

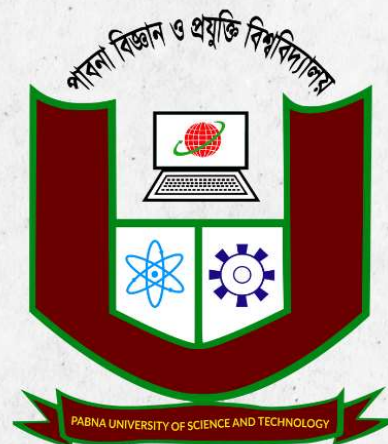
INFORMATION AND COMMUNICATION ENGINEERING

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Part B

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6(a)	Compare the properties of ionic and covalent compounds. Give two examples of each type of compounds.
6(b)	What is a co-ordinate covalent bond? How does it differ from a normal covalent bond?
7(a)	What do you understand by hydrogen bonds? Classify them with examples. Explain why water has abnormally high boiling point.
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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?

Defects of Rutherford's Atomic Model:

1. Instability of Electrons:

Rutherford's model suggested that electrons orbit the nucleus in circular paths, much like planets orbiting the Sun. However, according to classical electromagnetic theory, an accelerating electron would continuously lose energy in the form of electromagnetic radiation. This would cause the electron to spiral inward, eventually collapsing into the nucleus. Yet, this did not occur in reality, as atoms remain stable. This flaw made the Rutherford model unable to explain the stability of atoms.

2. Lack of Explanation for Atomic Spectra:

Based on Rutherford's model, electrons would emit a continuous spectrum of radiation as they lost energy while spiraling towards the nucleus. However, experimental results showed that atoms emit distinct, discrete spectral lines instead of a continuous spectrum. Rutherford's model was inadequate in explaining why atoms emit only specific wavelengths of light.

3. Failure to Account for Discrete Spectral Lines:

Rutherford's theory could not explain why atomic spectra, like those of hydrogen, consist of only specific lines (e.g., the Balmer series), rather than a smooth, continuous range of wavelengths. The atomic spectra observed in experiments presented a clear contrast to the continuous emission predicted by Rutherford's model.

Bohr's Revisions to Address the Defects:

1. Introduction of Quantized Orbits:

Bohr proposed that electrons do not occupy just any orbit around the nucleus but instead exist in specific, fixed orbits (energy levels). In these quantized orbits, electrons remain stable without radiating energy. The concept of quantized orbits explained why only certain wavelengths of light were emitted or absorbed, as electrons could only transition between specific energy levels.

2. Energy Emission During Electron Transitions:

In Bohr's model, the emission or absorption of energy happens only when an electron jumps from one orbit to another. The energy difference between the orbits corresponds

to the frequency of the emitted or absorbed light. This concept successfully explained the discrete spectral lines, as the energy gaps between the orbits are responsible for the specific wavelengths observed in atomic spectra.

3. Electron Stability in Fixed Orbits:

To resolve the issue of electron instability in Rutherford's model, Bohr suggested that electrons do not emit radiation while moving in stable orbits. It is only during transitions between these quantized orbits that electrons release or absorb energy. This allowed electrons to remain in stable orbits without spiraling inward, thereby maintaining the atom's structural integrity and stability.

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 ($n = 1, 2, 3, \dots$).
- 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 (p_x , p_y , p_z).

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
Structure	Ionic compounds have a regular, ordered, and rigid crystalline lattice structure.	Covalent compounds have distinct molecules, often arranged in molecular structures.
Melting and Boiling Points	Generally have high melting and boiling points due to the strong electrostatic forces between ions.	Typically have low melting and boiling points due to weaker intermolecular forces.
Solubility	Often soluble in water and other polar solvents.	Solubility depends on the polarity of the compound (polar covalent compounds are soluble in polar solvents).
Electrical Conductivity	Conduct electricity when molten or dissolved in water due to the movement of ions.	Do not conduct electricity in solid or liquid form as they lack free ions.

Hardness and Brittleness	Usually hard but brittle, as the ionic bonds break easily under stress.	Soft and flexible (in many cases), as covalent bonds are generally weaker.
State at Room Temperature	Most are solid at room temperature.	Can be solid, liquid, or gas at room temperature.
Examples	Sodium chloride (NaCl), Magnesium oxide (MgO).	Water (H ₂ O), Carbon dioxide (CO ₂).

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

Feature	Co-ordinate Covalent Bond	Normal Covalent Bond
Formation	Formed when one atom donates a lone pair of electrons to another atom that has an empty orbital.	Formed when two atoms share one or more pairs of electrons equally.
Example	Ammonium ion (NH ₄ ⁺), Carbon monoxide (CO), $\text{BF}_3 + \text{NH}_3 \rightarrow \text{BF}_3\text{NH}_3$.	Hydrogen (H ₂), Oxygen (O ₂), Methane (CH ₄).
Electron Donor and Acceptor	One atom (donor) provides both electrons, the other atom (acceptor) accepts them.	Each atom provides one electron to form the bond.
Bond Strength	Generally weaker than normal covalent bonds, since only one atom contributes both electrons.	Generally stronger due to the equal sharing of electrons between two atoms.
Type of Atoms Involved	Typically occurs between a Lewis base (which has a lone pair of electrons) and a Lewis acid (which has an empty orbital).	Typically occurs between two non-metals.

Nature of the Bond	A special case of covalent bonding where both electrons come from one atom.	A typical covalent bond formed by mutual sharing of electrons.
Polarity	May be polar, depending on the atoms involved.	Can be polar or non-polar depending on the electronegativity difference.
Bond Example in Molecules	In carbon monoxide (CO), carbon donates a pair of electrons to oxygen.	In H ₂ O (water), the oxygen shares electrons with two hydrogen atoms.

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

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

- **Formation:** Hydrogen bonds form when the hydrogen atom, which is bonded to a more electronegative atom (e.g., oxygen, nitrogen, or fluorine), develops a partial positive charge. This hydrogen then interacts with a lone pair of electrons on a nearby electronegative atom (such as oxygen or nitrogen), resulting in an attraction between the hydrogen and the electronegative atom.
- **Strength:** Hydrogen bonds are generally weaker than covalent bonds but stronger than most other types of intermolecular forces like van der Waals forces.

Classification of Hydrogen Bonds:

Hydrogen bonds can be classified into two types:

1. Intermolecular Hydrogen Bonds:

- **Definition:** These hydrogen bonds occur between molecules.
- **Example:** In **water (H₂O)**, the hydrogen atoms of one water molecule are attracted to the oxygen atom of another water molecule, forming intermolecular hydrogen bonds.

- **Example: Ammonia (NH_3)** also forms intermolecular hydrogen bonds, where the hydrogen atoms of one molecule are attracted to the nitrogen atoms of another.

2. Intramolecular Hydrogen Bonds:

- **Definition:** These hydrogen bonds occur within a single molecule, between different parts of the same molecule.
- **Example: In DNA,** hydrogen bonds form between the nitrogenous bases (adenine-thymine and guanine-cytosine) to hold the two strands of the DNA together.
- **Example: In salicylic acid,** hydrogen bonding occurs between the hydroxyl group ($-\text{OH}$) and the carboxyl group ($-\text{COOH}$) of the same molecule.

Why Water Has an Abnormally High Boiling Point:

Water (H_2O) has an **exceptionally high boiling point** compared to other molecules of similar size because of the presence of strong **hydrogen bonding** between water molecules. Here's how hydrogen bonding influences water's boiling point:

1. Strong Hydrogen Bonds Between Water Molecules:

- In water, each hydrogen atom of a molecule forms a hydrogen bond with the oxygen atom of a neighboring water molecule. These bonds are relatively strong intermolecular forces, which require a large amount of energy to break.

2. Energy Required to Break Bonds:

- To convert water from liquid to gas, the hydrogen bonds must be broken. This requires a significant amount of energy, which raises water's boiling point.

3. Comparison with Other Molecules:

- Other molecules, such as hydrogen sulfide (H_2S) or methane (CH_4), which are of similar molecular size to water, do not exhibit hydrogen bonding and thus have much lower boiling points. For example, hydrogen sulfide boils at -60°C , while water boils at 100°C .

4. High Surface Tension:

- The strong hydrogen bonds also contribute to water's high surface tension, further illustrating the energy required to break the intermolecular attractions.

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:

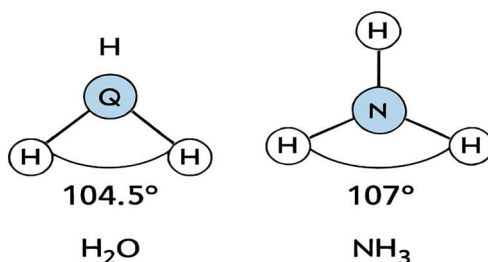
1. sp^3 Hybridization:

- 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?

f-Block Elements:

The **f-block elements** are those elements in the periodic table that have electrons filling the **f orbitals**. These elements are divided into two groups:

1. **Lanthanides:** These are elements with atomic numbers 57 to 71, from **Lanthanum (La)** to **Lutetium (Lu)**. They are also called **rare earth elements**.
2. **Actinides:** These elements have atomic numbers 89 to 103, ranging from **Actinium (Ac)** to **Lawrencium (Lr)**. Many of these elements, like **Uranium (U)** and **Plutonium (Pu)**, are radioactive.

In these elements, the **f orbitals** are progressively filled with electrons as we move from left to right across the periodic table.

Why f-Block Elements Are Called Inner Transition Elements:

The **f-block elements** are referred to as **inner transition elements** for the following reasons:

1. **Location in the Periodic Table:**

The f-block elements are placed **separately** below the main body of the periodic table. This makes them appear like a "transition" from the main body of elements, but they are positioned **inside** the table, which is why they are called "inner" transition elements.

2. **Electron Configuration:**

The defining feature of these elements is that their **f orbitals** (which are inner orbitals) are being filled with electrons. These orbitals are not as outer as the **d orbitals** (of d-block elements), and as such, they are referred to as "inner" because they are filled after the d-block elements in the periodic table.