Pabna University of Science and Technology

Information and Communication Engineering

Assignment

Course Name : Chemistry

Course Code : CHEM-2201

Submitted By:

Name: M.Q.A Haniya Khanam

Roll: 220632

Session:2021-2022

2nd year 2nd semester

Department of ICE, PUST

Submitted to :

Abrar Yasir Abir

Lecturer

Department of Chemistry,PUST

Date of submission:27-04-2025

Answer to the question no: 05

(a)Give the defects of Rutherford‟s model of atom. What suggestions were given by Bohr to remove these defects.

Rutherford's model of the atom, while groundbreaking, had several key defects. These defects were mainly related to the stability of the atom and the emission spectrum of elements. Let's discuss the defects and then look at the suggestions given by Niels Bohr to improve upon Rutherford's model.

Defects of Rutherford's Model:

1. **Instability of the Atom**:
   - According to Rutherford's model, electrons revolve around the nucleus in circular orbits, much like planets orbiting the sun. This implies that an electron, which is a charged particle, would experience centripetal acceleration as it moves in its orbit.
   - According to classical electromagnetic theory, any charged particle undergoing acceleration emits radiation (electromagnetic radiation), which would cause the electron to lose energy.
   - As the electron loses energy, it would spiral inward toward the nucleus, eventually collapsing into the nucleus. This would result in the atom's collapse, which contradicts the observed stability of atoms.
2. **Continuous Spectrum**:
   - Rutherford's model could not explain why atoms emitted only specific wavelengths of light (discrete spectral lines) when heated or excited. According to classical theory, the electron would emit a continuous spectrum of radiation, not the discrete lines observed in atomic spectra.
3. **Electron Orbits**:
   - Rutherford's model did not specify any quantization of electron orbits. It did not explain why electrons in specific orbits do not radiate energy or why only certain orbits (and hence certain energy levels) are allowed.

Bohr's Suggestions to Remove These Defects:

Niels Bohr introduced a new atomic model in 1913 to overcome the defects of Rutherford's model. His model was based on quantum principles and the idea of quantized electron orbits. The key suggestions Bohr made were:

1. **Quantization of Electron Orbits**:

- Bohr proposed that electrons could only occupy certain orbits or energy levels around the nucleus, and these orbits were quantized. Electrons in these stable orbits did not emit radiation, which resolved the problem of atomic instability.
- The radii of these orbits were determined by the integer values of quantum numbers ($n = 1, 2, 3,...$), and only certain specific orbits were allowed.

2. **Energy Emission and Absorption**:
   - Bohr postulated that electrons could move from one orbit to another, and when they did so, they would absorb or emit energy in discrete amounts (quanta). The energy of the emitted or absorbed radiation corresponded to the difference in energy between the two orbits.
   - This explained the discrete spectral lines observed in atomic spectra. Each transition between orbits emitted a photon of specific energy, and thus only specific wavelengths of light were observed.
3. **Electron Stability in Quantized Orbits**:
   - In Bohr's model, electrons in stable orbits do not lose energy by radiation. This solved the problem of the electron spiraling into the nucleus, as Rutherford's model had predicted.
4. **Bohr's Postulates**:
   - **First Postulate**: Electrons revolve in stable orbits without radiating energy, corresponding to certain allowed energy levels.
   - **Second Postulate**: An electron can jump from one orbit to another by absorbing or emitting a quantum of electromagnetic radiation, corresponding to the difference in energy between the two orbits.

(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.

Quantum Number:

A **quantum number** is a number that describes the properties and behavior of an electron in an atom. It defines the electron's energy, its orbital shape, its orientation in space, and its spin. Quantum numbers arise from quantum mechanics and are essential for understanding atomic structure.

How Many Quantum Numbers Does an Electron Have in an Orbital?

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

1. **Principal Quantum Number (n)**
2. **Angular Momentum Quantum Number (l)**
3. **Magnetic Quantum Number (m_l)**
4. **Spin Quantum Number (m_s)**

Significance of Each Quantum Number:

1. **Principal Quantum Number (n):**
   - Describes the **energy level** or **shell** where the electron is located.
   - Determines the **size** and **energy** of the orbital.
   - Can take any positive integer value: n=1,2,3,…
2. **Angular Momentum Quantum Number (l):**
   - Defines the **shape** of the orbital.
   - For each n, l can range from 0 to n−1
   - Values of l correspond to specific orbital types: l=0 (s orbital), l=1(p orbital), l=2 (d orbital), etc.
3. **Magnetic Quantum Number (m_l):**
   - Describes the **orientation** of the orbital in space.
   - For each l, m can range from −l to +l including 0.
4. **Spin Quantum Number (m_s):**
   - Describes the **spin** of the electron.
   - Can have values of +1/2 or −1/2, indicating the two possible spin states of an electron.

Answer to the question no : 06

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

Comparison of Properties of Ionic and Covalent Compounds:

Property	Ionic Compounds	Covalent Compounds
Formation	Formed by the transfer of electrons between metals and non-metals.	Formed by the sharing of electrons between non-metals.
Bond Type	Ionic bond (electrostatic attraction between oppositely charged ions).	Covalent bond (sharing of electrons).
Melting and Boiling Points	Generally high due to strong electrostatic forces between ions.	Generally low, as intermolecular forces are weaker.
Electrical Conductivity	Conducts electricity in molten or dissolved state due to free ions.	Does not conduct electricity, as there are no free ions or electrons.

Property	Ionic Compounds	Covalent Compounds
Solubility	Usually soluble in water but insoluble in non-polar solvents.	Often soluble in non-polar solvents, but may not dissolve well in water.
State at Room Temperature	Usually solid (forming crystalline structures).	Can be solid, liquid, or gas, depending on the substance.

Examples of Each Type of Compound:

- **Ionic Compounds**:
  1. **Sodium chloride (NaCl)**: Formed by the transfer of an electron from sodium (Na) to chlorine (Cl), creating Na^+ and Cl^- ions.
  2. **Magnesium oxide (MgO)**: Formed by the transfer of electrons from magnesium (Mg) to oxygen (O), creating Mg^{2+} and O^{2-} ions.
- **Covalent Compounds**:
  1. **Water (H_2O)**: Formed by the sharing of electrons between hydrogen (H) and oxygen (O) atoms.
  2. **Carbon dioxide (CO_2)**: Formed by the sharing of electrons between carbon (C) and oxygen (O) atoms.

(b)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**.

Here's how it differs from a normal covalent bond:

Feature	Normal Covalent Bond	Coordinate Covalent Bond
Electron Contribution	Each atom contributes one electron to the bond	One atom contributes **both** electrons
Formation	Happens between atoms needing to share electrons to complete their octet	Happens when one atom with a lone pair donates to another atom with an empty orbital
Example	H–Cl (both H and Cl share one electron)	NH_4^+ (ammonia donates a lone pair to H^+)

Answer to the question no : 07

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

Hydrogen Bonds: A hydrogen bond **is a type of** weak electrostatic attraction **that occurs when a** hydrogen atom**, already** covalently bonded **to a highly** electronegative atom **(like** nitrogen, oxygen, or fluorine**), is attracted to another** electronegative atom **nearby.**

Classification of Hydrogen Bonds:

1. Intermolecular Hydrogen Bond:

- **Occurs between molecules.**
- Example: **Water (H_2O)**
  In water, the hydrogen of one molecule forms a hydrogen bond with the oxygen of another molecule.
- Other examples: Hydrogen fluoride (HF), ammonia (NH_3), alcohols.

2. Intramolecular Hydrogen Bond:

- **Occurs within the same molecule.**
- Common when two groups capable of hydrogen bonding are close together.
- Example: **o-nitrophenol**
  Here, the -OH group and -NO_2 group within the same molecule form a hydrogen bond.

Water has a much **higher boiling point** than expected for a molecule of its size. This is because of **strong hydrogen bonding** between water molecules:

- Each water molecule can form **up to four hydrogen bonds** (two through hydrogen atoms, two through lone pairs on oxygen).
- These hydrogen bonds require **extra energy** to break before water molecules can escape into the vapor phase.
- Therefore, a lot more heat is needed to boil water compared to other similar-sized molecules (like H_2S, which doesn't form strong hydrogen bonds).

(b)Why bond angles of H2O and NH3 are 104.5° and 107° respectively although central atoms are sp 3 hybridized.

Both **H_2O (water)** and **NH_3 (ammonia)** have **sp^3 hybridized central atoms**, which ideally form a **tetrahedral geometry** with a bond angle of **109.5°**.
However, their actual bond angles are **less** than 109.5° due to the presence of **lone pairs** of electrons.

In NH_3 (Ammonia):

- Nitrogen has **3 bonding pairs** and **1 lone pair**.
- The lone pair exerts **greater repulsion** than bonding pairs.
- This **compresses the H–N–H bond angle** from 109.5° to **~107°**.

In H_2O (Water):

- Oxygen has **2 bonding pairs** and **2 lone pairs**.
- The **two lone pairs** repel even more strongly.
- This pushes the hydrogen atoms **closer together**, reducing the H–O–H angle to **~104.5°**.

<u>Answer to the question no : 08</u>

(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?

Ionization potential (also called **ionization energy**) is the **energy required to remove the most loosely bound electron** from an isolated gaseous atom to form a **positive ion**.

The **first ionization potential** is less than the **second** because, after the first electron is removed, the remaining electrons are more tightly held by the nucleus due to a **higher effective nuclear charge**. This stronger attraction makes it harder to remove the second electron, requiring more energy.

The ionization potential of an element vary with atomic volume-The **ionization potential** of an element **decreases** as its **atomic volume increases**.

This is because, as the atomic volume increases, the **outermost electrons** are **farther** from the nucleus and experience **less attraction** from the nucleus. Consequently, **less energy** is required to remove an electron, leading to a **lower ionization potential**.

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

The **f-block elements** are elements in which the **f-orbitals** are being filled with electrons. These elements are located in the **two rows at the bottom of the periodic table**, which are:

- **Lanthanides (rare earth elements)**: These are the 14 elements from **Lanthanum (La) to Lutetium (Lu)**, with the electron configuration filling the 4f orbitals.
- **Actinides**: These are the 14 elements from **Thorium (Th) to Lawrencium (Lr)**, with the electron configuration filling the 5f orbitals.

The **f-block elements** are called **inner transition elements** because they are part of the **transition series** of elements, but their electrons fill the **inner f orbitals** (compared to the d-block elements, which fill d orbitals).

- **Transition elements** typically fill the d-orbitals in the second-last electron shell (n-1).
- **Inner transition elements** fill the f-orbitals in the third-last electron shell (n-2).

This makes them **"inner"** transition elements, as their electron configurations involve filling inner orbitals (f orbitals) rather than the outer ones (d orbitals).