ASSIGNMENT

CHEM-2201
Chemistry

Submitted By :
Al Kavy Joy
Roll : 220627
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Submitted To :
Abrar Yasir Abir
Lecturer
Department of Chemistry
Pabna University of Science and Technology

PABNA UNIVERSITY OF SCIENCE AND TECHNOLOGY

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:

1. **Stability Issue**: According to classical electromagnetic theory, a charged particle in motion (such as an electron) should emit electromagnetic radiation, causing it to lose energy. As a result, the electron would spiral inward and eventually collapse into the nucleus. However, atoms are stable, which contradicts Rutherford's model.
2. **Continuous Spectrum Problem**: Rutherford's model could not explain the discrete line spectra observed in atomic emissions, particularly in hydrogen, where the light emitted by atoms is not continuous but consists of distinct lines.

Bohr's Suggestions to Remove These Defects:

1. **Quantization of Electron Orbits**: Bohr proposed that electrons exist in certain fixed orbits around the nucleus where they do not emit radiation. These orbits are called "stationary states."
2. **Electron Energy Emission and Absorption**: Electrons can absorb or emit energy only when they move between these stationary orbits. The energy difference between the orbits corresponds to the frequency of the emitted or absorbed radiation.
3. **Angular Momentum Quantization**: Bohr also suggested that the angular momentum of an electron in a stationary orbit is quantized and is given by L=nℏL = n\hbarL=nℏ, where nnn is a positive integer and ℏ\hbarℏ is the reduced Planck's constant.

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.
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:

- The f-block elements are those elements in which the last electron enters the f-orbital. These elements are found in the two rows placed at the bottom of the periodic table.
- The f-block elements consist of two series:
    1. **Lanthanides** (Rare Earth Elements): These are the 14 elements with atomic numbers from 57 (Lanthanum) to 71 (Lutetium).
    2. **Actinides**: These are the 14 elements with atomic numbers from 89 (Actinium) to 103 (Lawrencium).

Why f-block elements are called inner transition elements:

- The f-block elements are also called **inner transition elements** because they involve the filling of the **f-orbitals**, which are inner orbitals that are located below the main body of the periodic table (hence, they are "inner").
- These elements transition through the filling of the f-orbitals, which is why they are called **transition elements**, but they are "inner" because they are placed in the inner sections of the periodic table, separate from the transition metals (d-block elements).