PABNA UNIVERSITY OF SCIENCE AND TECHNOLOGY

Department of
Information and Communication Engineering

Faculty of Engineering and Technology

Course Tittle: Chemistry
Course Code: CHEM-2201

Submitted By:	**Submitted To:**
Name**: Siam Khan** Roll: 220620 Session: 2021-2022 2nd Year 2nd Semester Dept. of Information and Communication Engineering, PUST.	**Abrar Yasir Abir** Lecturer Department of Chemistry, PUST.

Answer to the Question No- 05

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

Ans:

Defects of Rutherford's Model of the Atom:

1. **Instability of Atom:**
   According to classical electromagnetic theory, a charged particle (like an electron) moving in a circular orbit around the nucleus should continuously emit energy in the form of electromagnetic radiation. As a result, the electron would spiral inward and eventually fall into the nucleus. This means atoms should not be stable — but in reality, atoms are stable.
2. **Continuous Spectrum Problem:**
   If electrons were spiraling into the nucleus while emitting radiation continuously, atoms should produce a continuous spectrum of light (all colors mixed). However, experiments showed that atoms emit light at only specific, discrete wavelengths (a line spectrum).

Bohr's Suggestions to Remove These Defects:

1. **Quantized Orbits (Energy Levels):**
   Bohr proposed that electrons can revolve only in certain special orbits (called stationary orbits or energy levels) without radiating energy. These orbits have fixed energies.
2. **Energy Emission/Absorption Rule:**
   Electrons emit or absorb energy only when they jump from one allowed orbit to another. The energy change corresponds exactly to the difference between the energy levels, explaining the discrete line spectra.

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

<u>Ans</u>:

Quantum Number:

- Quantum numbers are **numbers that describe the position and energy of an electron in an atom**.
- They specify the size, shape, orientation, and spin of the electron's orbitals.

The Four Quantum Numbers and Their Significance:

1. **Principal Quantum Number (n):**
    - It describes the **main energy level** or **shell** of the electron.
    - It also indicates the **size** of the orbital.
    - Values: n=1,2,3,4,…
    - Larger n means higher energy and larger orbit.
2. **Azimuthal Quantum Number (l):**
    - It describes the **shape** of the orbital.
    - Values: l=0,1,2,…,(n−1)l = 0, 1, 2, \dots, (n-1)l=0,1,2,…,(n−1)
        - l=0l = 0l=0 → **s-orbital** (spherical)
        - l=1l = 1l=1 → **p-orbital** (dumbbell)
        - l=2l = 2l=2 → **d-orbital** (cloverleaf)
        - l=3l = 3l=3 → **f-orbital** (complex shape)
3. **Magnetic Quantum Number (m):**
    - It describes the **orientation** of the orbital in space.
    - Values of m= -l to +l (including zero)
4. **Spin Quantum Number (s):**
    - It describes the **spin direction** of the electron.
    - Values : +1/2 , -1/2
    - Important for explaining how two electrons can occupy the same orbital (with opposite spins).

Answer to the Question No- 06

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

Ans:

Property	Ionic Compounds	Covalent Compounds
Nature of Bond	Formed by **transfer** of electrons	Formed by **sharing** of electrons
Constituents	Ions (cations and anions)	Molecules
Physical State	Generally **solid** (hard and brittle)	Usually **gases, liquids, or soft solids**
Melting and Boiling Points	**High** (due to strong electrostatic forces)	**Low** (due to weak intermolecular forces)
Solubility	Soluble in **water**, insoluble in organic solvents	Soluble in **organic solvents**, less soluble in water
Electrical Conductivity	Conduct electricity in molten or aqueous state	Generally **do not conduct electricity**
Examples	Sodium chloride ($NaCl$), Magnesium oxide (MgO)	Water (H_2O), Carbon dioxide (CO_2)

Two Examples of Each:

- **Ionic Compounds:** Sodium chloride ($NaCl$), Magnesium oxide (MgO)
- **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?

Ans:

Co-ordinate Covalent Bond:

- A **co-ordinate covalent bond** (also called a **dative bond**) is a type of covalent bond in which **both electrons** shared between two atoms come **from the same atom**.
- Example: In the ammonium ion, the nitrogen atom donates a lone pair of electrons to bond with a proton.

Difference from Normal Covalent Bond:

Aspect	Normal Covalent Bond	Co-ordinate Covalent Bond
Electron contribution	Each atom contributes **one electron** to the shared pair	**Both electrons** are donated by **one atom**
Formation	Mutual sharing of electrons	One-sided donation of a lone pair
Representation	Shown by a simple line (–)	Often shown by an arrow (→) pointing from donor to acceptor

Quick Example:

- **Normal covalent bond:** H_2 molecule (each H atom gives one electron).
- **Co-ordinate covalent bond:** NH4+ ion (N donates a lone pair to H+).

<u>**Answer to the Question No- 07**</u>

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

<u>**Ans:**</u>

Hydrogen Bond:

- A **hydrogen bond** is a **special type of weak bond** formed when a hydrogen atom, which is already bonded to a highly electronegative atom (like nitrogen, oxygen, or fluorine), gets attracted to another nearby electronegative atom.
- It is **weaker than a covalent bond**, but **stronger than most other intermolecular forces** (like Van der Waals forces

Types of Hydrogen Bonding:

1. **Intermolecular Hydrogen Bonding:**
   - Occurs **between molecules**.
   - Example: In water (H_2O), one water molecule forms hydrogen bonds with neighboring water molecules.
2. **Intramolecular Hydrogen Bonding:**
   - Occurs **within the same molecule**.
   - Example: In o-nitrophenol, a hydrogen bond forms between the –OH group and the $–NO_2$ group within the same molecule.

Why Water Has an Abnormally High Boiling Point:

- In water, each molecule forms **extensive hydrogen bonds** with neighboring molecules.
- These strong intermolecular hydrogen bonds **hold water molecules together tightly**.
- As a result, **more energy** (heat) is needed to break these bonds and turn water into vapor, leading to a **higher boiling point** compared to other similar-sized molecules (like H_2S, which has much weaker forces).

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

<u>**ANS:**</u>

Reason for Different Bond Angles in H_2O and NH_3 (even though both are sp^3 hybridized):

- **In NH_3 (Ammonia):**
    - Nitrogen is **sp^3 hybridized**, forming three N–H bonds and has **one lone pair**.
    - The lone pair-bond pair repulsion is **stronger** than bond pair-bond pair repulsion.
    - This pushes the three bond pairs closer together, **reducing the bond angle** slightly from the ideal 109.5° (for perfect sp^3) to about **107°**.
- **In H_2O (Water):**
    - Oxygen is also **sp^3 hybridized**, forming two O–H bonds and having **two lone pairs**.
    - Now, there are **two lone pairs**, and lone pair-lone pair repulsion is even **stronger**.
    - These strong repulsions push the bond pairs closer together **even more**, reducing the bond angle further to about **104.5°**.

Summary:

- **More lone pairs = greater repulsion = smaller bond angle.**
- NH_3 (one lone pair) → 107°
- H_2O (two lone pairs) → 104.5°

Answer to the Question No- 08

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?

ANS:

Ionization Potential (Ionization Energy):

- It is the **amount of energy required** to **remove the outermost electron** from a gaseous atom to form a positive ion.

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

- After removing the first electron, the atom becomes a **positively charged ion**.
- The positive charge **increases the attraction** between the nucleus and the remaining electrons.
- As a result, **more energy** is needed to remove the second electron compared to the first.

Variation of Ionization Potential with Atomic Volume:

- **Atomic volume increases** as you move **down a group** in the periodic table.
- As atomic size increases, the outer electrons are **farther from the nucleus** and experience **less electrostatic attraction**.
- Therefore, **ionization potential decreases** with **increasing atomic volume**.

In short:
Bigger atom = Lower ionization potential because it's easier to remove an electron.

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

<u>**Ans:**</u>

f-block Elements:

- The **f-block elements** are those elements in which the last electron enters into the **f-orbital** (the (n-2)f orbital).
- They include two series:
  - **Lanthanides** (elements with atomic numbers 58 to 71)
  - **Actinides** (elements with atomic numbers 90 to 103)
- In the periodic table, they are usually shown **separately at the bottom**.

Why f-block Elements are Called Inner Transition Elements:

- They are called **inner transition elements** because:
  - In these elements, the **f-orbitals are being filled**, which are **inner orbitals**, lying **deeper inside** the atom compared to outer s- and p-orbitals.
  - While d-block (transition) elements involve filling of outer (n-1)d orbitals, f-block elements involve filling of **(n-2)f orbitals**, making them "inner" transitions.

In Short:

- **f-block = filling f-orbitals**
- **Inner transition** because they involve inner shell electron transitions during their chemistry.