



**Department of
Information and Communication Engineering**

ASSIGNMENT

CHEM-2201
Chemistry

Submitted By

Azmira Akter Simla

Roll : 220626

Dept. of Information and
Communication Engineering

Submitted To

Abrar Yasir Abir

Lecturer

Department of Chemistry
Pabna University of Science and
Technology

Q-(5.a): Give the defects of Rutherford's model of atom. What suggestions were given by Bohr to remove these defects?

Answer: Defects of Rutherford's Model of Atom:

1. Could not explain atomic stability:

According to classical physics, an electron moving around the nucleus should continuously lose energy as radiation. Eventually, it would spiral into the nucleus, and the atom would collapse. But in reality, atoms are stable, which Rutherford's model could not explain.

2. Failed to explain discrete line spectra:

Experiments showed that atoms emit light only at certain specific wavelengths (giving line spectra), not a continuous range. Rutherford's model, with electrons moving freely, should have produced continuous spectra, but it didn't match observations.

3. No idea of energy levels:

Rutherford did not propose that electrons stay in specific orbits with fixed energies. Without fixed orbits, electrons would not have stable, regular motions.

4. No explanation for chemical properties:

The model could not explain why different elements show different chemical behaviors (for example, why hydrogen reacts differently than oxygen).

5. Could not explain electron arrangement:

Rutherford said electrons revolve around the nucleus but did not explain how many electrons are there, how they are distributed, or how they move systematically.

6. Did not consider quantum theory:

Rutherford's model was fully based on classical physics, but later it was found that atomic behavior follows quantum rules (like quantization of energy), which his model ignored.

Bohr's Suggestions to Remove Rutherford's Defects:

Electrons move in fixed orbits (energy levels):

Bohr suggested that electrons revolve around the nucleus only in certain allowed circular paths called orbits or energy levels, without radiating energy.

No energy loss in stable orbits:

As long as the electron stays in a fixed orbit, it does not lose energy. That's why atoms are stable and do not collapse.

Energy is absorbed or emitted during transitions:

Electrons can jump from one orbit to another.

- When an electron jumps to a higher orbit, it absorbs energy.
- When it falls to a lower orbit, it emits energy.

The energy difference between the two orbits appears as light of a particular wavelength — explaining the line spectra.

Formula for energy absorbed or emitted:

$$\Delta E = h \nu$$

where:

- ΔE = energy difference between two orbits
- h = Planck's constant
- ν = frequency of emitted or absorbed light

Quantization of energy:

Bohr said that the angular momentum (motion property) of an electron is quantized. It can have only specific values, not any random value.

(This idea brought quantum theory into the atomic model.)

Formula for angular momentum quantization:

$$mvr = n \frac{h}{2\pi}$$

where:

- m = mass of the electron
- v = speed of the electron
- r = radius of the orbit
- n = orbit number (1, 2, 3, ...)
- h = Planck's constant

Formula for energy levels:

Bohr also gave a formula for the energy of each orbit, which correctly explained the spectra of hydrogen and some other atoms.

Formula for energy of n-th orbit:

$$E_n = - \frac{13.6 \text{ eV}}{n^2}$$

where:

- E_n = energy of electron in the n-th orbit
- n = orbit number ($n = 1, 2, 3, \dots$)

Q-(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.

Answer:

A quantum number is a set of numbers used to describe the energy state and location of an electron in an atom. It helps define various properties of electrons, such as their energy, angular momentum, and position within orbitals.

There are four quantum numbers that describe the behavior and position of an electron in an orbital.

The Four Quantum Numbers and Their Significance:

1. Principal Quantum Number (n):

Significance: It determines the energy level or shell where the electron is located. The larger the value of n, the higher the energy and the larger the orbital.

Values: $n=1,2,3,\dots$ $n = 1, 2, 3, n=1,2,3,\dots$ (positive integers)

Formula for energy:

For a hydrogen atom, the energy of an electron in the n-th orbit is given by:

$$E_n = - \frac{13.6 \text{ eV}}{n^2}$$

where:

- E_n = energy of the electron in the n-th orbit
- n = principal quantum number

2. Azimuthal Quantum Number (ℓ):

Significance: It defines the shape of the orbital. This quantum number is also related to the orbital angular momentum of the electron.

Values: $\ell = 0, 1, 2, \dots, (n-1)$

For each value of n, can take any integer value from 0 to n-1.

- ✓ $\ell = 0$ corresponds to an s orbital (spherical shape)
- ✓ $\ell = 1$ corresponds to a p orbital (dumbbell shape)
- ✓ $\ell = 2$ corresponds to a d orbital (clover shape)
- ✓ $\ell = 3$ corresponds to an f orbital (complex shape)

3. Magnetic Quantum Number (m):

Significance: It defines the orientation of the orbital in space. This quantum number specifies the direction in which the orbital is aligned in a magnetic field.

Values: $m_l = -l, -l+1, \dots, +l$

For each value of l , m can take integer values from $-l$ to $+l$, including 0. For example, if $l=1$ (p orbital), m can be $-1, 0, +1$ corresponding to the three possible orientations of p orbitals.

4. Spin Quantum Number (s):

Significance: It describes the spin of the electron. Electrons have a property called spin, which can be in one of two directions: "up" or "down".

Values: $s = +\frac{1}{2}$ or $s = -\frac{1}{2}$

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

Answer:

Property	Ionic Compounds	Covalent Compounds
Formation	Formed by the transfer of electrons from one atom to another, creating ions (positive and negative).	Formed by the sharing of electrons between atoms.
Bond Type	Ionic bonds (electrostatic attraction between oppositely charged ions).	Covalent bonds (shared pair of electrons).
Melting and Boiling Points	High melting and boiling points due to strong	Low to moderate melting and boiling points (due to

	electrostatic forces between ions.	weaker intermolecular forces).
Solubility	Usually soluble in water but insoluble in non-polar solvents.	Generally soluble in non-polar solvents, but insoluble in water.
Electrical Conductivity	Conducts electricity when molten or dissolved in water (because ions are free to move).	Does not conduct electricity (no free ions or electrons in solid or liquid form).
State at Room Temperature	Solid (forms crystalline solids).	Can be solid, liquid, or gas.
Example Compounds	NaCl (Sodium chloride), MgO (Magnesium oxide).	H ₂ O (Water), CO ₂ (Carbon dioxide).

Example of Ionic compounds:

1. NaCl (Sodium Chloride):

Sodium (Na) has one electron in its outer shell, while chlorine (Cl) has seven. Sodium gives up its one electron to chlorine, forming Na⁺ and Cl⁻ ions. The opposite charges of Na⁺ and Cl⁻ create an ionic bond, forming a stable compound, NaCl.

2. MgO (Magnesium Oxide):

Magnesium (Mg) has two electrons in its outer shell, while oxygen (O) has six. Magnesium loses its two electrons to oxygen, forming Mg²⁺ and O²⁻ ions. The electrostatic attraction between these ions holds them together in an ionic bond, forming MgO.

Example of Covalent compounds:

1. H₂O (Water):

In water, each hydrogen (H) atom shares one electron with the oxygen (O) atom. The oxygen atom shares two of its electrons, one with each hydrogen atom, to form two covalent bonds. This gives each hydrogen atom a stable electron configuration (like helium), and the oxygen atom gets a stable octet (like neon).

2. CO₂ (Carbon Dioxide):

Carbon (C) shares four electrons with two oxygen (O) atoms. Each oxygen atom shares two of its electrons with carbon, forming two double bonds. This sharing results in a stable configuration for all atoms involved.

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

Answer: A coordinate covalent bond (also known as a dative bond) is a type of covalent bond where both electrons involved in the bond come from the same atom. This occurs when one atom donates a lone pair of electrons to another atom that does not have a complete electron pair to share.

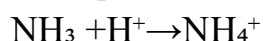
Formation of Coordinate Covalent Bond:

- Donor atom: The atom that donates the lone pair of electrons.
- Acceptor atom: The atom that accepts the lone pair of electrons.

For example:

In the formation of ammonium ion (NH₄⁺), the nitrogen atom (which has a lone pair of electrons) donates its lone pair to bond with a hydrogen ion (H⁺), forming a coordinate covalent bond.

Example:



In this reaction, nitrogen donates its lone pair to bond with the H⁺ ion.

Differences Between Coordinate Covalent Bond and Normal Covalent Bond:

Property	Coordinate Covalent Bond	Normal Covalent Bond
Electron Donation	Both electrons in the bond come from one atom (the donor).	Each atom donates one electron to the bond, so both atoms share the electrons equally.
Electron Pair Contribution	Only one atom contributes a pair of electrons.	Both atoms contribute one electron each, forming a pair.
Example of Bond Formation	$\text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+$	$\text{H} + \text{Cl} \rightarrow \text{HCl}$
Bond Type	This bond is sometimes referred to as a dative bond because the electron pair is donated.	This is simply called a covalent bond, where electrons are shared.
Nature of Bonding	Formed when an atom donates a lone pair of electrons to form a bond with another atom that doesn't have a complete pair.	Formed when atoms share electrons to achieve a full outer shell (octet rule).
Formation of Complex Ions	Often seen in the formation of complex ions, where a central atom bonds with ligands via coordinate covalent bonds (e.g., $[\text{Fe}(\text{CO})_6]^{2-}$)	Rarely involved in complex ion formation; typically seen in simpler molecules.
Bond Strength	The strength of the bond is similar to that of a normal covalent bond, but it depends on the donor-acceptor pair.	Normal covalent bonds can be strong or weak, depending on the elements involved.
Electron Distribution	In coordinate covalent bonds, the electrons are still shared, but	In normal covalent bonds, electrons are shared equally

	the donor atom retains both electrons.	or unevenly, depending on electronegativity differences.
Examples of Molecules	NH_4^+ , CO_2 (when reacting with a base), $[\text{Fe}(\text{CO})_6]^{2-}$	H_2O , CO ,
Appearance in Polar Molecules	Often found in polar molecules or ions where the donor atom has lone pairs of electrons.	Normal covalent bonds can also form polar molecules (e.g., H_2O), but they are typically neutral.
Impact on Molecular Geometry	The donor atom typically has a lone pair, which can influence the geometry of the molecule.	The molecular geometry is determined by the number of electron pairs involved in bonding (VSEPR theory).

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

Answer:

Hydrogen Bonds: A hydrogen bond is a special type of dipole-dipole interaction that occurs between molecules when a hydrogen atom that is covalently bonded to an electronegative atom (like oxygen, nitrogen, or fluorine) is attracted to another electronegative atom with a lone pair of electrons. These bonds are generally weaker than covalent bonds but stronger than van der Waals forces (dispersion forces).

There are Two types of hydrogen bonds. They are-

1. **Intermolecular Hydrogen bonding:** This type of hydrogen bonding is formed between two different molecules of the same or different substances e.g. hydrogen bonding in HF , H_2O , NH_3 etc, the substances with intermolecular hydrogen bonding have high melting points, boiling points, viscosity, surface tension etc.

2. **Intramolecular Hydrogen bonding:** This type of hydrogen bonding is formed between the hydrogen atom and the electronegative atom present within the same molecule. Intramolecular hydrogen bonding has no effect on physical properties like melting point, boiling point, viscosity, surface tension, solubility etc. For example, intramolecular hydrogen bonding exists in o-nitrophenol, 2-nitrobenzoic acid etc.

Reason for which water has abnormally high boiling point:

Water's abnormally high boiling point is largely due to the presence of hydrogen bonding between its molecules. Each water molecule can form up to four hydrogen bonds with surrounding molecules, creating a strong network that holds the molecules tightly together. This bonding requires a significant amount of energy to break, which is why water has a boiling point of 100°C at 1 atmosphere of pressure, much higher than other molecules of similar size. For example, methane (CH_4), a much smaller molecule, has a boiling point of -161°C , as it lacks hydrogen bonding. Therefore, the strong hydrogen bonds in water increase the energy needed to transition it from a liquid to a gas, resulting in its relatively high boiling point.

Q-(7.b): Why bond angles of H_2O and NH_3 are 104.5° and 107° respectively although central atoms are sp^3 hybridized?

Answer: The bond angles of H_2O (water) and NH_3 (ammonia) are 104.5° and 107° , respectively, despite both having sp^3 hybridization because of the presence of lone pairs of electrons on the central atoms (oxygen in water and nitrogen in ammonia). These lone pairs exert a greater repulsive force on the bonding pairs of electrons, which results in a reduction in the bond angles compared to the ideal tetrahedral angle of 109.5° that would be expected if there were no lone pairs.

In H_2O , the central oxygen atom is sp^3 hybridized, meaning it forms four hybrid orbitals. However, two of these orbitals contain lone pairs, and the remaining two are involved in bonding with hydrogen atoms. The lone pairs of electrons occupy more space than the bonding pairs because lone pairs are localized closer to the

nucleus, whereas bonding pairs are shared between atoms. As a result, the two bonding pairs in H_2O are pushed closer together, reducing the bond angle to approximately 104.5° .

Similarly, in NH_3 , the nitrogen atom is also sp^3 hybridized, with one lone pair of electrons and three bonding pairs with hydrogen atoms. The lone pair exerts repulsion on the bonding pairs, but because there is only one lone pair in NH_3 (compared to two in H_2O), the repulsion is less, and the bond angle is slightly larger, around 107° . This is still smaller than the ideal tetrahedral angle of 109.5° , but it is larger than the bond angle in water due to the fewer number of lone pairs.

In summary, the bond angles in H_2O and NH_3 are smaller than the ideal tetrahedral bond angle (109.5°) because the lone pairs take up more space and cause electron-pair repulsion, but the difference in bond angles between water and ammonia is due to the different numbers of lone pairs present on the central atoms.

Q-(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?

Answer:

Ionization Potential (Ionization Energy): The ionization potential (or ionization energy) of an element refers to the amount of energy required to remove an electron from an isolated gaseous atom in its ground state. In simpler terms, it is the energy needed to overcome the attractive force between the electron and the nucleus of the atom, allowing the electron to be removed.

- **First Ionization Energy:** The energy required to remove the first electron from a neutral atom.
- **Second Ionization Energy:** The energy required to remove the second electron from the ion after the first electron has already been removed.

Reason for why First Ionization Potential Less Than the Second:

The first ionization potential is generally less than the second ionization potential due to the following reasons:

1. Electron Shielding:

- ❖ In a neutral atom, the outermost electron is shielded by the inner electrons, making it easier for the electron to be removed. The nucleus's attractive force is partly counteracted by the repulsion from the inner electrons.
- ❖ After the removal of the first electron, the remaining electrons experience less shielding, meaning they feel a stronger attraction from the nucleus. This makes it harder to remove the second electron.

2. Effective Nuclear Charge:

- ❖ After the first electron is removed, the number of electrons decreases, but the number of protons in the nucleus stays the same. This increases the effective nuclear charge (the net positive charge experienced by the remaining electrons), resulting in a stronger pull on the remaining electrons.
- ❖ Consequently, the second electron is more tightly bound to the nucleus and requires more energy to be removed.

3. Electron Configuration:

- ❖ After the first ionization, the atom becomes an ion with a **more stable electron configuration**, especially if the removal results in a noble gas configuration or a stable ion. The second electron, being in a more stable configuration, is harder to remove.

Thus, it takes more energy to remove the second electron than the first one because the electron is more tightly held by the nucleus after the first ionization.

Ionization Potential and Atomic Volume:

The ionization potential of an element is inversely related to its atomic volume (or atomic size). As the atomic volume increases, the ionization potential generally decreases, and this relationship can be explained as follows:

1. Atomic Volume and Distance from the Nucleus:

- ❖ In larger atoms, the outermost electrons are farther from the nucleus. This distance reduces the attractive force between the nucleus and the outer electrons, making it easier to remove an electron (lower ionization energy).
- ❖ As the atomic volume increases, the outermost electron experiences less attraction from the nucleus, and the ionization potential becomes lower.

2. Electron Shielding:

- ❖ In atoms with larger atomic volumes, the inner electrons shield the outermost electrons more effectively. This shielding reduces the effective nuclear charge experienced by the outer electrons, which lowers the ionization potential.

3. Trends Across the Periodic Table:

- ❖ As you move **down a group** in the periodic table, the atomic volume increases due to the addition of electron shells. Since the outer electrons are farther from the nucleus, ionization potential decreases.
- ❖ As you move **across a period** (from left to right), atomic volume decreases, leading to a higher ionization potential because the outer electrons are closer to the nucleus and experience a stronger pull.

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

Answer:

f-Block Elements: The f-block elements are those elements in the periodic table whose electrons fill the f-orbitals. These elements are located in the two rows at the

bottom of the periodic table, known as the lanthanide series (elements 57 to 71) and the actinide series (elements 89 to 103). The f-orbitals are a set of seven orbitals that can hold up to 14 electrons.

- ❖ Lanthanides (elements 57 to 71): These are also known as the rare earth elements.
- ❖ Actinides (elements 89 to 103): These include elements like uranium (U) and thorium (Th), and many of them are radioactive.

Reason for why are f-Block Elements Called Inner Transition Elements:

The f-block elements are called inner transition elements because they involve the transition of electrons into the f-orbitals, which are inner orbitals, as opposed to the outer orbitals involved in the s and p blocks.

Here's why they are referred to as "inner transition elements":

1. **Electron Configuration:** The f-block elements transition between filling the 4f and 5f orbitals (lanthanides and actinides, respectively), which are deeper in the electron configuration. These orbitals are in the inner part of the atom, not the outermost shell.
2. **Position in the Periodic Table:** While these elements are placed at the bottom of the periodic table, they are technically part of the same periods as the d-block transition elements. The d-block elements fill the d-orbitals, whereas the f-block elements fill the f-orbitals.

The term "transition" reflects the change in the filling of orbitals as you move across a period. For the f-block elements, this transition occurs within the inner part of the atom (hence the term inner transition).

3. **Lanthanoids and Actinoids:** The lanthanide and actinide series are sometimes referred to as "inner transition metals" because their electron configurations involve filling f-orbitals, which are part of the atom's inner shell. This is in contrast to the d-block transition elements (elements 3 to 12) that fill the d-orbitals in the outer shells.

