



Department of

INFORMATION AND COMMUNICATION ENGINEERING Assignment

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

Rutherford proposed that:

- The atom consists of a small, dense, positively charged nucleus.
- Electrons revolve around the nucleus in circular paths (like planets around the sun).

However, this model had **two major flaws**:

1. Violation of Classical Electromagnetic Theory:

- According to **Maxwell's electromagnetic theory**, a charged particle (like an electron) moving in a circular path should continuously emit electromagnetic radiation.
 - As it emits radiation, the electron should **lose energy** and **spiral inward** toward the nucleus.
 - This means atoms should be **highly unstable** and **collapse within a very short time** (estimated lifetime: 10^{-8} seconds).
 - **Contradiction:** Atoms are actually **stable** in nature. Matter around us doesn't collapse. So Rutherford's model failed to explain atomic stability.
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2. Inability to Explain Atomic Spectra:

- When gases like hydrogen are excited, they emit **line spectra** — sharp, distinct lines at specific wavelengths.
- Rutherford's model could not explain:
 - Why only **certain frequencies** (lines) of light are emitted.
 - Why each element has a **unique spectrum**.
- His model predicted a **continuous spectrum** due to continuous energy loss, which is **not observed**.

2- 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 number that describes the **properties and behavior of an electron** in an atom.

They arise from the **quantum mechanical model of the atom**, particularly from solving the Schrödinger equation for electrons.

These numbers help us understand:

- **Where** an electron is likely to be (its position in an atom)
 - **How** it behaves (its energy, shape of its orbital, orientation, and spin)
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How Many Quantum Numbers Does an Electron Have?

An electron in an atom is described by **four quantum numbers**:

1. **Principal quantum number (n)**
2. **Azimuthal (angular momentum) quantum number (ℓ)**
3. **Magnetic quantum number (m_l)**
4. **Spin quantum number (m_s)**

Each electron in an atom has a unique set of these four quantum numbers (based on the **Pauli exclusion principle**).

Explanation and Significance of Each Quantum Number:

◆ **1. Principal Quantum Number (n)**

- **Symbol:** n

- **Values:** $n=1,2,3,4,\dots$, $n = 1, 2, 3, 4, \dots$ (positive integers)
- **Significance:**
 - Determines the **main energy level** or **shell** of the electron.
 - Indicates the **size** of the orbital: higher n means farther from the nucleus.
 - Determines the **energy** of the orbital (approximate).

 **Example:**

If $n=1$, it's the **first energy level** (K shell);

If $n=2$, it's the **second energy level** (L shell), and so on.

◆ **2. Azimuthal Quantum Number (ℓ)**

- **Symbol:** ℓ
- **Values:** $\ell=0,1,2,\dots,(n-1)$, $\ell = 0, 1, 2, \dots, (n-1)$
- **Significance:**
 - Defines the **shape of the orbital** (subshell).
 - Also helps determine the **energy difference** between subshells (especially in multi-electron atoms).

 **Subshell Types:**

ℓ value Subshell Orbital Shape

0	s	Spherical
1	p	Dumbbell
2	d	Cloverleaf
3	f	Complex

 **Example:**

If $n=3$, ℓ can be 0 (3s), 1 (3p), or 2 (3d).

◆ **3. Magnetic Quantum Number (m_l)**

- **Symbol:** m_l
- **Values:** $m_l = -l, \dots, 0, \dots, +l$
- **Significance:**
 - Describes the **orientation of the orbital** in 3D space.
 - Tells us **how many orbitals** are in a given subshell.

Example: If $l=1$, then $m_l = -1, 0, +1$ (p orbital).
 → That means **3 p orbitals** (p_x, p_y, p_z).

◆ 4. Spin Quantum Number (m_s)

- **Symbol:** m_s
- **Values:** $+1/2$ or $-1/2$
- **Significance:**
 - Represents the **spin direction** of the electron.
 - Every orbital can hold **2 electrons with opposite spins**.
 - Basis of the **Pauli Exclusion Principle**: no two electrons in an atom can have the same set of all 4 quantum numbers.

Example:

If two electrons occupy the same orbital, one must have $m_s = +1/2$ and the other $m_s = -1/2$.

Summary Table:

Quantum Number	Symbol	Values	Significance
Principal	n	1, 2, 3, ...	Main energy level; size and energy
Azimuthal	ℓ	0 to $n-1$	Subshell type; orbital shape
Magnetic	m_l	$-\ell \text{ to } +\ell$	Orientation of orbital

Quantum Number	Symbol	Values	Significance
Spin	msm_sms	+1/2, -1/2	Spin direction of the electron

3- Compare the properties of ionic and covalent compounds. Give two examples of each type of compounds.

Comparison of Properties: Ionic vs Covalent Compounds

Property	Ionic Compounds	Covalent Compounds
1. Bond Formation	Formed by transfer of electrons from metal to non-metal	Formed by sharing of electrons between non-metal atoms
2. Types of Elements Involved	Metal + Non-metal	Non-metal + Non-metal
3. Physical State	Generally solid at room temperature	Can be solid, liquid, or gas
4. Melting and Boiling Points	High (due to strong electrostatic forces)	Low (due to weak intermolecular forces)
5. Solubility	Soluble in water , insoluble in organic solvents	Soluble in organic solvents , less soluble in water
6. Electrical Conductivity	Conducts electricity when molten or dissolved in water	Generally non-conductors , except some polar compounds
7. Bond Strength	Strong bonds due to ionic attraction	Weaker individual bonds, but stable molecules
8. Structure	Forms crystalline lattice structures	Exists as discrete molecules

Examples:

Ionic Compounds:

1. **Sodium chloride (NaCl)** – table salt
2. **Magnesium oxide (MgO)**

Covalent Compounds:

1. **Water (H₂O)**
2. **Carbon dioxide (CO₂)**

4-What is a co-ordinate covalent bond? How does it differ from a normal covalent bond?

A **coordinate covalent bond** (also called a **dative bond**) is a type of **covalent bond** in which **both electrons** in the shared pair come **from the same atom**.

Definition:

“A coordinate covalent bond is formed when one atom donates a lone pair of electrons to another atom or ion which is electron-deficient.”

How It Works:

- One atom has a **lone pair** (non-bonding pair of electrons).
- The other atom has an **empty orbital** and needs electrons.
- The atom with the lone pair **donates both electrons** to form a bond.

Once formed, the coordinate bond is **indistinguishable** from a normal covalent bond in terms of bond strength and length.

Example: Formation of Ammonium Ion (NH₄⁺)

- **NH₃ (Ammonia)** has a lone pair on nitrogen.

- H^+ (**Proton**) has no electrons.
- Nitrogen donates its lone pair to the H^+ , forming an **NH_4^+ ion**.

Coordinate bond:



Arrow notation (\rightarrow) is used to indicate the direction of electron donation.

Difference Between Coordinate Covalent Bond and Normal Covalent Bond:

Property	Normal Covalent Bond	Coordinate Covalent Bond
Electron Contribution	Each atom contributes one electron	One atom contributes both electrons
Nature of Bond	Mutual sharing	Donor-acceptor sharing
Notation	Represented by a single line (-)	Represented by an arrow (\rightarrow) from donor to acceptor
Example	H_2 , O_2 , H_2O , CH_4	NH_4^+ , H_3O^+ , $\text{BF}_3 \cdot \text{NH}_3$

Common Compounds with Coordinate Bonds:

- **Ammonium ion (NH_4^+)**
- **Hydronium ion (H_3O^+)**
- **Carbon monoxide (CO)**
- **Complexes like $[\text{Cu}(\text{NH}_3)_4]^{2+}$**

5- 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 electrostatic attraction** that occurs between:

- A hydrogen atom **covalently bonded** to a highly **electronegative atom** (like **N, O, or F**), **and**
- Another **electronegative atom** (with a lone pair of electrons) in the **same or a different molecule**.

Definition:

“A hydrogen bond is the attractive force between a hydrogen atom attached to a highly electronegative atom and another electronegative atom with a lone pair.”

How It Works:

- The bond between H and an electronegative atom (like O–H or N–H) is **polar**, making H partially positive (δ^+).
 - This δ^+ hydrogen is attracted to the lone pair on a nearby electronegative atom (like another O or N).
 - The result is a **hydrogen bond** (weaker than covalent bonds but stronger than van der Waals forces).
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Types of Hydrogen Bonds:

1. Intermolecular Hydrogen Bonding

- Between **two different molecules**.
- Responsible for many physical properties (e.g., boiling points, solubility).

Examples:

- **Water (H₂O)**: Each molecule forms hydrogen bonds with 3–4 neighboring molecules.
- **Hydrogen fluoride (HF)**
- **Alcohols (e.g., ethanol)**

2. Intramolecular Hydrogen Bonding

- Within the **same molecule** (between different parts of the molecule).

- Occurs when hydrogen is between two functional groups within a single molecule.

Examples:

- **Ortho-nitrophenol** ($-\text{OH}$ and $-\text{NO}_2$ groups form a bond)
 - **Salicylaldehyde**
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Why Does Water Have an Abnormally High Boiling Point?

Water's boiling point (100°C) is **unusually high** for a molecule of its size. Let's explore **why**:

Reason: Extensive Hydrogen Bonding

- Each water molecule can form **up to 4 hydrogen bonds**:
 - 2 through its lone pairs (on O)
 - 2 through its hydrogen atoms
- This leads to a **strong network** of intermolecular hydrogen bonds.

These hydrogen bonds:

- Hold the water molecules **tightly together**
- Require **a lot of energy** to break when water is boiled

So, despite its **low molecular mass (18 g/mol)**, water has a **much higher boiling point** compared to similar-sized molecules like:

- **H_2S (boiling point $\sim -60^\circ\text{C}$)** → No hydrogen bonding
 - **NH_3 (boiling point $\sim -33^\circ\text{C}$)** → Fewer hydrogen bonds
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Summary:

Aspect	Details
Hydrogen Bond	Attraction between H (attached to N, O, or F) and another electronegative atom
Types	Intermolecular (e.g., H_2O , HF), Intramolecular (e.g., o-nitrophenol)

Aspect	Details
Importance in Water	Creates a strong network that requires high energy to break → high boiling point

6- Why bond angles of H₂O and NH₃ are 104.5° and 107° respectively although central atoms are *sp*³ hybridized.

1. Hybridization of Central Atom:

- Both water (H₂O) and ammonia (NH₃) have central atoms (O and N) that are **sp³ hybridized**.
- In ideal sp³ hybridization (like in CH₄), the bond angle is **109.5°** — forming a perfect **tetrahedral geometry**.

But H₂O and NH₃ don't maintain this ideal angle. Why?

2. Role of Lone Pairs:

Molecule	Central Atom	Total Electron Pairs	Bonding Pairs	Lone Pairs
NH ₃	Nitrogen	4	3	1
H ₂ O	Oxygen	4	2	2

- NH₃ has **1 lone pair** and **3 bond pairs** → **trigonal pyramidal shape**
 - H₂O has **2 lone pairs** and **2 bond pairs** → **bent or V-shape**
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3. Why Lone Pairs Reduce Bond Angles:

- **Lone pairs occupy more space** than bonding pairs because they're not shared between atoms.
- They exert **greater repulsive force**, pushing bonding pairs **closer together**.

The more lone pairs, the more **compression** of bond angles:

Repulsion order:

Lone pair–lone pair > lone pair–bond pair > bond pair–bond pair

Actual Bond Angles:

Molecule	Theoretical Angle (sp ³)	Actual Bond Angle	Reason
CH ₄	109.5°	109.5°	No lone pairs → perfect tetrahedral
NH ₃	109.5°	107°	1 lone pair pushes bond pairs slightly closer
H ₂ O	109.5°	104.5°	2 lone pairs → more repulsion → greater compression

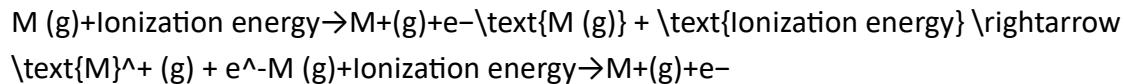
Summary:

- Both H₂O and NH₃ are sp³ hybridized.
- **Lone pairs distort bond angles** from the ideal tetrahedral angle.
- **H₂O has more lone pairs** (2 vs 1), so its bond angle is **smaller** than NH₃'s.

7- 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?

Definition:

Ionization potential (or ionization energy) is the **amount of energy required to remove the most loosely bound electron** (usually from the outermost shell) from a **neutral gaseous atom** to form a **positively charged ion**.



- It is measured in **kJ/mol** or **eV**.
 - The process must be done in the **gaseous state** to avoid interference from intermolecular forces.
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Why is the First Ionization Potential Less Than the Second?

Definitions:

- **First Ionization Potential (I_1):** Energy required to remove the first electron.
- **Second Ionization Potential (I_2):** Energy required to remove a second electron from the already positively charged ion.

Reason for the Increase:

- After removing one electron, the atom becomes a **positive ion (M^+)**.
- This ion has **greater effective nuclear charge** acting on the remaining electrons.
- The **electrons are held more tightly, so more energy is needed** to remove the second one.

So:

$$I_2 > I_1$$

This trend continues:

$$I_3 > I_2 > I_1$$

Exception Note:

If the second electron has to be removed from a **new inner shell** (after a stable configuration), the ionization energy jumps **sharply**.

How Does Ionization Potential Vary with Atomic Volume?

Atomic volume refers to **atomic size** or the **distance between the nucleus and outermost electrons**.

Trend:

- As atomic volume increases, the outer electrons are farther from the nucleus.
- The electrostatic attraction between the nucleus and these electrons decreases.
- Hence, less energy is required to remove an outer electron.

Conclusion:

$\text{Ionization potential} \propto \frac{1}{\text{Atomic volume}}$ $\text{Ionization potential} \propto \frac{1}{\text{Atomic volume}}$

In the Periodic Table:

- **Across a period:**
Atomic size **decreases**, so ionization potential **increases**.
 - **Down a group:**
Atomic size **increases**, so ionization potential **decreases**.
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Summary:

Concept	Explanation
Ionization potential	Energy needed to remove an electron from a gaseous atom
$I_1 < I_2$	Because the remaining electrons are more tightly held

Relation with atomic volume Larger atomic volume \rightarrow lower ionization potential

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

f-block elements are the elements in the periodic table that are placed in the **two rows at the bottom** (often separated from the main table). These elements have their **outermost electrons** filling the **f-orbitals**.

- The f-block is made up of two series:
 1. **Lanthanides (Rare Earth Elements):** These elements have their electrons filling the 4f orbitals. They range from **La (Lanthanum)** to **Lu (Lutetium)**.
 2. **Actinides:** These elements have their electrons filling the 5f orbitals. They range from **Ac (Actinium)** to **Lr (Lawrencium)**.

General Electronic Configuration:

- For **lanthanides**: $(n-2)f^{1-14} (n-1)s^2 \backslash \text{text}\{(n-2)\}f^{\{1-14\}} \backslash, (n-1)s^2(n-2)f^{1-14}(n-1)s^2$
 - For **actinides**: $(n-1)f^{1-14} ns^2 \backslash \text{text}\{(n-1)\}f^{\{1-14\}} \backslash, ns^2(n-1)f^{1-14}ns^2$
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Why are f-block Elements Called Inner Transition Elements?

The term "inner transition elements" is used for f-block elements because:

1. Electron Configuration:

The f-block elements have their **f-orbitals** being filled **after the d-orbitals**, and these **f-orbitals** are situated in the inner part of the atom. Thus, they are "inner" compared to the **transition metals** that fill the d-orbitals.

2. Transition Between s- and d-block:

The f-block elements transition from the **s-block** elements (like the alkali and alkaline earth metals) to the **d-block** transition metals. This is why they are also referred to as "transition" elements.

3. Position in the Periodic Table:

These elements are placed **below** the main body of the periodic table, forming two separate rows, making them the "inner" transition elements in terms of their placement.

Summary:

Aspect	f-block Elements
Definition	Elements with electrons filling the f-orbitals.
Location	Below the main periodic table (separate two rows).
Lanthanides	Fill the 4f orbitals (La to Lu).

Aspect	f-block Elements
Actinides	Fill the 5f orbitals (Ac to Lr).
Why "Inner Transition"?	Their electrons are filling inner orbitals (f-orbitals), between the s- and d-block.