PABNA UNIVERSITY OF SCIENCE AND TECHNOLOGY

Faculty of Engineering & Technology

Department of Information and Communication Engineering

ASSIGNMENT ON QUESTIONS

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Question 01:

Give the defects of Rutherford's model of atom. What suggestions were given by Bohr to remove these defects?

Answer:

Defects of Rutherford's Atomic Model:

- **Instability of Electrons (According to Classical Physics):** Rutherford suggested that electrons revolve around the nucleus in circular orbits. However, according to classical electromagnetic theory, any charged particle undergoing acceleration (like an electron in circular motion) should emit radiation. As it loses energy, the electron would spiral inward and eventually collapse into the nucleus. This implies atoms should be unstable — but in reality, atoms are stable.
- **Failure to Explain Atomic Spectra:** Rutherford's model couldn't explain the discrete lines observed in atomic emission spectra, especially for hydrogen. If electrons were to lose energy continuously, the spectrum would be continuous — but experiments showed line spectra (specific wavelengths), which Rutherford's model failed to account for.

Bohr's Suggestions to Overcome These Defects:

- **Quantized Orbits:** Bohr proposed that electrons revolve around the nucleus in certain fixed orbits (called "stationary states") without emitting radiation. These orbits have quantized energy levels.
- **Energy Emission/Absorption Only During Transitions:** Electrons emit or absorb energy only when they jump from one allowed orbit to another. The energy absorbed or emitted corresponds to the difference between energy levels, and is given by the formula:
  $E=h\nu$
  where h is Planck's constant and ν is the frequency of the radiation.

Question 02:

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:

A **quantum number** is a value that describes the unique quantum state of an electron in an atom. An electron in an orbital is described by **four quantum numbers**.

Significance of Each Quantum Number:

1. **Principal Quantum Number (n):**
    - Indicates the **main energy level** or shell.
    - Determines the size and energy of the orbital.
    - Values: n=1,2,3,…n
2. **Azimuthal Quantum Number (l):**
    - Indicates the **subshell** or shape of the orbital.
    - Values: l=0 to n−1
      (0 = s, 1 = p, 2 = d, 3 = f)
3. **Magnetic Quantum Number (m or mlm_lml):**
    - Indicates the **orientation** of the orbital in space.
    - Values: ml=−l to +l
4. **Spin Quantum Number (s or msm_sms):**
    - Indicates the **spin** direction of the electron.

- Values: +1/2,-1/2

Question 03:

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

Answer:

The properties of ionic and covalente compounds are given below:

Properties	**Ionic Compounds**	**Covalent Compounds**
Bond Type	Transfer of electrons	Sharing of electrons
Physical State	Usually solid	Can be solid, liquid, or gas
Melting/Boiling Point	High	Low to moderate
Solubility	Soluble in water	Often insoluble in water
Electrical Conductivity	Conduct in molten/solution state	Usually do not conduct electricity

Examples:

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

Question 04:

What is a co-ordinate covalent bond? How does it differ from a normal covalent bond?

Answer:

A coordinate covalent bond is a type of covalent bond in which both electrons in the shared pair come from the same atom.

Difference from a Normal Covalent Bond:

Feature	Normal Covalent Bond	Coordinate Covalent Bond
Electron contribution	Each atom contributes one electron	One atom donates both electrons
Formation	Between atoms with similar electronegativity	Usually involves a lone pair donor and an acceptor
Example	H_2, O_2 (e.g. H–H, O=O)	NH_4^+, H_3O^+ (e.g. N → H^+)

Question 05:

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 special type of dipole-dipole interaction that occurs when a hydrogen atom, which is covalently bonded to a highly electronegative atom (like fluorine (F), oxygen (O), or nitrogen (N)), interacts with another electronegative atom having a lone pair of electrons.

Hydrogen bonding is weaker than covalent bonds but stronger than van der Waals forces.

Types of Hydrogen Bonds:

1. Intermolecular Hydrogen Bond:

- This occurs between two different molecules.
- Example: In water (H_2O), each molecule forms hydrogen bonds with neighboring molecules.

- This leads to the formation of a network of molecules, contributing to water's unique properties.

2. Intramolecular Hydrogen Bond:

- This occurs within the same molecule, usually between two functional groups.
- Example: In *ortho*-nitrophenol, a hydrogen bond forms between the hydroxyl group (-OH) and the nitro group ($-NO_2$) within the same molecule.

Water Has Abnormally High Boiling Point:

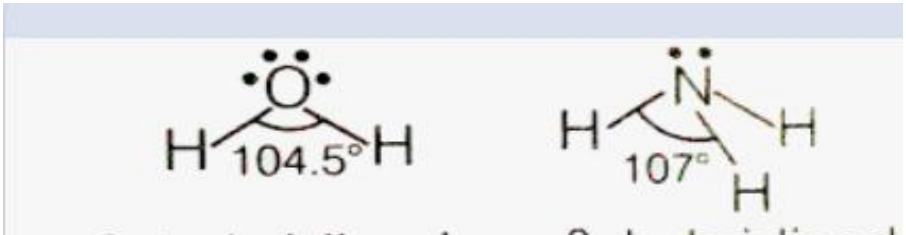
- Water molecules form extensive intermolecular hydrogen bonds due to the high electronegativity of oxygen and the presence of two hydrogen atoms per molecule.
- These hydrogen bonds strongly hold the molecules together, requiring a large amount of heat energy to break them and convert water into vapor.
- As a result, water has a much higher boiling point (100°C) than other similar-sized molecules like H_2S (boiling point: -60°C), which do not form hydrogen bonds.

Question 06:

Why bond angles of H2O and NH3 are 104.5° and 107° respectively although central atoms are sp3 hybridized.

Answer:

Water (H_2O) and ammonia (NH_3) both have central atoms (oxygen and nitrogen, respectively) that are sp^3 hybridized. According to the sp^3 hybridization theory, the electron pairs should arrange themselves in a tetrahedral geometry to minimize repulsion, with an ideal bond angle of 109.5°.



In the case of ammonia (NH_3), nitrogen has three bond pairs and one lone pair of electrons. The lone pair occupies more space than the bond pairs because it is only attracted to one nucleus instead of two. This increased repulsion pushes the bonding pairs slightly closer together, reducing the bond angle to 107° and giving ammonia a trigonal pyramidal shape.

For water (H_2O), oxygen has two bond pairs and two lone pairs. Lone pair–lone pair repulsion is stronger than lone pair–bond pair or bond pair–bond pair repulsion. As a result, the two lone pairs push the hydrogen atoms even closer, reducing the bond angle to 104.5°. This gives water a bent or V-shaped molecular geometry.

So, although both molecules are sp^3 hybridized, the presence and number of lone pairs cause variations in bond angles due to differing repulsion forces. The more lone pairs, the greater the repulsion, and the smaller the bond angle compared to the ideal 109.5°.

Question 07:

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 (or ionization energy) of an element is the minimum amount of energy required to remove the most loosely bound electron from an isolated gaseous atom in its ground state.

Mathematically:

$$M(g) + \text{Ionization energy} \rightarrow M^+(g) + e$$

It is measured in electron volts (eV) or kilojoules per mole (kJ/mol). This value indicates how strongly an atom holds onto its electrons.

the First Ionization Potential Less Than the Second because:

- The first ionization potential refers to the energy required to remove the first electron from a neutral atom.
- After removing this electron, the atom becomes a positively charged ion (M^+). This cation holds its remaining electrons more tightly due to:
    - Increased effective nuclear charge
    - Decreased electron-electron repulsion
- Therefore, removing a second electron requires more energy, making the second ionization potential higher than the first.

$$M(g) \rightarrow M^+(g) + e^-$$

$$M^+(g) \rightarrow M^+(g) + e^-$$

This trend continues; each subsequent ionization requires more energy than the previous one

Variation of Ionization Potential with Atomic Volume:

There is a clear trend in how ionization potential varies with atomic size (volume):

1. Across a Period (Left to Right):
    - Atomic size decreases
    - Nuclear charge increases
    - Electrons are more tightly held
    - Ionization potential increases
2. Down a Group (Top to Bottom):
    - Atomic size increases (more electron shells)
    - Outer electrons are farther from the nucleus
    - They are more easily removed
    - Ionization potential decreases

Thus, smaller atoms with higher nuclear attraction have higher ionization potentials, while larger atoms with loosely held outer electrons have lower ionization potentials.

Question 08:

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

Answer:

f-block elements are called inner transition elements because:

1. They lie between the s-block and d-block elements in terms of atomic number and electron configuration.
2. In the periodic table, they are placed separately below the main table, but they actually belong to the transition series, just like d-block elements.
3. The term "inner" indicates that the f-orbitals (which are deep within the atom) are being filled, unlike d-orbitals which are more outer in nature.
4. These elements show variable oxidation states, form colored ions, and often exhibit magnetic properties, similar to transition elements.

f-block elements are elements in which the last electron enters an f-orbital. They are known as inner transition elements because their electron transitions occur in the inner f-subshell, beneath the outermost energy levels.