Pabna University of Science and Technology

Faculty of Engineering and Technology
Department of Information and Communication Engineering

Assignment

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1. Defects of Rutherford's Model of Atom and Bohr’s Improvements

Defects of Rutherford’s Model:

- Atomic Collapse Problem:
  According to classical electromagnetic theory, a moving electron should emit radiation, lose energy, and spiral into the nucleus, making the atom unstable. But atoms are stable.
- No Explanation for Line Spectra:
  Rutherford's model could not explain why atoms emitted light at specific wavelengths, seen as discrete lines in an atomic emission spectrum (like hydrogen's Balmer series).
- Lack of Quantization:
  The model did not incorporate any idea of quantized energy levels.

Bohr’s Modifications:

- Stable Orbits:
  Electrons revolve around the nucleus in stable orbits called "stationary states" without radiating energy.
- Quantized Energy Levels:
  Each orbit has a definite, fixed energy level. Electrons can only exist in these allowed orbits.
- Emission/Absorption of Radiation:
  Energy is absorbed or emitted when an electron jumps from one orbit to another, explaining discrete spectral lines.
- Formula Introduced:
  Bohr introduced formulas linking energy levels and emission frequencies using Planck’s quantum theory.

2. Quantum Numbers and Their Importance

Definition:
Quantum numbers specify the properties and behavior of atomic orbitals and the electrons inside them.
The Four Quantum Numbers:

Quantum Number	Symbol	What it Represents	Allowed Values	Example
Principal	n	Energy level, size	1, 2, 3, 4…	n=1 for 1st shell
Azimuthal	l	Shape of orbital	0 to (n-1)	l=0 (s), 1 (p), 2 (d)
Magnetic	m	Orientation	-l to +l	For l=1: m=-1,0,+1
Spin	s	Electron spin direction	+½ or -½	

Significance:
- Defines the energy, shape, orientation, and spin of an electron.
- No two electrons in the same atom can have identical values of all four quantum numbers (Pauli's Exclusion Principle).

3. Comparison Between Ionic and Covalent Compounds
Properties of Ionic Compounds:
- Formed by transfer of electrons from metal to non-metal.
- Strong electrostatic force between oppositely charged ions.
- High melting and boiling points.
- Conduct electricity in molten or aqueous state.
- Generally soluble in water but insoluble in organic solvents.

Examples:
Sodium chloride ($NaCl$), Calcium chloride ($CaCl_2$)

Properties of Covalent Compounds:
- Formed by sharing of electrons between non-metal atoms.
- Weak intermolecular forces (except network solids like diamond).
- Low melting and boiling points (except for giant molecules like SiO_2).
- Poor conductors of electricity.
- Soluble in organic solvents, insoluble in water (mostly).

Examples:
Methane (CH_4), Carbon dioxide (CO_2)

4. Coordinate Covalent Bond

Definition:
A type of covalent bond where both electrons shared in the bond originate from the same atom.

How Different from a Normal Covalent Bond:

- Normal Covalent Bond: Each atom contributes one electron.
- Coordinate Covalent Bond: One atom donates both electrons.

Example:
In ammonium ion (NH_4^+), nitrogen donates a lone pair to a proton (H^+).

Diagram:
yaml
CopyEdit

```
  H
  |
H–N → H
  |
  H
```

(Arrow shows coordinate bond.)

5. Hydrogen Bonding and Water's High Boiling Point

Definition:
Attractive force between a hydrogen atom covalently bonded to a highly electronegative atom (N, O, or F) and another electronegative atom.

Types of Hydrogen Bonding:

- Intermolecular Hydrogen Bonding:
  e.g., H_2O molecules bonding with each other.
- Intramolecular Hydrogen Bonding:
  e.g., o-nitrophenol, hydrogen bonding within a single molecule.

Why Water Has a High Boiling Point:

- Extensive hydrogen bonding between water molecules requires significant energy to break.
- This strong intermolecular force leads to a much higher boiling point compared to other similar-sized molecules like H_2S.

6. Bond Angles of H_2O and NH_3 and Explanation

Bond Angles:

- H_2O (Water): 104.5°
- NH_3 (Ammonia): 107°

Why They Differ:

- Both oxygen and nitrogen atoms are sp^3 hybridized.
- In water (H_2O):
  - Oxygen has two lone pairs.
  - Lone pair-lone pair > lone pair-bond pair > bond pair-bond pair repulsions.
  - This causes compression of bond angle to 104.5°.
- In ammonia (NH_3):
  - Nitrogen has only one lone pair, so the bond angle remains larger at 107°.

Diagram:

- H_2O: Bent shape.
- NH_3: Trigonal pyramidal shape.

7. Ionization Potential of an Element

Definition:

The minimum amount of energy required to remove an electron from a gaseous atom or ion.

First vs Second Ionization Potential:

- First Ionization Potential: Energy to remove the first electron.
- Second Ionization Potential: Higher, because removing an electron from a positively charged ion is more difficult due to stronger nuclear attraction.

Variation with Atomic Volume:

- Down a Group:
  - Atomic radius increases.
  - Valence electrons are farther from the nucleus.
  - Ionization energy decreases.
- Across a Period:

    - Nuclear charge increases.
    - Atomic radius decreases.
    - Ionization energy increases.

Graph:

```
Ionization Energy
|
|            /
|          /
|       /
|     /
|   /
| /
|/_________________________________________→ Atomic Number
```

8. f-Block Elements and Inner Transition Elements

Definition of f-Block Elements:

- Elements where the last electron enters an f-orbital.
- Include Lanthanides (58-71) and Actinides (90-103).

Characteristics:

- Show variable oxidation states.
- Have high melting and boiling points.
- Most lanthanides are silvery-white metals.
- Actinides are mostly radioactive.

Why Called Inner Transition Elements:

- They involve the filling of (n-2)f orbitals.

- Transition occurs in the inner orbitals, not the outermost ones, differentiating them from d-block transition metals.

Examples:

- Lanthanides: Cerium (Ce), Neodymium (Nd)
- Actinides: Uranium (U), Thorium (Th)