b>	Formed by the sharing of electrons from both atoms.	Formed when an atom with a lone pair of electrons donates them to an atom with an empty orbital.
<b>Properties of the Bond Once Formed</b>	Once formed, it is indistinguishable from a normal single covalent bond in terms of its properties (bond length, bond strength, etc.).	Once formed, it is also indistinguishable from a normal single covalent bond. The origin of the electrons doesn't affect the nature of the bond itself.

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### Examples of Co-ordinate Covalent Bonds:

- **Formation of the Ammonium Ion ( $\text{NH}_4^+$ ):** Ammonia ( $\text{NH}_3$ ) has a lone pair of electrons on the nitrogen atom. A proton ( $\text{H}^+$ ) has an empty 1s orbital. The nitrogen atom donates its lone pair to form a co-ordinate covalent bond with the proton:  $\text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+$ . In the ammonium ion, all four N-H bonds are equivalent once formed, even though one originated from a co-ordinate bond.
- **Formation of the Hydronium Ion ( $\text{H}_3\text{O}^+$ ):** A water molecule ( $\text{H}_2\text{O}$ ) has two lone pairs of electrons on the oxygen atom. It can donate one of these lone pairs to a proton ( $\text{H}^+$ ) with an empty 1s orbital:  $\text{H}_2\text{O} + \text{H}^+ \rightarrow \text{H}_3\text{O}^+$ . Again, all three O-H bonds in the hydronium ion become equivalent.
- **Formation of Metal Complexes:** Many transition metal ions form co-ordinate covalent bonds with ligands (molecules or ions with lone pairs of electrons), such as in the formation of  $[\text{Cu}(\text{NH}_3)_4]^{2+}$ .

In essence, the distinction between a normal covalent bond and a co-ordinate covalent bond lies in the origin of the shared electron pair during bond formation. Once the bond is formed, it behaves like an

### Question 7:

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

Hydrogen bonds are a special type of intermolecular force that occurs when a hydrogen atom bonded to a highly electronegative atom (such as oxygen, nitrogen, or fluorine) is attracted to another electronegative atom in a different molecule. This attraction arises due to the significant difference in electronegativity, which creates a partial positive charge ( $\delta^+$ ) on the hydrogen atom and a partial negative charge ( $\delta^-$ ) on the electronegative atom. The positively charged hydrogen is then attracted to the lone pair of electrons on the electronegative atom of a neighboring molecule.

Hydrogen bonds can be classified into two main types:

- **Intermolecular Hydrogen Bonds:** These occur between different molecules. Examples include:
  - Hydrogen bonds between water molecules ( $\text{H}_2\text{O} \cdots \text{H}_2\text{O}$ ).
  - Hydrogen bonds between ammonia molecules ( $\text{NH}_3 \cdots \text{NH}_3$ ).
  - Hydrogen bonds between alcohol molecules ( $\text{ROH} \cdots \text{ROH}$ ).
  - Hydrogen bonds between water and alcohol molecules ( $\text{H}_2\text{O} \cdots \text{ROH}$ ).
  - Hydrogen bonds between the hydrogen of one HF molecule and the fluorine of another ( $\text{HF} \cdots \text{HF}$ ).
- **Intramolecular Hydrogen Bonds:** These occur within the same molecule when a hydrogen atom and an electronegative atom are located in close proximity. Examples include:
  - In *o*-nitrophenol, where the hydrogen of the hydroxyl group forms a hydrogen bond with the oxygen of the nitro group.
  - In the enol form of beta-diketones.

Water ( $\text{H}_2\text{O}$ ) has an abnormally high boiling point compared to other hydrides of Group 16 elements (like  $\text{H}_2\text{S}$ ,  $\text{H}_2\text{Se}$ , and  $\text{H}_2\text{Te}$ ) due to the extensive network of



intermolecular hydrogen bonds between its molecules. Each water molecule can form up to four hydrogen bonds with neighboring water molecules (two through its hydrogen atoms and two through the lone pairs on its oxygen atom). These strong intermolecular forces require a significant amount of energy to overcome, leading to a much higher boiling point ( $100^{\circ}\text{C}$ ) than expected based on its molar mass. The other hydrides of Group 16 have much weaker dipole-dipole interactions and van der Waals forces, resulting in lower boiling points.

**(b) Why bond angles of  $\text{H}_2\text{O}$  and  $\text{NH}_3$  are  $104.5^{\circ}$  and  $107^{\circ}$  respectively although central atoms are  $\text{sp}^3$  hybridized.**

In both water ( $\text{H}_2\text{O}$ ) and ammonia ( $\text{NH}_3$ ), the central atoms (oxygen and nitrogen, respectively) undergo  $\text{sp}^3$  hybridization. This hybridization would ideally result in a tetrahedral electron geometry with bond angles of  $109.5^{\circ}$ . However, the actual bond angles are smaller due to the presence of lone pairs of electrons on the central atom.

- **Water ( $\text{H}_2\text{O}$ ):** Oxygen has six valence electrons. In  $\text{H}_2\text{O}$ , it forms two sigma bonds with two hydrogen atoms and has two lone pairs of electrons. The electron geometry is tetrahedral (four electron domains), but the molecular geometry is bent or V-shaped due to the arrangement of the atoms. The two lone pairs on the oxygen atom exert greater repulsion than the bond pairs. This stronger lone pair-lone pair repulsion and lone pair-bond pair repulsion compress the bond angle to approximately  $104.5^{\circ}$ , which is smaller than the ideal tetrahedral angle of  $109.5^{\circ}$ .
- **Ammonia ( $\text{NH}_3$ ):** Nitrogen has five valence electrons. In  $\text{NH}_3$ , it forms three sigma bonds with three hydrogen atoms and has one lone pair of electrons. The electron geometry is also tetrahedral (four electron domains), but the molecular geometry is trigonal pyramidal. The single lone pair on the nitrogen atom exerts repulsion on the bond pairs, causing the H-N-H bond angles to be compressed to approximately  $107^{\circ}$ , which is smaller than the ideal tetrahedral angle.

The difference in bond angles between water ( $104.5^{\circ}$ ) and ammonia ( $107^{\circ}$ ) can be attributed to the number of lone pairs on the central atom. Oxygen in water has two lone pairs, leading to stronger repulsive forces and a greater reduction in bond

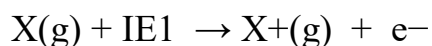
angle compared to nitrogen in ammonia, which has only one lone pair. The order of repulsion strength is: lone pair-lone pair > lone pair-bond pair > bond pair-bond pair.

## Question 8

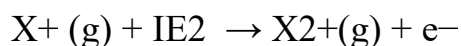
**8(a) What do you mean by the ‘ionization potential’ of an element? Why the first ionization potential of an element is less than the second ionization potential? How does the ionization potential of an element vary with atomic volume?**

The ionization potential (or ionization energy) of an element is the minimum amount of energy required to remove an electron from a gaseous atom or ion in its ground electronic state. It is usually expressed in units of kilojoules per mole (kJ/mol) or electron volts (eV).

The first ionization potential (IE<sub>1</sub>) is the energy required to remove the outermost electron from a neutral gaseous atom:



The second ionization potential (IE<sub>2</sub>) is the energy required to remove an electron from the singly charged gaseous ion:



The first ionization potential of an element is always less than the second ionization potential (IE<sub>1</sub> < IE<sub>2</sub>) because:

1. **Increased Nuclear Attraction:** When the first electron is removed, the remaining electrons are held more tightly by the nucleus because the effective nuclear charge (the net positive charge experienced by the remaining electrons) increases. There are fewer electrons to shield the full nuclear charge.
2. **Smaller Ionic Radius:** The removal of an electron generally leads to a smaller ionic radius because the remaining electrons are pulled closer to the nucleus by the increased effective nuclear charge. Removing an electron from a smaller, more tightly held electron cloud requires more energy.

3. **Removal from a More Stable Configuration:** In some cases, the removal of the first electron might result in a more stable electronic configuration (e.g., a half-filled or fully filled subshell). Removing a second electron would then disrupt this stable configuration, requiring significantly more energy.

### **Variation of Ionization Potential with Atomic Volume:**

Atomic volume is the volume occupied by one mole of an element in its solid state. It is inversely related to the effective nuclear charge and the strength of the attraction between the nucleus and the outermost electrons.

Generally, there is an inverse relationship between ionization potential and atomic volume:

- **Larger Atomic Volume (Smaller Effective Nuclear Charge):** Elements with larger atomic volumes tend to have lower ionization potentials. This is because the outermost electrons are farther from the nucleus and are shielded more effectively by the inner electrons, experiencing a weaker effective nuclear charge. Consequently, less energy is required to remove these loosely held electrons.
- **Smaller Atomic Volume (Larger Effective Nuclear Charge):** Elements with smaller atomic volumes tend to have higher ionization potentials. In these atoms, the outermost electrons are closer to the nucleus and experience a stronger effective nuclear charge. Therefore, more energy is needed to overcome the strong attraction and remove an electron.

### **Trends in the Periodic Table:**

- **Across a Period (Left to Right):** Atomic volume generally decreases due to increasing effective nuclear charge, and ionization potential generally increases.
- **Down a Group (Top to Bottom):** Atomic volume generally increases due to the addition of electron shells, and ionization potential generally decreases.

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

f-block elements are a group of elements in the periodic table where the last electron added to the atom enters the f-subshell. These elements are located in two series placed below the main body of the periodic table:

- **Lanthanide series (or Lanthanoids):** These elements follow Lanthanum (La, atomic number 57) and involve the filling of the 4f subshell (from Cerium, Ce, atomic number 58, to Lutetium, Lu, atomic number 71).
- **Actinide series (or Actinoids):** These elements follow Actinium (Ac, atomic number 89) and involve the filling of the 5f subshell (from Thorium, Th, atomic number 90, to Lawrencium, Lr, atomic number 103).

f-block elements are called inner transition elements because their position in the periodic table is considered to be within the d-block (transition elements).

Consider the placement:

- The lanthanides are inserted between Lanthanum (a d-block element in Group 3, Period 6) and Hafnium (a d-block element in Group 4, Period 6).
- The actinides are inserted between Actinium (a d-block element in Group 3, Period 7) and Rutherfordium (a d-block element in Group 4, Period 7).

The differentiating electron in these series enters the  $(n-2)f$  subshell, which is two energy levels inner than the outermost valence shell ( $n$ ) and one energy level inner than the d-subshell of the preceding transition element  $(n-1)d$ . This "inner" filling of the f-orbitals distinguishes them from the typical d-block transition elements where the differentiating electron enters the  $(n-1)d$  subshell. Due to this characteristic electronic configuration and their position within the periodic table relative to the d-block elements, they are appropriately termed inner transition elements.