Pabna University of Science and Technology

Faculty of Engineering and Technology

Department of Information and Communication Engineering

Assignment

Course Code: **CHE-2201**

Course title**: Chemistry**

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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:

Here is the answer:

<u>Defects of Rutherford’s Model of Atom:</u>

Rutherford said electrons revolve around the nucleus in circular paths.

But according to electromagnetic theory, a moving electron should emit energy.

As it emits energy, the electron would lose energy and spiral into the nucleus.

This means atoms should collapse, but in reality, atoms are stable.

Rutherford’s model could not explain the stability of atoms.

Also, it could not explain the line spectra (discrete colors) observed in hydrogen and other elements.

Therefore, Rutherford’s model was incomplete.

<u>Bohr’s Suggestions to Remove These Defects:</u>

Bohr proposed that electrons revolve around the nucleus only in fixed orbits called energy levels.

While moving in these orbits, electrons do not radiate any energy.

These orbits are associated with fixed energy, and are called stationary states.

Electrons can jump from one orbit to another.

When an electron jumps from a higher energy orbit to a lower one, it emits energy in the form of light.

When it jumps from a lower to a higher orbit, it absorbs energy.

This explained the stability of atoms and the discrete line spectra observed.

5(b): What do you understand by the term, " Quantum number". How many quantum numbers has an electron in an orbital? And its significance.

Answer:

<u>What is meant by "Quantum Number"?</u>

Quantum numbers are a set of numbers that give complete information about the state of an electron in an atom.

They describe the position, energy, size, shape, and orientation of the orbital where the electron is found.

Each electron in an atom has a unique set of quantum numbers, just like a person's identity card.

<u>How many Quantum Numbers are there for an electron?</u>

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

- <u>Principal Quantum Number (n):</u>

It indicates the main energy level or shell where the electron is present.

It shows the size of the orbital and the energy of the electron.

Values of 'n' are positive integers: n = 1, 2, 3, 4,...

Example: n=1 is the first shell (K-shell), n=2 is the second shell (L-shell), etc.

- <u>Azimuthal Quantum Number (l):</u>

It is also called the orbital angular momentum quantum number.

It tells the shape of the orbital (like spherical, dumbbell, etc.).

For a given 'n', 'l' can have values from 0 to (n-1).

Each value of 'l' represents a type of orbital:

l = 0 → s orbital (spherical)

l = 1 → p orbital (dumbbell-shaped)

l = 2 → d orbital (clover-shaped)

l = 3 → f orbital (complex shape)

- Magnetic Quantum Number (m):

It gives the orientation of the orbital in space.

For a given 'l', 'm' can have values from -l to +l, including zero.

Example: If l=1 (p orbital), m can be -1, 0, or +1, meaning the p orbital can orient in three different directions.

- Spin Quantum Number (s):

It describes the spin or rotation of the electron around its own axis.

An electron can spin in two directions:

s = +½ (clockwise spin)

s = –½ (anti-clockwise spin)

This explains why each orbital can hold a maximum of two electrons with opposite spins.

Significance of Quantum Numbers:

They uniquely identify every electron in an atom.

They explain the distribution of electrons in shells, subshells, and orbitals.

They describe the energy of electrons and help in understanding atomic structure.

Quantum numbers help explain chemical behavior, bonding, and the properties of elements based on electron configuration.

They also explain phenomena like the structure of the periodic table, magnetic properties, and spectral lines.

<u>In short:</u>

Quantum numbers act like the address of an electron, telling where exactly it is inside the atom and how it behaves.

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

Answer: Comparison of Properties: Ionic Compounds vs Covalent Compounds

Property	Ionic Compounds	Covalent Compounds
Nature of bond	Formed by **transfer of electrons** from one atom to another (metal + non-metal).	Formed by **sharing of electrons** between atoms (non-metal + non-metal).
Physical state	Usually **solid** and crystalline at room temperature.	Can be **solid, liquid, or gas** at room temperature.
Melting and boiling points	Have **high** melting and boiling points.	Have **low** melting and boiling points (except few like diamond).
Electrical conductivity	Conduct electricity when **dissolved in water** or in **molten** state.	Generally **do not conduct** electricity.
Solubility	Mostly soluble in **water**, insoluble in organic solvents.	Mostly soluble in **organic solvents**, insoluble in water.
Bond strength	Strong electrostatic forces between ions.	Bonds may be strong, but forces between molecules are weaker.

Ionic Compounds:

1. Sodium chloride ($NaCl$)
2. Magnesium oxide (MgO)

Covalent Compounds:

3. Water (H_2O) 4. Carbon dioxide (CO_2)

6(b): What is a co-ordinate covalent bond? How does it differ from a normal covalent bond?

Answer:

<u>What is a Co-ordinate Covalent Bond?</u>

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

One atom donates a lone pair of electrons, and the other atom accepts it.

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

<u>Difference between Co-ordinate Covalent Bond and Normal Covalent Bond:</u>

Normal Covalent Bond	Co-ordinate Covalent Bond
Each atom contributes one electron	One atom contributes both electrons
Formed by mutual sharing of electrons	Formed by donation of lone pair
Shown by a single line (–)	Often shown by an arrow (→) from donor to acceptor
Example H_2, O_2, CO_2	Example NH_4^+, H_3O^+, CO

In short:

Normal covalent bond = mutual sharing.

Co-ordinate covalent bond = one-sided donation of electron pair.

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

Answer:

<u>What is a Hydrogen Bond?</u>

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.

It is weaker than a covalent bond but stronger than van der Waals forces.

Classification of Hydrogen Bonds: Hydrogen bonds are of two types:

Intermolecular Hydrogen Bond:

Hydrogen bond forms between two different molecules.

<u>Example:</u> In water (H_2O), hydrogen bonds form between one water molecule and another.

<u>Intramolecular Hydrogen Bond:</u>

Hydrogen bond forms within the same molecule.

Example: In ortho-nitrophenol ($C_6H_4(OH)(NO_2)$), hydrogen bonding occurs inside the molecule between the -OH and -NO_2 groups.

Why does Water have an Abnormally High Boiling Point? Water molecules are strongly held together by hydrogen bonds.

A lot of energy is needed to break these hydrogen bonds.

As a result, water boils at a much higher temperature compared to other similar-sized molecules.

Without hydrogen bonding, water would have boiled at a much lower temperature.

<u>In short:</u> Hydrogen bonding makes water molecules "stick together," needing more heat to separate them, which causes a high boiling point.

7(b): Why bond angles of H2O and NH3 are 104.5° and 107° respectively although central atoms are sp3 hybridized.

Answer:

Why are bond angles of H_2O and NH_3 different even though both are sp^3 hybridized?

In both H_2O and NH_3, the central atoms (Oxygen in H_2O and Nitrogen in NH_3) are sp^3 hybridized.

In ideal sp^3 hybridization, the bond angle should be 109.5°.

But in reality, the bond angles are different because of lone pairs.

Reason in detail:

Ammonia (NH_3):

Nitrogen has one lone pair and three bond pairs (attached to three hydrogens).

Lone pairs repel more strongly than bond pairs.

This strong repulsion compresses the bond angle slightly from 109.5° to 107°.

Water (H_2O):

Oxygen has two lone pairs and two bond pairs (attached to two hydrogens).

Two lone pairs cause even stronger repulsion.

As a result, the bond angle gets further reduced from 109.5° to 104.5°.

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:

<u>What is the "Ionization Potential" of an Element?</u>

The ionization potential (also known as ionization energy) is the energy required to remove an electron from a neutral atom or molecule in its gaseous state.

It is measured in electron volts (eV) or kilojoules per mole (kJ/mol).

First ionization potential refers to the energy required to remove the first electron, second ionization potential refers to removing the second electron, and so on.

<u>Why is the First Ionization Potential Less Than the Second Ionization Potential?</u>

The first ionization potential is the energy required to remove the first electron from a neutral atom.

After the first electron is removed, the atom becomes a positively charged ion (cation). The second ionization potential is the energy needed to remove an electron from this positively charged ion.

<u>The second ionization potential is higher because:</u>

The cation is more positively charged than the neutral atom, so the electrostatic attraction between the nucleus and the remaining electrons is stronger.

The electrons are held more tightly by the positively charged ion, requiring more energy to remove the next electron.

<u>How Does the Ionization Potential of an Element Vary with Atomic Volume?</u>

The atomic volume refers to the size or volume of an atom.

Ionization potential and atomic volume are related in the following way:

As atomic volume increases, the distance between the nucleus and the outermost electrons also increases.

This means that the outermost electrons are farther from the nucleus, which results in less force of attraction between the nucleus and the electrons.

As a result, the ionization potential decreases because it becomes easier to remove an electron from a larger atom.

In smaller atoms, the electrons are closer to the nucleus, and the attraction is stronger, making it harder to remove an electron, leading to higher ionization potential.

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

Answer:

<u>What do you mean by f-block elements?</u>

The f-block elements are elements in the periodic table where the f-orbitals (inner orbitals) are being filled with electrons.

These elements are found in the two rows at the bottom of the periodic table, which are:

Lanthanides (rare earth metals), elements 57 to 71.

Actinides, elements 89 to 103.

The f-block elements include metals that have their electrons filling the 4f and 5f orbitals, respectively.

<u>Why are f-block elements called Inner Transition Elements?</u>

f-block elements are called inner transition elements because they are transitioning between the s-block and d-block elements, but the transition occurs inside the periodic table, involving the f-orbitals.

These elements are called inner transition elements because their electrons are filling inner orbitals (f-orbitals), rather than the outer orbitals.

The lanthanides and actinides are placed separately at the bottom of the periodic table to keep the main body of the table compact. Although they are chemically similar to transition metals, their filling of f-orbitals distinguishes them from the d-block transition metals.

Summary:

f-block elements: Elements where electrons fill the f-orbitals (4f and 5f), consisting of lanthanides and actinides.

Inner transition elements: These elements are called "inner" because their transition involves the inner f-orbitals, not the outermost s- or d-orbitals.

In short: f-block elements are inner transition elements because their electron filling takes place in the inner f-orbitals, and they are positioned at the bottom of the periodic table as a separate block.