PABNA UNIVERSITY OF SCIENCE AND TECHNOLOGY

পাবনা বিজ্ঞান ও প্রযুক্তি বিশ্ববিদ্যালয়

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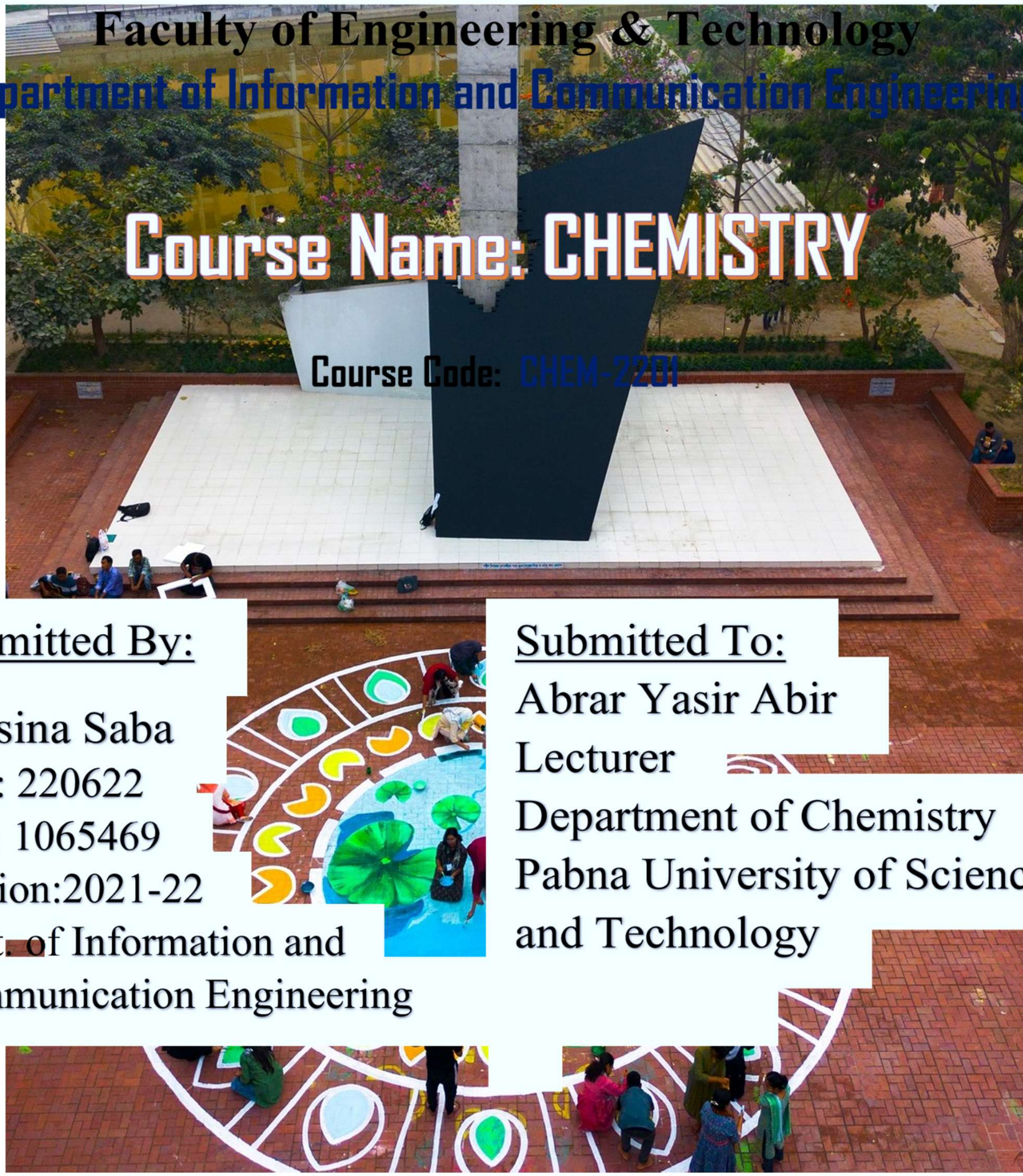

Faculty of Engineering & Technology

Department of Information and Communication Engineering

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Answer to the question no: 05

05(a):

Defects of Rutherford's Model of Atom:

1. **Stability Problem**: According to classical electromagnetic theory, an electron revolving around the nucleus should continuously emit energy in the form of radiation. As a result, the electron would spiral inward and eventually fall into the nucleus. But in reality, atoms are stable.
2. **Spectrum Problem**: Rutherford's model could not explain the discrete line spectra (distinct lines of specific wavelengths) observed for atoms like hydrogen. According to his model, the atom should emit a continuous spectrum, not discrete lines.

Bohr's Suggestions to Remove These Defects:

1. **Quantized Orbits**: Bohr proposed that electrons revolve around the nucleus only in certain fixed circular orbits called energy levels without emitting radiation. These orbits are associated with definite energies.
2. **Energy Absorption/Emission**: An electron can jump from one orbit to another by absorbing or emitting a fixed amount of energy, equal to the

difference in energies of the two orbits. This explained the discrete line spectra.

05(b):

Quantum numbers are numbers that describe the size, shape, orientation, and spin of an electron's orbit in an atom.

Each electron in an atom is completely described by four quantum numbers.

The Four Quantum Numbers and Their Significance:

Quantum Number	**Symbol**	**Significance**
1. Principal Quantum Number	**n**	Describes the main energy level or shell of the electron. It shows the size and energy of the orbital. (n = 1, 2, 3, ...)
2. Azimuthal (or Angular Momentum)	**l**	Describes the shape of the orbital. It depends on the value of **n**. (l = 0 to n−1). (s, p, d, f orbitals)

Quantum Number		
3. Magnetic Quantum Number	**m** (or **m_l**)	Describes the orientation of the orbital in space. ($m = -l$ to $+l$)
4. Spin Quantum Number	**s** (or **m_s**)	Describes the direction of the electron's spin: either +½ (clockwise) or −½ (anticlockwise).

Answer to the question no:06

06(a)

Comparison between Ionic and Covalent Compounds:

Property	**Ionic Compounds**	**Covalent Compounds**
Nature of Bond	Formed by the complete transfer of electrons from one atom to another.	Formed by the mutual sharing of electrons between two atoms.
Physical State	Generally exist as hard crystalline	May exist as solids, liquids, or gases.

Property	Ionic Compounds	Covalent Compounds
	solids at room temperature.	
Melting and Boiling Points	Have high melting and boiling points due to strong electrostatic forces.	Usually have low melting and boiling points as intermolecular forces are weaker.
Electrical Conductivity	Conduct electricity in molten state or in aqueous solution.	Generally do not conduct electricity, except some polar covalent compounds.
Solubility	Soluble in polar solvents like water.	Soluble in non-polar solvents like benzene and ether.

Examples:

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

06(b) :

Coordinate Covalent Bond:

A coordinate covalent bond is a type of covalent bond in which both the shared electrons are donated by one atom only.
In this bond, one atom having a lone pair donates it to another atom which is electron-deficient.

Example:
In the ammonium ion (NH_4^+), the nitrogen atom donates a lone pair of electrons to a hydrogen ion (H^+) to form a coordinate bond.

Difference between Coordinate Covalent Bond and Normal Covalent Bond:

Basis	**Coordinate Covalent Bond**	**Normal Covalent Bond**
Electron Contribution	Both electrons in the bond are provided by the same atom.	Each atom contributes one

Basis	Coordinate Covalent Bond	Normal Covalent Bond
		electron to the bond.
Formation	Forms between an atom with a lone pair and another atom needing electrons.	

Answer to the question no:07

07(a):

A **hydrogen bond** is a type of weak chemical bond that forms when a hydrogen atom, which is covalently bonded to a highly electronegative atom like nitrogen (N), oxygen (O), or fluorine (F), is attracted to another electronegative atom nearby.

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

Classification of Hydrogen Bonds:

1. **Intermolecular Hydrogen Bond:**
   This bond forms between hydrogen atoms of one molecule and electronegative atoms of another

molecule.
Example: Hydrogen bonding between two water (H_2O) molecules.

2. **Intramolecular Hydrogen Bond:**
   This bond forms within the same molecule when hydrogen is attracted to an electronegative atom in the same molecule.
   Example: In ortho-nitrophenol ($C_6H_4(OH)(NO_2)$), hydrogen bonding occurs within the molecule.

Water Has Abnormally High Boiling Point:

Water molecules are linked together by strong intermolecular hydrogen bonds.
To boil water, these hydrogen bonds must be broken, which requires a large amount of energy.
As a result, water has an unusually high boiling point compared to other molecules of similar size and mass.

07(b):

Reason for Different Bond Angles of H_2O and NH_3:

Both water (H_2O) and ammonia (NH_3) have sp^3 hybridization at their central atoms, but their bond angles are different because of the difference in the number of lone pairs.

- **In H_2O**:
  Oxygen has two lone pairs and two bond pairs. Lone pairs exert greater repulsion than bond pairs, which pushes the bond pairs closer together, reducing the bond angle to 104.5°.
- **In NH_3**:
  Nitrogen has one lone pair and three bond pairs. There is less lone pair repulsion compared to water, so the bond angle is slightly larger, around 107°.

Answer to the question no:08

08(a):

The ionization potential (or 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 positively charged ion.

In simple words, it measures how tightly an atom holds onto its outermost electron.

First Ionization Potential is Less than Second:

- The first ionization potential is the energy needed to remove the first electron.
- After removing the first electron, the atom becomes a positively charged ion.

- Removing a second electron from a positively charged ion requires more energy because the nucleus holds the remaining electrons more tightly due to increased effective nuclear charge.
- Therefore, the second ionization potential is higher than the first.

Variation of Ionization Potential with Atomic Volume:

- As atomic volume increases (i.e., size of the atom increases), the outermost electrons are farther from the nucleus.
- As a result, the attractive force between the nucleus and the outer electrons becomes weaker.
- Hence, ionization potential decreases with increase in atomic volume.
- Across a period (left to right): atomic volume decreases, so ionization potential increases.
- Down a group: atomic volume increases, so ionization potential decreases.

08(b):

f-block Elements:

The f-block elements are those elements in which the last electron enters the f-orbital of the atom.
They include the lanthanides (atomic numbers 58–71) and actinides (atomic numbers 90–103).

Why f-block Elements are Called Inner Transition Elements:

- In f-block elements, the f-orbitals, which are inner orbitals, are being filled.
- The filling of electrons takes place in an inner $(n-2)$ shell, even though outer shells are also present.
- Thus, the transition of electrons occurs in the inner shells and not the outermost shell.
- Therefore, they are called inner transition elements.