PABNA UNIVERSITY OF SCIENCE AND TECHNOLOGY

INFORMATION AND COMMUNICATION ENGINEERING

Course Name: CHEMISTRY

CHEM-2201
ASSIGNMENT

SUBMITTED BY

JUBAYER AHMMED

ROLL : 220638

SUBMITTED TO:

Abrar Yasir Abir

Lecturer

Department of Chemistry

Pabna University of Science and Technology

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

Answer: Rutherford's nuclear model (1911) proposed that:

- An atom has a tiny, dense, positively charged nucleus.
- Electrons revolve around the nucleus like planets around the sun.

However, it had major defects:

(1) Rutherford's model was unable to explain the stability of an atom. According to Rutherford's postulate, electrons revolve at a very high speed around a nucleus of an atom in a fixed orbit. However, Maxwell explained accelerated charged particles release electromagnetic radiations. Therefore, electrons revolving around the nucleus will release electromagnetic radiation.

(2) The electromagnetic radiation will have energy from the electronic motion as a result of which the orbits will gradually shrink. Finally, the orbits will shrink and collapse in the nucleus of an atom. According to the calculations, if Maxwell's explanation is followed Rutherford's model will collapse with 10-8 seconds. Therefore, Rutherford atomic model was not following Maxwell's theory and it was unable to explain an atom's stability.

(3) Rutherford's theory was incomplete because it did not mention anything about the arrangement of electrons in the orbit. This was one of the major drawbacks of Rutherford atomic model.

Bohr's Atomic Model (1913) – Fixing the Defects

Niels Bohr proposed quantized electron orbits to solve these issues. His key postulates:

In 1913 Niels Bohr, an eminent scientist of Denmark successfully explained the stability of the atom and the reason behind the appearance of line spectra with the help of Plank's quantum theory. The theory, put forward by Bohr regarding the structure of H-atom, is based on three revolutionary postulates

Postulate 1: When an electron revolves in any selected orbits, it neither emits nor absorbs energy. The energy of an electron in a particular orbit is constant. These orbits are therefore, called stationary orbits. Depending on the distance of from the nucleus, these orbits are divided into energy levels such as K, L, M, N … etc. and these are designated respectively by the numbers 1, 2, 3, 4…etc.

Postulate 2: The electron in the hydrogen atom revolves around the nucleus only in certain selected circular paths (called orbit) which are associated with definite energies. Only those orbits are permitted for the revolving electron in which the angular momentum of the electron is a whole number multiple of $h/2\pi$ angular momentum of electron, Where, n = 1, 2, 3, 4…etc. m = mass of the electron v = velocity of the electron r = radius of the orbit h = Plank's constant mvr = $n \times h\ 2\pi$ Thus the angular momentum of electrons in an atom is quantized.

Postulate 3: As long as an electron remains in a particular orbit, it neither emits (i.e., radiates or losses) nor absorbs energy. But when the electron is excited from a lower energy level to a higher energy level, it absorbs energy. On the other hand when it comes back from a higher energy level to a lower energy level, it emits energy. Now if ν the frequency of the radiation absorbed by the electron, then, according to Plank's quantum theory of radiation, energy absorbed by the electron will be

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

This atomic model overcomes the deffect of Rutherford‟s model since describe about:

1. Stability of atom
2. Line spectra of H-atom
3. Radius of the first orbit of H-atom
4. Principal quantum number
5. Energy of an electron

Question-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: Certain identification numbers which are necessary to give complete information about any revolving electron in an atom are quantum numbers. There are four quantum numbers used to describe the exact position and the nature of an electron. They are:

1.Principal quantum number: Different energy levels in an atom according to Bohr theory have been represented by n which can have non-zero positive integer values upto infinity i.e.,

$$n = 1, 2, 3, 4 \ldots \infty$$

Thus n can have infinite values and it has been called principal quantum number.

2.Azimuthal quantum number: In order to explain the formation of fine structure, Sommerfeld proposed the existence of elliptical orbits, besides the circular orbits of Bohr. In order to specify the shape of the elliptical orbit, another supplementary quantum number is necessary. This supplementary quantum number which indicates the ellipticity of the electronic orbit (sub-shell) is called azimuthal or subsidiary quantum number denoted by 'l'. i.e., l = '0' → (n-1)

Values of l	→	0	1	2	3
Designation of sub-shell	→	s	p	d	f

Magnetic quantum number: Due to the angular motion of electron around the nucleus, a magnetic field is produced which interacts with the external magnetic field. As a result sub-shell of definite energy split into three dimensional spatial regions which can be expressed by magnetic quantum number. It is denoted by 'm'. i.e., m = -l → +l

Spin quantum number: It is observed that the electron in an atom is not only revolved around the nucleus but is also spinning around its own axis like the earth. As this quantum number describes the spin motion of the electron. It is designated by ‘s’. Since the electron can spin only clockwise and therefore the spin quantum number can take only two values be +1/2 or -1/2.

The Four Quantum Numbers & Their Significance

Quantum Number	Symbol	Represents	Possible Values	Significance
Principal (n)	n	Energy level & size of orbital	n=1,2,3,4,....	Determines the shell (e.g., K, L, M) and average distance from the nucleus. Higher n = higher energy.
Azimuthal (l)	l	Orbital shape (subshell)	l= 0 to n-1	Defines the shape of the orbital: l=0 ⟶ s-orbital (spherical)- l=1 ⟶ p-orbital (dumbbell) l=2 ⟶ d-orbital (cloverleaf) l=3 ⟶ f-orbital (complex)
Magnetic (m_l)	m_l	Orbital orientation	m_l = -l to +l	Describes how the orbital is oriented in space: for l = 1 (p-orbital), m_l = -1,0,+1 ⟶ 3 orientations(p_x, p_y,p_z)
Spin(m_s)	m_s	Electron spin	$m_s = +\frac{1}{2}$ or $-\frac{1}{2}$	Indicates clockwise or counterclockwise spin of the electron (↑ or ↓). Explains magnetic properties and Pauli’s exclusion principle.

Significance of Quantum Numbers

1. Define Electron Address:
    - Just like a postal address, quantum numbers pinpoint an electron's location in an atom.
2. Explain Atomic Structure:
    - They determine orbital shapes, energies, and electron arrangements (e.g., why s-block has 2 electrons, p-block has 6).
3. Pauli's Exclusion Principle:
    - No two electrons in an atom can have the same set of four quantum numbers (prevents electron crowding).
4. Chemical Bonding & Spectroscopy:
    - Differences in quantum numbers explain bond formation and spectral lines.

Question-6.(a) Compare the properties of ionic and covalent compounds. Give two examples of each type of compounds.

Answer: The comparison is described below:

Properties	Ionic Compounds	Covalent Compounds
Nature of Bond	Formed by transfer of electrons (metal + non-metal)	Formed by sharing of electrons (non-metals)
Physical State	Usually solids (crystalline) at room temperature	Can be gases, liquids, or solids
Melting and Boiling Points	High melting and boiling points	Generally low melting and boiling points
Electrical Conductivity	Conduct electricity when molten or dissolved in water	Generally poor conductors of electricity
Solubility	Soluble in polar solvents like water	Soluble in non-polar solvents like benzene

Examples	Sodium chloride (NaCl), Magnesium oxide (MgO)	Water (H_2O), Carbon dioxide (CO_2)

Question-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 in the bond are donated by the same atom.In a normal covalent bond, each atom contributes one electron to form the bond.
Example of co-ordinate bond: In the formation of the ammonium ion (NH_4^+), the nitrogen atom donates a lone pair to bond with an H^+ ion.

Difference Summary:

Aspect	**Normal Covalent Bond**	**Co-ordinate Covalent Bond**
Electron Contribution	One electron from each atom	Both electrons from one atom

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

Answer:

Hydrogen bond: An attractive force between a hydrogen atom covalently bonded to a highly electronegative atom (like F, O, or N) and another electronegative atom.

Types of Hydrogen Bonds:

- **Intermolecular Hydrogen Bond:** Between molecules.
  Example: Between two water (H_2O) molecules.

- **Intramolecular Hydrogen Bond:** Within the same molecule.
  Example: In o-nitrophenol.

Reason for High Boiling Point of Water:

- Water molecules form strong intermolecular hydrogen bonds, which require more energy to break.

- This results in water having an abnormally high boiling point compared to other similar-sized molecules.

Question-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 involve sp^3 hybridization of the central atom. However, the bond angles differ because of the number of lone pairs:

- **Ammonia (NH_3):**
  One lone pair of electrons is present. Due to the lone pair-bond pair repulsion, the bond angle is slightly reduced to 107° from the ideal 109.5°.

- **Water (H_2O):**
  Two lone pairs of electrons are present. The lone pairs exert greater repulsion on bond pairs,

further reducing the bond angle to 104.5°.

Thus, greater the number of lone pairs, greater the repulsion, and smaller the bond angle.

Question-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: The Ionization Potential (also called Ionization Energy) of an element is the amount of energy required to remove the most loosely bound electron from an isolated gaseous atom to form a cation.

Why First Ionization Potential is Less than Second:

The first ionization removes one electron, making the atom positively charged. The second ionization requires removing another electron from an already positively charged ion, which has a greater hold on its remaining electrons. Hence, the second ionization potential is always greater than the first.

Variation of Ionization Potential with Atomic Volume:

- As atomic volume increases (i.e., moving down a group), ionization potential decreases because outer electrons are farther from the nucleus and experience less electrostatic attraction.
- As atomic volume decreases (moving across a period from left to right), ionization potential increases due to stronger nuclear attraction.

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

Answer: f-Block Elements are elements in which the last electron enters into the f-orbital of their atoms.

- These include the lanthanides (elements 58 to 71) and actinides (elements 90 to 103).
- They are placed separately at the bottom of the periodic table to maintain the table's structure.

Why Called Inner Transition Elements: They are called "inner transition elements" because their f-electrons are being filled in an inner (n-2) shell, while the outer shells remain unchanged. They lie between the s-block and d-block elements, representing a transition within the inner electron shells.