



Faculty of Engineering & Technology

Department of Information and Communication Engineering

Course name : Chemistry

Course Code : CHEM -2201

Assignment On Chemistry

Submitted By:

Name: Arifur Rahman

Roll: 220612

Session: 2021-22

2nd Year 2nd Semester

Department of ICE, PUST

Submitted To:

Abrar Yasir Abir

Lecturer

Department of Chemistry,

Pabna University of Science and

Technology, Pabna

Date of Submission : 27/4/25

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

Defects of Rutherford's model of atom:

1. Instability of the atom according to classical theory:

- a. Rutherford's model proposed that electrons revolve around the nucleus in circular orbits.
- b. However, classical electromagnetic theory states that a charged particle (like an electron) moving in a circular path emits radiation continuously.
- c. This loss of energy would cause the electron to spiral into the nucleus, leading to the collapse of the atom. But in reality, atoms are stable.

2. Inability to explain atomic spectra:

- a. If electrons could revolve in any orbit around the nucleus, they should emit a continuous spectrum of radiation.
- b. However, experimental observations show that atoms, such as hydrogen, emit line spectra (distinct wavelengths or frequencies).
- c. Rutherford's model failed to explain why only certain lines appear in the atomic spectrum.

Suggestions given by bohr to remove these defects:

1. Postulate of stationary orbits:

- a. Bohr proposed that electrons move in specific, fixed orbits called stationary states or energy levels, where they do not emit radiation.
- b. These orbits correspond to specific energy levels denoted by the principal quantum number $n = 1, 2, 3, \dots$

2. Quantization of angular momentum:

- a. Bohr stated that the angular momentum of an electron is quantized and given by the formula:

$$mvr = \frac{nh}{2\pi}$$

where m is the mass of the electron, v its velocity, r the radius of the orbit, h Planck's constant, and n a positive integer.

3. Energy absorption and emission:

- a. Electrons absorb or emit energy only when they move from one allowed orbit (energy level) to another.
- b. The energy difference between the two orbits is emitted or absorbed as electromagnetic radiation:

$$\Delta E = E_2 - E_1 = h\nu$$

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.

Definition of quantum number:

Quantum numbers are a set of numerical values that describe the unique quantum state of an electron in an atom. These numbers define the energy, shape, orientation, and spin of an electron's orbital.

An electron in an orbital is described by four quantum numbers:

1. Principal Quantum Number (n):

- a. Indicates the main energy level or shell of the electron.
- b. Values: $n = 1, 2, 3, \dots$
- c. Significance: Determines the size and energy of the orbital. Higher n means larger orbit and higher energy.

2. Azimuthal (or Angular Momentum) Quantum Number (l):

- a. Describes the subshell or shape of the orbital.
- b. Values: $l = 0 \text{ to } n - 1$
- c. For example:
 - i. $l = 0$: s orbital (spherical)
 - ii. $l = 1$: p orbital (dumbbell-shaped)
 - iii. $l = 2$: d orbital (clover-shaped)

iv. $l = 3$: f orbital (complex shape)

- d. Significance: Determines the orbital's shape and contributes to the energy difference within the same shell.

3. Magnetic Quantum Number (m):

- a. Indicates the orientation of the orbital in space.
- b. Values: $m = -l$ to $+l$
- c. For $l = 1$, possible values are $m = -1, 0, +1$
- d. Significance: Tells us the number of orbitals and their orientation in a magnetic field.

4. Spin Quantum Number (s):

- a. Describes the spin direction of the electron.
- b. Values: $+\frac{1}{2}$ or $-\frac{1}{2}$
- c. Significance: Accounts for the two possible spin states of an electron.
No two electrons in the same orbital can have the same set of all four quantum numbers (Pauli Exclusion Principle).

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
Formation	Formed by transfer of electrons from metal to non-metal.	Formed by sharing of electrons between two non-metals.
Constituent particles	Ions (cations and anions).	Neutral atoms or molecules.
Bond strength	Strong electrostatic attraction between oppositely charged ions.	Generally weaker compared to ionic bonds.
Melting and boiling points	Usually high due to strong forces.	Generally low to moderate due to weaker intermolecular forces.
Physical state	Usually crystalline solids at room temperature.	Can be gases, liquids, or soft solids.
Solubility	Soluble in water (polar solvents), insoluble in organic solvents.	Soluble in organic solvents, usually insoluble in water.
Electrical conductivity	Conduct electricity in molten or aqueous state (due to free ions).	Poor conductors of electricity (no free ions).

Examples:

- Ionic Compounds:

- Sodium chloride (NaCl)
 - Magnesium oxide (MgO)
-
- Covalent Compounds:
 - Carbon dioxide (CO₂)
 - Water (H₂O)

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

Definition of coordinate covalent bond:

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

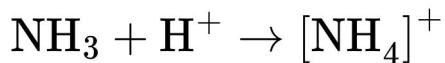
This bond is typically formed when:

- One atom with a lone pair of electrons donates both electrons to an electron-deficient atom or ion.

Example:

Formation of ammonium ion (NH₄⁺):

- Ammonia (NH₃) has a lone pair on nitrogen.
- A hydrogen ion (H⁺) has no electrons.
- The lone pair from nitrogen is donated to H⁺, forming NH₄⁺.



Difference between Coordinate Covalent Bond and Normal Covalent Bond:

Feature	Coordinate Covalent Bond	Normal Covalent Bond
Electron contribution	Both shared electrons are donated by one atom.	Each atom contributes one electron to the shared pair.
Formation condition	One atom must have a lone pair, and the other must be electron-deficient.	Formed between atoms with incomplete outer shells.
Nature of bond	A special case of covalent bond.	Standard bond involving mutual sharing.

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

Definition of hydrogen bond:

A hydrogen bond is a type of weak electrostatic attraction that occurs between:

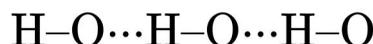
- A hydrogen atom covalently bonded to a highly electronegative atom (such as F, O, or N), and
- Another electronegative atom (often with a lone pair of electrons).

Though weaker than covalent or ionic bonds, hydrogen bonds significantly affect physical properties such as boiling point, melting point, and solubility.

Types of hydrogen bonding:

1. Intermolecular Hydrogen Bonding:

- a. Occurs between different molecules.
- b. Example: In water (H_2O), hydrogen atoms of one molecule are attracted to the oxygen atoms of neighboring water molecules.



2. Intramolecular Hydrogen Bonding:

- a. Occurs within the same molecule, often forming a ring-like structure.
- b. Example: In ortho-nitrophenol, a hydrogen bond forms between the -OH group and the $-\text{NO}_2$ group within the same molecule.

Why water has abnormally high boiling point:

- Water molecules form extensive hydrogen bonding with each other.
- These hydrogen bonds create strong intermolecular attractions that require a large amount of energy to break.
- As a result, water has an unusually high boiling point (100°C) compared to other molecules of similar molecular weight.
- For example, hydrogen sulfide (H_2S), which lacks hydrogen bonding, boils at -60°C , even though it has a similar structure.

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

Explanation:

Both water (H_2O) and ammonia (NH_3) have central atoms that are sp^3 hybridized, which ideally gives a tetrahedral geometry with a bond angle of 109.5° . However, in reality:

- NH_3 (Ammonia):
 - Has three bond pairs and one lone pair.
 - Lone pairs repel more strongly than bond pairs.
 - The increased repulsion pushes the bond pairs closer, reducing the bond angle from 109.5° to 107° .
- H_2O (Water):
 - Has two bond pairs and two lone pairs.
 - The two lone pairs exert even greater repulsion on the bond pairs.
 - This causes the bond angle to reduce further to 104.5° .

Conclusion:

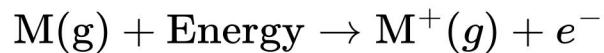
- The greater the number of lone pairs, the greater the repulsion, and hence the smaller the bond angle.

- So even though both molecules are based on a tetrahedral arrangement (sp^3 hybridization), the actual shape and bond angle are affected by the presence and repulsion of lone pairs.

8. a) What do you mean by '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 of ionization potential:

- Ionization potential (or ionization energy) is the minimum amount of energy required to remove the most loosely bound electron from an isolated gaseous atom in its ground state.
- It is usually expressed in kilojoules per mole (kJ/mol).



First vs. Second ionization potential:

- First Ionization Potential (IP_1):
 - Energy needed to remove the first electron from a neutral atom.
- Second Ionization Potential (IP_2):

- Energy needed to remove a second electron from a unipositive ion (M^+).



Why IP_2 is greater than IP_1 :

- After removing the first electron, the atom becomes positively charged.
- The nuclear attraction on the remaining electrons increases.
- Therefore, it requires more energy to remove the second electron from a more positively charged ion.

Variation of ionization potential with atomic volume:

- As atomic volume (or atomic radius) increases, the outermost electrons are farther from the nucleus.
- This results in weaker attraction between the nucleus and the outer electrons.
- Hence, ionization potential decreases with increasing atomic volume.

Across a period (left to right):

- Atomic size decreases, nuclear charge increases.
- Ionization potential increases.

Down a group (top to bottom):

- Atomic size increases due to the addition of new shells.

- Ionization potential decreases.

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

Definition of f-block elements:

- f-block elements are those in which the last electron enters the $(n-2)f$ orbital.
- These elements are placed in the two rows below the main periodic table.
- They include:
 - Lanthanides (atomic numbers 58 to 71)
 - Actinides (atomic numbers 90 to 103)

Why they are called inner transition elements:

- These elements are called inner transition elements because:
 - The electron enters an inner (f) subshell, even though outer shells are still being filled.
 - Unlike d-block (transition) elements where d orbitals are filled after outer s orbitals, in f-block elements the filling of f orbitals happens before the completion of outer d or p orbitals.
- They show characteristics similar to transition elements such as:
 - Variable oxidation states
 - Formation of colored ions
 - Paramagnetism