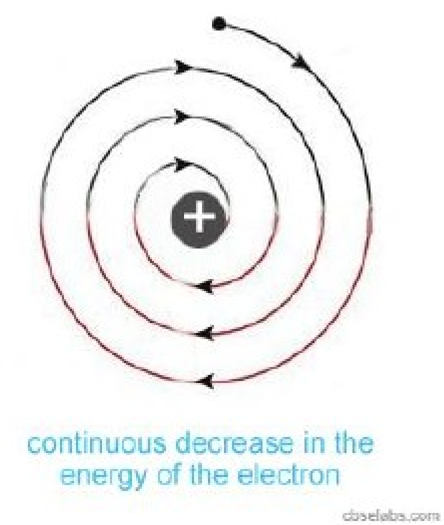
CHEMISTRY ASSINGMENT

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) Rutherford's model was unable to explain the stability of an atom. According to Rutherford's postulate, electrons revolve at a very high speed around a nucleus of an atom in a fixed orbit. However, Maxwell explained accelerated charged particles release electromagnetic radiations. Therefore, electrons revolving around the nucleus will release electromagnetic radiation.

(2) The electromagnetic radiation will have energy from the electronic motion as a result of which the orbits will gradually shrink. Finally, the orbits will shrink and collapse in the nucleus of an atom. According to the calculations, if Maxwell's explanation is followed Rutherford's model will collapse with 10-8 seconds. Therefore, Rutherford atomic model was not following Maxwell's theory and it was unable to explain an atom's stability.



(3) Rutherford's theory was incomplete because it did not mention anything about the arrangement of electrons in the orbit. This was one of the major drawbacks of Rutherford atomic model.

Bohr's Modifications (Bohr's Atomic Model - 1913):

To fix these defects, Niels Bohr proposed a new model based on quantum theory. His key suggestions were:

Postulate 1: When an electron revolves in any selected orbits, it neither emits nor absorbs energy. The energy of an electron in a particular orbit is constant. These orbits are therefore, called stationary orbits. Depending on the distance of from the nucleus, these orbits are divided into energy levels such as **K, L, M, N** … etc. and these are designated respectively by the numbers **1, 2, 3, 4**…etc.

Postulate 2: The electron in the hydrogen atom revolves around the nucleus only in certain selected circular paths (called orbit) which are associated with definite energies. Only those orbits are permitted for the revolving electron in which the angular momentum of the electron is a whole number multiple of h/2π angular momentum of electron,

$$\mathbf{mvr} = n \times h/2\pi$$

Where, **n** = 1, 2, 3, 4…etc.

m = mass of the electron

v = velocity of the electron

r = radius of the orbit

h = Plank's constant

Thus the angular momentum of electrons in an atom is quantized.

Postulate 3: As long as an electron remains in a particular orbit, it neither emits (i.e., radiates or losses) nor absorbs energy. But when the electron is excited from a lower energy level to a higher energy level, it absorbs energy. On the other hand when it comes back from a higher energy level to a lower energy level, it emits energy. Now if ν the frequency of the radiation absorbed by the electron, then, according to Plank's quantum theory of radiation, energy absorbed by the electron will be

$$\Delta E = E2 - E1 = h\nu$$

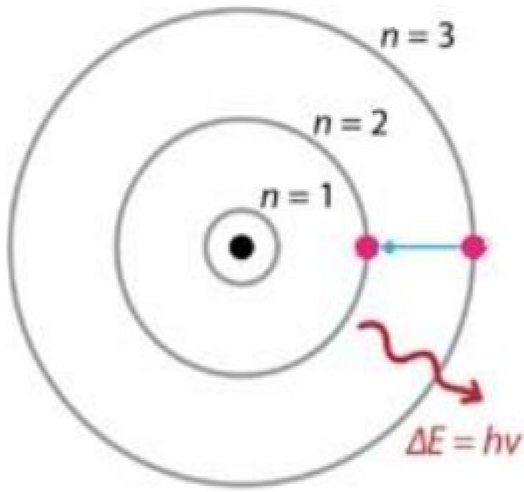


Figure: Bohr's Atom Model

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.

Quantum numbers: Certain identification numbers which are necessary to give complete information about any revolving electron in an atom are quantum numbers.

There are four quantum numbers used to describe the exact position and the nature of an electron.

They are:

Principal Quantum Number (n):

Different energy levels in an atom according to Bohr theory have been represented by n which can have non-zero positive integer values upto infinity i.e.,

$$n = 1, 2, 3, 4 \ldots \infty$$

Thus **n** can have infinite values and it has been called principal quantum number.

Significance: Determines the size and energy of the orbital. A larger n means the electron is farther from the nucleus and has more energy.

Azimuthal Quantum Number (l):

In order to explain the formation of fine structure, **Sommerfeld** proposed the existence of elliptical orbits, besides the circular orbits of Bohr. In order to specify the shape of the elliptical orbit, another supplementary quantum number is necessary. This supplementary quantum number which indicates the ellipticity of the electronic orbit (sub-shell) is called azimuthal or subsidiary quantum number denoted by '**l**'. i.e., l **= '0' → (n-1)**

Values of l	→	**0**	**1**	**2**	**3**
Designation of sub-shell	→	**s**	**p**	**d**	**f**

Significance: Defines the **shape** of the orbital.

- **l**=**0** → **s**-orbital (spherical)
- **l**=**1** → **p**-orbital (dumbbell-shaped)
- **l**=**2** → **d**-orbital
- **l**=**3** → **f**-orbital

Magnetic Quantum Number (m or ml):

Due to the angular motion of electron around the nucleus, a magnetic field is produced which interacts with the external magnetic field. As a result sub-shell of definite energy split into three dimensional spatial regions which can be expressed by magnetic quantum number. It is denoted

by '**m**'. i.e., **m = -l → +l**

Significance: Determines the number of orbitals and their orientation within a subshell.

Spin Quantum Number (s or ms):

It is observed that the electron in an atom is not only revolved around the nucleus but is also spinning around its own axis like the earth. As this quantum number describes the spin motion of the electron. It is designated by '**s**'. Since the electron can spin only clockwise and therefore the spin quantum number can take only two values be **+1/2** or **-1/2**.

Significance: Accounts for the two possible spin directions of an electron. It explains phenomena like **Pauli Exclusion Principle**, where no two electrons in the same atom can have the same set of all four quantum numbers.

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

Comparison Between Ionic and Covalent Compounds:

Property	Ionic Compounds	Covalent Compounds
Nature of Bonding	Formed by **transfer** of electrons	Formed by **sharing** of electrons
Constituent Particles	Made up of **ions** (cations and anions)	Made up of **molecules**
Physical State	Usually **solid** at room temperature	Can be **solid, liquid, or gas**
Melting and Boiling Points	Generally **high**	Generally **low**
Electrical Conductivity	Conducts electricity in **molten state or solution**	Poor conductors of electricity
Solubility	Soluble in **water** (polar solvents)	Soluble in **organic solvents** (like benzene)
Bond Strength	Strong **electrostatic** force of attraction	Weaker **intermolecular** forces compared to ionic bonds

Examples of Ionic Compounds:

1. **Sodium Chloride (NaCl):**
   - **Sodium (Na)** donates one electron to **chlorine (Cl)**.
   - This forms **Na^+ (cation)** and **Cl^- (anion)**.
   - The opposite charges attract, forming a **strong ionic bond**.
2. **Calcium Oxide (CaO):**
   - **Calcium (Ca)** loses two electrons to form **Ca^{2+}**.
   - **Oxygen (O)** gains two electrons to form **O^{2-}**.
   - The electrostatic attraction between Ca^{2+} and O^{2-} makes this an **ionic compound**.

Examples of Covalent Compounds:

1. **Water (H_2O):**
   - Each **hydrogen (H)** shares one electron with **oxygen (O)**.

- Oxygen shares one electron with each hydrogen, forming **two covalent bonds**.
- Electrons are shared, not transferred, so it's a **covalent compound**.

2. **Carbon Dioxide (CO_2):**
    - **Carbon (C)** shares **two pairs of electrons** with each **oxygen (O)** atom.
    - This forms **double covalent bonds** (C=O).
    - Since electrons are shared, it is a **covalent compound**.

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

A **co-ordinate covalent bond** (also called a **dative bond**) is a type of covalent bond in which **both electrons** in the shared pair come from **one atom only**.

- It still involves sharing of electrons, like a regular covalent bond.
- Once formed, it behaves just like a normal covalent bond.

How It Forms:

- One atom has a **lone pair of electrons** (a pair not involved in bonding).
- Another atom has an **empty orbital** and needs electrons to complete its octet.
- The atom with the lone pair **donates** both electrons to form the bond.

Example:

Ammonium ion (NH_4^+):

- Ammonia (NH_3) has a lone pair on nitrogen.
- A proton (H^+) has no electrons.
- Nitrogen donates its lone pair to bond with H^+, forming NH_4^+.
- The bond between nitrogen and the fourth hydrogen is a **co-ordinate covalent bond**.

Difference Between Co-ordinate and Normal Covalent Bond:

Property	Normal Covalent Bond	Co-ordinate Covalent Bond
Electron Contribution	Each atom donates **one electron**	One atom donates **both electrons**
Bond Formation	Mutual sharing between two atoms	One-sided sharing (donor → acceptor)
Example	H_2, O_2, CH_4	NH_4^+, CO (carbon monoxide), SO_2
Initial Electron Ownership	Electrons come from **both atoms**	Electrons come from **one atom only**

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 weak electrostatic force of attraction between:

- A **hydrogen atom** that is **covalently bonded to a highly electronegative atom** (like N, O, or F), and
- Another **electronegative atom** (also usually N, O, or F) of a different molecule or a different part of the same molecule.

Hydrogen bonding is **stronger than van der Waals forces** but **weaker than ionic or covalent bonds**.

Classification of Hydrogen Bonds:

1. **Intermolecular Hydrogen Bond:**
    - Occurs **between molecules**.
    - **Example:** In water (H_2O), hydrogen bonds form between the hydrogen of one water molecule and the oxygen of another.
2. **Intramolecular Hydrogen Bond:**
    - Occurs **within the same molecule**, between different parts of it.
    - **Example:** In **ortho-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:

- Water molecules form **strong intermolecular hydrogen bonds**.
- These bonds **hold the molecules tightly together**, making it **harder to separate them** by heating.
- As a result, **more energy (higher temperature)** is needed to break these bonds and convert water from liquid to gas.
- This is why water has a **boiling point of 100°C**, which is **much higher** than other group 16 hydrides like H_2S or H_2Se.

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

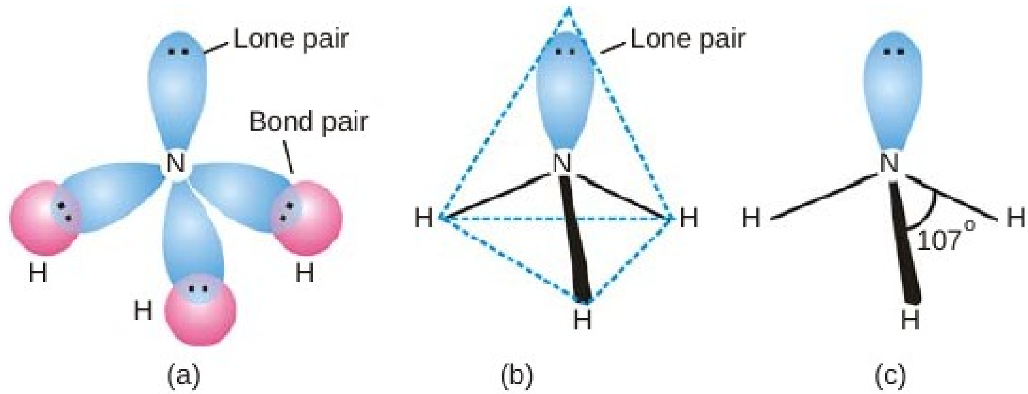
Both **water (H_2O)** and **ammonia (NH_3)** have a **central atom** (Oxygen in H_2O, Nitrogen in NH_3) that is **sp^3 hybridized**, meaning:

- The central atom forms a **tetrahedral arrangement** with four electron pairs (bonding or non-bonding).

But their **bond angles differ** due to the **number of lone pairs** and the **repulsion between electron pairs**:

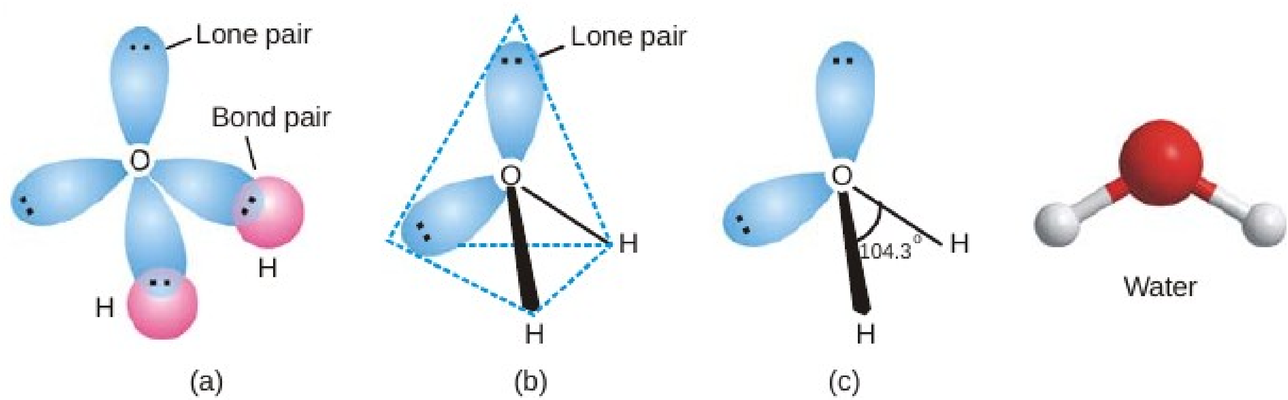
In NH_3 (Ammonia):

- Central atom: **Nitrogen (N)**
- Electron pairs: 3 **bond pairs** (N-H) + 1 **lone pair**
- Ideal tetrahedral angle: **109.5°**
- The lone pair-bond pair repulsion **pushes the bond pairs closer**, reducing the angle slightly.
- **Observed bond angle: 107°**



In H_2O (Water):

- Central atom: **Oxygen (O)**
- Electron pairs: 2 **bond pairs** (O-H) + 2 **lone pairs**
- The two lone pairs cause **even more repulsion**, compressing the bond angle further.
- **Observed bond angle: 104.5°**



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 (or ionization energy) is the **amount of energy required** to remove the **most loosely bound electron** from an **isolated gaseous atom** to form a **cation**.

- For example:

  $Na(g) \rightarrow Na^{+}(g) + e-$

  The energy required for this reaction is the **first ionization potential** of sodium.

Why is the First Ionization Potential Less Than the Second?

- The **first ionization potential** involves removing an electron from a **neutral atom**.
- After removing the first electron, the atom becomes a **positively charged ion**.
- In the **second ionization**, you're trying to remove an electron from a **positively charged ion**, where the remaining electrons are held more tightly by the nucleus.
- Hence, **more energy** is needed to remove the second electron.

1st I.P.<2nd I.P.<3rd I.P.…

Variation of Ionization Potential with Atomic Volume:

- **As atomic volume increases**, the **outer electrons are farther from the nucleus** and are **less tightly held**.
- This means **less energy** is required to remove them.
- So, **ionization potential decreases** with **increasing atomic volume**.

In the Periodic Table:

- **Across a period (left to right):**
    - Atomic size decreases, so ionization potential **increases**.
- **Down a group (top to bottom):**
    - Atomic size increases, so ionization potential **decreases**.

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

What Are f-Block Elements?

The f-block elements are a group of metals in the periodic table where electrons fill the f-orbitals (4f or 5f). They include:

•Lanthanides (4f series, Atomic numbers 58–71)

•Example: Cerium (Ce), Europium (Eu)

•Actinides (5f series, Atomic numbers 90–103)

•Example: Uranium (U), Plutonium (Pu)

Why Are They Called Inner Transition Elements?

1.**Position in the Periodic Table**

- They are placed below the main periodic table (in two separate rows) to maintain its compactness.
- They "transition" between Group 2 (alkaline earth metals) and Group 3 (transition metals).

2.**Electron Filling in Inner f-Orbitals**

- Unlike d-block (transition metals), where electrons fill outer d-orbitals, f-block elements fill inner 4f or 5f orbitals.
- Example:

Lanthanum (La): [Xe] $5d^1$ $6s^2$ → Next element (Ce) fills 4f instead of 5d.

3.**Chemical Similarity**

- Due to similar +3 oxidation states and close atomic sizes, they exhibit near-identical chemistry, making them a distinct "inner" series.