ASSIGNMENT

CHEM-2201
Chemistry

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Session : 2021-22
Dept. of Information and Communication Engineering

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Submission Date : 27-04-2025

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5 (a)

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

Answer:

Rutherford proposed that atoms consist of a small, dense, positively charged nucleus around which electrons revolve in circular orbits. Though revolutionary, his model had some serious defects:

- **Instability according to classical physics:** According to Maxwell's electromagnetic theory, a moving charged particle emits radiation. Hence, an electron moving in a circular orbit around the nucleus should continuously emit energy, lose speed, and spiral into the nucleus, causing the atom to collapse. But this doesn't happen, meaning Rutherford's model could not explain atomic stability.
- **No explanation of atomic spectra:** Rutherford's model could not explain the discrete line spectra observed in the light emitted by atoms. If energy was lost continuously, the emission spectrum should be continuous, not consisting of specific lines.

Bohr's Suggestions to Remove These Defects:
Niels Bohr modified Rutherford's model by introducing new ideas:

- Electrons revolve in specific, stable orbits called stationary states without emitting energy.
- Energy is absorbed or emitted only when an electron moves between two stationary orbits, and the energy change corresponds to the difference between the energy levels.
- The angular momentum of an electron is quantized — it can only have certain fixed values (an integral multiple of $h/2\pi$).
  Thus, Bohr's model successfully explained atomic stability and the appearance of 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.

Answer:

Quantum numbers are a set of four numbers that describe the unique position and behavior of an electron within an atom. They define the size, shape, orientation, and spin of the electron's orbital.

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

1. **Principal Quantum Number (n):**
    - Describes the main energy level or shell.
    - Higher 'n' value means the electron is farther from the nucleus and has more energy.
2. **Azimuthal Quantum Number (l):**
    - Describes the shape of the orbital (s, p, d, f types).
    - Values range from 0 to (n–1). (0 = s, 1 = p, 2 = d, etc.)
3. **Magnetic Quantum Number (m):**
    - Describes the orientation of the orbital in space.
    - Values range from –l to +l.
4. **Spin Quantum Number (s):**
    - Describes the spin of the electron — either +½ (up) or –½ (down).

Significance:
These quantum numbers uniquely identify every electron in an atom and explain the structure of the periodic table and chemical behavior of elements.

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 transfer of electrons from a metal to a non-metal.	Formed by sharing of electrons between non-metals.
Melting/Boiling Points	Generally high.	Generally low to moderate.
Physical State	Solid at room temperature; crystalline structure.	Can be solid, liquid, or gas.

Property	Ionic Compounds	Covalent Compounds
Electrical Conductivity	Conduct in molten or aqueous form.	Generally poor conductors.
Solubility	Soluble in polar solvents like water.	Soluble in non-polar solvents like benzene.

Examples:

- **Ionic Compounds:** Sodium chloride ($NaCl$), Magnesium oxide (MgO).
- **Covalent Compounds:** Water (H_2O), Methane (CH_4).

6 (b)

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

Answer:

A co-ordinate covalent bond (also called a dative bond) is a type of covalent bond where both electrons shared in the bond come from the same atom.

- In a normal covalent bond, each atom supplies one electron to the shared pair.
- In a coordinate bond, one atom donates both electrons.

Example:
In the ammonium ion (NH_4^+), the nitrogen atom donates a lone pair of electrons to bond with an H^+ ion.

Thus, although the nature of bonding is covalent (sharing of electrons), the origin of the shared pair distinguishes coordinate covalent bonds.

7 (a)

> ***What do you understand by hydrogen bonds? Classify them with examples. Explain why water has abnormally high boiling point.***

Answer:

A **hydrogen bond** is a weak bond formed when a hydrogen atom attached to a highly electronegative atom (like oxygen, nitrogen, or fluorine) is attracted to another electronegative atom.

Types:

1. **Intermolecular Hydrogen Bonding:** Between different molecules.
    - Example: Water (H_2O) molecules bonding together.
2. **Intramolecular Hydrogen Bonding:** Within the same molecule.
    - Example: Ortho-nitrophenol.

Reason for Water's High Boiling Point:
Water molecules form extensive hydrogen bonding with each other. A large amount of energy is needed to break these strong bonds during boiling. Thus, water has an unusually high boiling point compared to similar-sized molecules.

7 (b)

> ***Why bond angles of H2O and NH3 are 104.5° and 107° respectively although central atoms are sp3 hybridized?***

Answer:

Both H_2O and NH_3 molecules are sp^3 hybridized, meaning they ideally should have a bond angle of 109.5°.

- In NH_3 (Ammonia), there are three bond pairs and one lone pair. The lone pair exerts more repulsion, slightly reducing the bond angle to 107°.
- In H_2O (Water), there are two lone pairs and two bond pairs. The two lone pairs exert even greater repulsion, reducing the bond angle further to 104.5°.

Thus, the bond angle decreases with an increase in the number of lone pairs due to stronger repulsion forces.

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

First vs Second Ionization Potential:

- The first ionization energy is less because it removes an electron from a neutral atom.
- After removing one electron, the atom becomes positively charged. Removing a second electron requires more energy because of stronger electrostatic attraction between the nucleus and the remaining electrons.

Variation with Atomic Volume:

- **Down a group:** Atomic size increases, outer electrons are farther from the nucleus, so ionization energy decreases.
- **Across a period:** Atomic size decreases, nuclear charge increases, so ionization energy increases.

8 (b)

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

Answer:

f-block elements are those elements in which the last electron enters the f-orbital. They are generally placed separately at the bottom of the periodic table.

The f-block includes:

- **Lanthanides (4f series):** From cerium (Ce) to lutetium (Lu).
- **Actinides (5f series):** From thorium (Th) to lawrencium (Lr).

They are called inner transition elements because during their filling, the inner (n–2)f orbitals are being filled while the outer orbitals (s- and d-orbitals) remain incomplete, meaning the transition occurs internally.