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Department of Information and Communication Engineering

ASSIGNMENT

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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 : Defects of Rutherford's Model of the Atom:

1. Instability of Atom:

According to classical electromagnetic theory, an electron revolving around the nucleus should continuously emit energy (since a charged particle moving in a circle radiates energy).

As it loses energy, the electron should spiral inward and eventually fall into the nucleus, making the atom unstable. But real atoms are stable, so Rutherford's model couldn't explain this.

2. No Explanation of Atomic Spectra:

Rutherford's model couldn't explain why atoms emit light at specific wavelengths (i.e., line spectra).

If electrons lost energy continuously, atoms should give a continuous spectrum, not a line spectrum.

Suggestions by Bohr to Remove These Defects:

1. Quantized Orbits:

Bohr proposed that electrons revolve around the nucleus only in certain allowed circular orbits (called energy levels or shells) without emitting energy.

2. Energy Emission/Absorption:

Electrons emit or absorb energy only when they jump from one allowed orbit to another.

The energy absorbed or emitted is equal to the difference between the two energy levels.

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

where h is Planck's constant and ν is the frequency of radiation.

3. Stable Orbits:

As long as electrons remain in their allowed orbits, the atom remains stable and does not radiate energy.

**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 : The term "**Quantum Number**" refers to a set of numerical values that describe the properties of electrons within an atom. These numbers provide information about the electron's energy, position, and angular momentum in its orbital. Quantum numbers are essential for defining the allowed states of an electron in an atom and for understanding the arrangement of electrons in atomic orbitals.

There are four quantum numbers associated with an electron in an orbital:

1. Principal Quantum Number (n):

- **Description:** This quantum number determines the **energy level** of an electron and the **size** of the orbital. It is a positive integer (1, 2, 3, ...), with higher values of **n** corresponding to orbitals that are farther from the nucleus and have higher energy.
- **Significance:**
 - Indicates the **energy** of the electron: higher **n** values correspond to higher energy levels.
 - Determines the **size** of the orbital: larger **n** means a larger orbital, with the electron being farther from the nucleus.
 - Defines the electron's overall **energy**.

2. Angular Momentum Quantum Number (l):

- **Description:** This quantum number defines the **shape** of the orbital and is related to the **angular momentum** of the electron. It can take integer values from 0 to **n - 1** for each value of **n**.
 - If **l = 0**, the orbital is spherical (**s** orbital).
 - If **l = 1**, the orbital is dumbbell-shaped (**p** orbital).
 - If **l = 2**, the orbital is cloverleaf-shaped (**d** orbital).
 - If **l = 3**, the orbital has a more complex shape (**f** orbital).

- **Significance:**
 - Defines the **shape** of the electron's probability distribution (orbital).
 - Determines the **type** of orbital (s, p, d, f).
- 3. **Magnetic Quantum Number (m_l):**
 - **Description:** This quantum number defines the **orientation** of the orbital in space relative to the other orbitals. It can take integer values from **-l** to **+l**, including zero.
 - For example, if **l = 1** (a p orbital), then **m_l** can take the values -1, 0, or +1, corresponding to the three possible orientations of a p orbital.
 - **Significance:**
 - Determines the **orientation** of the orbital in 3-dimensional space.
 - Specifies how many orbitals exist for a particular **l** value. For example, if **l = 1**, there are three possible orbitals (p_{-1} , p_0 , p_1).
- 4. **Spin Quantum Number (m_s):**
 - **Description:** This quantum number describes the **spin** of the electron, which is a fundamental property of the electron. The spin quantum number can take two values: **+1/2** or **-1/2**, representing the two possible spin states of an electron.
 - **Significance:**
 - Describes the **spin orientation** of the electron (clockwise or counterclockwise).
 - Ensures that no two electrons in the same orbital have the same set of quantum numbers (Pauli Exclusion Principle). Each orbital can hold a maximum of two electrons, one with spin +1/2 and one with spin -1/2.

Summary of the Four Quantum Numbers:

1. **Principal Quantum Number (n):** Specifies the energy level and size of the orbital.

2. **Angular Momentum Quantum Number (l):** Specifies the shape of the orbital.
3. **Magnetic Quantum Number (m_l):** Specifies the orientation of the orbital in space.
4. **Spin Quantum Number (m_s):** Specifies the spin of the electron (clockwise or counterclockwise).

Each of these quantum numbers provides crucial information about the location and behavior of an electron within an atom, helping to define its state within a given orbital.

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

Answer:

Comparison between Ionic and Covalent Compounds:

Property	Ionic Compounds	Covalent Compounds
Nature of Bond	Formed by transfer of electrons	Formed by sharing of electrons
Physical State	Mostly solids (hard and brittle)	Can be gases, liquids, or solids
Melting and Boiling Points	High	Generally low
Electrical Conductivity	Conduct electricity in molten or solution form	Usually do not conduct electricity
Solubility	Soluble in water	Soluble in organic solvents

Examples:

- **Ionic Compounds:** Sodium chloride (NaCl), Magnesium oxide (MgO)
- **Covalent Compounds:** Water (H_2O), Carbon dioxide (CO_2)

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

Answer : Co-ordinate Covalent Bond:

A **co-ordinate covalent bond** (also called a **dative bond**) is a type of covalent bond in which **both electrons** shared between two atoms come **from the same atom**.

Example: In the formation of ammonium ion (NH_4^+), the nitrogen atom donates a lone pair of electrons to a proton (H^+) to form a co-ordinate bond.

Difference between Normal Covalent Bond and Co-ordinate Covalent Bond:

Property	Normal Covalent Bond	Co-ordinate Covalent Bond
Electron Contribution	Each atom contributes one electron	One atom donates both electrons
Formation	By mutual sharing of electrons	By donation of a lone pair to an electron-deficient atom
Example	H_2 , O_2 , Cl_2 (normal covalent molecules)	NH_4^+ , CO , SO_2 (examples with co-ordinate bonds)

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

Answer : Hydrogen Bond:

A **hydrogen bond** is a special type of weak chemical bond that forms between a **hydrogen atom** (attached to a highly electronegative atom like nitrogen, oxygen, or fluorine) and another **electronegative atom** nearby.

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

Types of Hydrogen Bonds:

1. Intermolecular Hydrogen Bond:

- Forms between **different molecules**.
- Example: In water (H_2O), hydrogen bonds form between different water molecules.

2. Intramolecular Hydrogen Bond:

- Forms within the **same molecule**.
- Example: In ortho-nitrophenol, a hydrogen bond forms within the molecule itself.

Why does water have an abnormally high boiling point?

- In water (H_2O), strong **intermolecular hydrogen bonds** hold the molecules together.
- A large amount of heat energy is needed to break these hydrogen bonds before the water molecules can escape into the vapor phase.
- Therefore, water has an **abnormally high boiling point** compared to other similar-sized molecules.

7. (b) Why bond angles of H_2O and NH_3 are 104.5° and 107° respectively although central atoms are sp^3 hybridized?

Answer : Reason:

Both H_2O and NH_3 molecules have **sp^3 hybridization**, which ideally gives a bond angle of **109.5°** (as in perfect tetrahedral geometry). However, in both cases, the bond angles are **less than 109.5°** because of the presence of **lone pairs** on the central atom:

- In NH_3 (Ammonia):
 - Nitrogen has **one lone pair** and **three bond pairs**.
 - Lone pairs repel more strongly than bond pairs, slightly pushing the bond pairs closer.

- As a result, the bond angle reduces from 109.5° to about **107°** .
- In **H₂O** (Water):
 - Oxygen has **two lone pairs** and **two bond pairs**.
 - Two lone pairs exert even **greater repulsion** on the bond pairs compared to NH₃.
 - Thus, the bond angle decreases further to about **104.5°** .

Summary:

- **More lone pairs = More repulsion = Smaller bond angle.**

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 (Ionization Energy) refers to the energy required to remove an electron from an atom or ion in the gas phase. It is typically expressed in electron volts (eV) or kilojoules per mole (kJ/mol). The ionization potential is a measure of the ease with which an atom can lose an electron, and it is influenced by factors such as atomic size and the effective nuclear charge experienced by the electrons.

Why is the first ionization potential less than the second?

The first ionization potential refers to the energy needed to remove the first electron from an atom. The second ionization potential is the energy required to remove a second electron, after the first one has been removed.

The **first ionization potential is less than the second** because, after the first electron is removed, the remaining atom or ion has a greater effective nuclear charge (more protons relative to the number of electrons). This increased positive charge on the nucleus pulls the remaining electrons more tightly, making it harder to remove another

electron. As a result, the second ionization energy is higher than the first.

For example:

- **First ionization energy:** The removal of an electron from a neutral atom (e.g., $\text{Na} \rightarrow \text{Na}^+ + \text{e}^-$).
- **Second ionization energy:** The removal of an electron from a positively charged ion (e.g., $\text{Na}^+ \rightarrow \text{Na}^{2+} + \text{e}^-$).

Ionization Potential and Atomic Volume

The **ionization potential typically increases with decreasing atomic volume**. This is because as the atomic radius decreases (the atom becomes smaller), the electrons are closer to the nucleus and experience a stronger attractive force from the positively charged protons. This stronger attraction makes it harder to remove an electron, which leads to a higher ionization potential.

In summary:

- **First ionization energy < Second ionization energy** due to the increased effective nuclear charge after the first electron is removed.
- **Ionization potential increases as atomic volume decreases** because the electrons are held more tightly by the nucleus in smaller atoms.

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

Answer : f-block elements are a group of elements in the periodic table that are located in the two rows at the bottom of the table: the **lanthanides** (rare earth elements) and the **actinides**. These elements are characterized by the filling of their **f-orbitals** (the third set of orbitals in an atom, after s- and p-orbitals).

- **Lanthanides:** Elements with atomic numbers 57 to 71 (from lanthanum to lutetium).
- **Actinides:** Elements with atomic numbers 89 to 103 (from actinium to lawrencium).

Why are f-block elements called "inner transition elements"?

The term "**inner transition elements**" comes from their position in the periodic table and the way their electron configurations work.

- **Transition elements** are those that fill the d-orbitals, but the f-block elements fill the f-orbitals, which are in a different energy level than the d-orbitals.
- The f-block elements are considered "**inner**" **transition elements** because, unlike the d-block transition metals that fill their orbitals in the main body of the table, the f-block elements are placed in a separate section below the main body of the table.
- This placement reflects the fact that their electrons are added to the **f-orbitals**, which are farther from the nucleus and have different energy characteristics compared to the d-orbitals.

These elements are called **inner transition elements** because, although they are part of the transition series, their electron configurations involve inner orbitals (the f-orbitals), distinguishing them from the transition elements in the d-block.

In summary:

- **f-block elements** are the lanthanides and actinides, which fill the f-orbitals.
- They are called **inner transition elements** because they transition by filling the f-orbitals, which are part of the inner shells of electrons in the atom.