Chapter 6:

Stoichiometry

Short Questions (Exercise)

2(i) What is a mole?

A mole is a unit used to measure the amount of a substance. It is defined as the number of entities (atoms, ions, or molecules) in 12 grams of carbon12, which is approximately 6.022×10^{23} (Avogadro's number).

2(ii) Differentiate between empirical formula and molecular formula.

Empirical Formula: Represents the simplest wholenumber ratio of atoms in a compound (e.g., CH_2O).

Molecular Formula: Represents the actual number of atoms of each element in a compound (e.g., $C_6H_{12}O_6$).

2(iii) What is the number of molecules in 9.0 g of steam?

Molar mass of H_2O = 18 g/mol.
Moles of water = Mass / Molar Mass = 9.0 / 18 = 0.5 moles.
Number of molecules = $0.5 \times 6.022 \times 10^{23} = 3.011 \times 10^{23}$ molecules.

2(iv) What are the molar masses of uranium238 and uranium235?

Uranium238: 238 g/mol.
Uranium235: 235 g/mol.

2(v) Why are one mole of hydrogen molecules and one mole of Hatoms different in masses?

Hydrogen molecule (H_2) has two hydrogen atoms (Molar mass = 2 g/mol).
Hydrogen atom (H) has a molar mass of 1 g/mol.
Thus, one mole of H_2 weighs 2 g, while one mole of H weighs 1 g.

3. Define ion, molecular ion, formula unit, free radical, atomic number, mass number, atomic mass unit.

Ion: An atom or molecule with a net electric charge due to the loss or gain of electrons.
Molecular Ion: A charged species formed by the addition or removal of electrons from a molecule.
Formula Unit: The lowest wholenumber ratio of ions in an ionic compound (e.g., NaCl).
Free Radical: A molecule or atom with an unpaired electron, making it highly reactive.
Atomic Number: The number of protons in the nucleus of an atom.

Mass Number: The total number of protons and neutrons in the nucleus of an atom.
Atomic Mass Unit (amu): A unit of mass equal to 1/12 the mass of a carbon12 atom (1 amu = 1.66×10^{-24} g).

4. Describe how Avogadro's number is related to a mole of any substance.

Avogadro's Number (6.022×10^{23}) represents the number of atoms, ions, or molecules in one mole of a substance. For example, one mole of water contains 6.022×10^{23} water molecules.

5. Calculate the number of moles of each substance in the given masses:

2.4 g of He: Moles = 2.4 / 4 = 0.6 moles.
250 mg of carbon: Moles = 0.25 / 12 = 0.0208 moles.
15 g of sodium chloride (NaCl): Moles = 15 / 58.5 = 0.2564 moles.
40 g of sulphur (S): Moles = 40 / 32 = 1.25 moles.
1.5 kg of MgO: Moles = 1500 / 40 = 37.5 moles.

6. Calculate the mass in grams of:

1.2 moles of H_2: Mass = $1.2 \times 2 = 2.4$ g.
75 moles of H_2: Mass = $75 \times 2 = 150$ g.
0.25 moles of steam (H_2O): Mass = $0.25 \times 18 = 4.5$ g.
1.05 moles of $CuSO_4 \cdot 5H_2O$: Mass = $1.05 \times 249 = 261.45$ g.
1.5 moles of H_2SO_4: Mass = $1.5 \times 98 = 147$ g.

7. Identify the substance that has formula mass of 133.5 amu:

Substance $AlCl_3$ has a formula mass of 133.5 amu. Calculations:
- (a) $MgCl_2$: 95.3 amu
- (b) S_2Cl_2: 135.2 amu
- (c) BCl_3: 117.3 amu
- (d) $AlCl_3$: 133.5 amu

8. Calculate the number of atoms in each of the following samples:

- (a) 3.4 moles of nitrogen atoms: 2.048×10^{24} atoms.
- (b) 23 g of Na: 6.022×10^{23} atoms.
- (c) 5 g of H-atoms: 3.011×10^{24} atoms.

9. Calculate the mass of the following:

- (a) 3.24×10^{20} atoms of Fe: 3.01×10^{-2} g.
- (b) 2×10^{20} molecules of N_2: 9.29×10^{-3} g.
- (c) 1.1×10^{4} molecules of H_2O: 3.29×10^{-20} g.
- (d) 3×10^{4} atoms of Al: 1.35×10^{-19} g.

10. Balance the following chemical equations:

- (a) $Na + H_2O \rightarrow NaOH + H_2$:
  Balanced: $2Na + 2H_2O \rightarrow 2NaOH + H_2$.
- (b) $NH_3 + HCl \rightarrow NH_4Cl$:
  Balanced: $NH_3 + HCl \rightarrow NH_4Cl$.

11. Potassium reacts with water:

(a) Formula of potassium oxide and potassium nitride:
- Potassium oxide: K_2O.

- Potassium nitride: K_3N.

(b) Reaction of 40.5 g K with 100 cm^3 of water:
1. Balanced equation: $2K + 2H_2O \rightarrow 2KOH + H_2$.
2. Ionic equation: $2K + 2H_2O \rightarrow 2K^+ + 2OH^- + H_2$.
3. Number of atoms of K: 6.23×10^{23} atoms (from 1.035 mol of K).
4. Period number of potassium: Period 4.

Extera Short Questions (Topic Wise)

6.1: Empirical Formula and Molecular Formula

1. **What is an empirical formula?**
   The empirical formula shows the simplest whole-number ratio of atoms in a compound.
   Example: CH_2O is the empirical formula for glucose.

2. **What is a molecular formula?**
   The molecular formula shows the actual number of atoms in a molecule.
   Example: Glucose has a molecular formula of $C_6H_{12}O_6$.

3. **How are empirical and molecular formulas related?**
   The molecular formula is a multiple of the empirical formula.
   Example: C_2H_4 is twice CH_2.

6.2: Molecular Mass and Formula Mass

1. **What is molecular mass?**
   The molecular mass is the sum of the atomic masses of all atoms in a molecule.
   Example: H_2O has a molecular mass of 18 u ($2 \times 1 + 16$).

2. **What is formula mass?**
   Formula mass is the sum of atomic masses in an ionic compound.
   Example: NaCl has a formula mass of 58.5 u (23 + 35.5).

3. **How do molecular and formula mass differ?**
   Molecular mass applies to covalent compounds, while formula mass is for ionic compounds.

6.3: Chemical Formula and Name of Binary Ionic Compounds

1. **What is a binary ionic compound?**
   A compound formed between two elements, typically a metal and a non-metal.
   Example: NaCl (sodium chloride).

2. **How do you name binary ionic compounds?**
   Use the metal name first, followed by the non-metal with an "-ide" ending.
   Example: MgO is magnesium oxide.

3. **What is the chemical formula of potassium bromide?**
   KBr.

6.4: Avogadro's Number and Mole

1. **What is Avogadro's number?**
   Avogadro's number is 6.022×10^{23}, representing the number of particles in one mole of a substance.

2. **What is a mole?**
   A mole is the amount of substance containing Avogadro's number of particles.
   Example: 1 mole of carbon atoms is 6.022×10^{23} atoms.

3. **How is a mole used in calculations?**
   It relates the mass of a substance to the number of particles or moles.
   Example: 12 g of carbon equals 1 mole.

6.5: Chemical Calculations

1. **How do you calculate the molar mass of a compound?**
   Add the atomic masses of all atoms in the compound.
   Example: CO_2 has a molar mass of 44 g/mol (12 + 16 × 2).

2. **How do you calculate moles from mass?**
   Use the formula:
   Moles=MassMolar Mass\text{Moles} = \frac{\text{Mass}}{\text{Molar Mass}}Moles=Molar MassMass
   Example: 22 g of CO_2 equals 2244=0.5\frac{22}{44} = 0.54422=0.5 moles.

3. **What is percentage composition?**
   The percentage by mass of each element in a compound.
   Example: H_2O is 11.1% H and 88.9% O.

6.6: Chemical Equations and Balancing

1. **What is a chemical equation?**
   A representation of a chemical reaction using symbols and formulas.
   Example: $H_2 + O_2 \rightarrow H_2O$.

2. **Why is balancing chemical equations important?**
   It ensures the conservation of mass and atoms during a reaction.

3. **How do you balance a chemical equation?**
   Adjust coefficients to equalize the number of atoms on both sides.
   Example: 2H2+O2→2H2O2\text{H}_2 + \text{O}_2 → 2\text{H}_2\text{O}2H2+O2→2H2O.

6.7: Molecular and Structural Formula

1. **What is a molecular formula?**
   A formula showing the exact number of atoms of each element in a molecule.
   Example: C_2H_6 for ethane.

2. **What is a structural formula?**
   A formula showing the arrangement of atoms within a molecule.
   Example: Ethane's structural formula is H3C−CH3\text{H}_3\text{C}-\text{CH}_3H3C−CH3.

3. **How do structural formulas differ from molecular formulas?**
   Structural formulas provide spatial information, while molecular formulas show the quantity of atoms.

Extera Long Questions (Topic Wise)

1. What is the difference between empirical and molecular formulas, and how are they related?

Empirical Formula
- The empirical formula represents the simplest whole-number ratio of atoms in a compound.
- It does not provide information about the exact number of atoms in a molecule, only their ratio.
- Example: The empirical formula of glucose is CH_2O, indicating a 1:2:1 ratio of carbon, hydrogen, and oxygen atoms.

Molecular Formula
- The molecular formula shows the actual number of atoms of each element in a molecule.
- It is a multiple of the empirical formula.
- Example: The molecular formula of glucose is $C_6H_{12}O_6$, which is six times the empirical formula CH_2O.

Relationship Between Empirical and Molecular Formulas
- The molecular formula can be derived by multiplying the empirical formula by a factor determined from the compound's molecular mass and empirical formula mass.
  Molecular Formula = (Empirical Formula) × (Molecular Mass / Empirical Formula Mass)
- Example Calculation: A compound has an empirical formula CH and molecular mass 78 g/mol.
  The empirical formula mass is 12 + 1 = 13 g/mol.
  The factor is 78 / 13 = 6.
  Molecular formula = CH × 6 = C_6H_6 (benzene).

2. What is Avogadro's number, and how is it used in chemical calculations?

Avogadro's Number
- Avogadro's number is 6.022×10^{23}, representing the number of particles (atoms, molecules, ions) in one mole of a substance.
- It provides a link between the microscopic scale (individual atoms or molecules) and the macroscopic scale (grams).

Significance in Chemistry
1. Counting Particles: One mole of any substance contains 6.022×10^{23} entities.
   Example: 1 mole of H_2O contains 6.022×10^{23} water molecules.

2. Mass-Particle Relationship: The molar mass of a substance (in grams) corresponds to one mole of particles.
   Example: 1 mole of oxygen gas (O_2) has a molar mass of 32 g and contains 6.022×10^{23} molecules.

Applications in Calculations
1. Converting Moles to Particles:

Number of Particles = Moles × 6.022×10^{23}
Example: 2 moles of H_2O = $2 \times 6.022 \times 10^{23} = 1.204 \times 10^{24}$ molecules.

2. Converting Particles to Moles:
Moles = Number of Particles / 6.022×10^{23}
Example: 1.204×10^{24} molecules of H_2O = $1.204 \times 10^{24} / 6.022 \times 10^{23} = 2$ moles.

3. Relating Mass and Particles:
First convert mass to moles using molar mass, then convert moles to particles.
Example: 18 g of H_2O = 18 / 18 = 1 mole = 6.022×10^{23} molecules.

3. How do you write chemical formulas for binary ionic compounds, and what rules govern their naming?

Binary Ionic Compounds
- Binary ionic compounds consist of two elements: a metal (cation) and a non-metal (anion).

Writing Chemical Formulas
1. Determine the Charges: Identify the charges of the cation and anion based on their group numbers.
Example: Sodium (Na^+) and chlorine (Cl^-).

2. Balance the Charges: Combine ions in a ratio that makes the compound electrically neutral.
Example: $Na^+ + Cl^- \rightarrow NaCl$ (1:1 ratio).

3. Use Subscripts if Necessary: Use subscripts to indicate the number of ions needed to balance charges.
Example: Magnesium (Mg^{2+}) and oxygen (O^{2-}) → MgO.
Example: Aluminum (Al^{3+}) and oxygen (O^{2-}) → Al_2O_3.

Naming Rules
1. Name the Metal First: Use the element name for the metal.
Example: Na = sodium.

2. Name the Non-Metal with an "-ide" Ending: Change the non-metal's suffix to "-ide."
Example: Cl = chloride.

Examples of Binary Ionic Compounds:
- NaCl = Sodium chloride.
- MgO = Magnesium oxide.
- Al_2O_3 = Aluminum oxide.

4. Why is it important to balance chemical equations, and how is it done?

Importance of Balancing Chemical Equations
- Balancing ensures the law of conservation of mass is followed, meaning the number of atoms of

each element is the same on both sides of the reaction.
- It represents real-world chemical processes accurately, ensuring proper stoichiometric ratios for calculations.

Steps to Balance a Chemical Equation

1. Write the Unbalanced Equation:
   Example: $H_2 + O_2 \rightarrow H_2O$.

2. Count Atoms on Both Sides:
   Reactants: H = 2, O = 2. Products: H = 2, O = 1.

3. Balance One Element at a Time:
   Adjust coefficients to balance oxygen: $H_2 + O_2 \rightarrow 2H_2O$.
   Reactants: H = 2, O = 2. Products: H = 4, O = 2.

4. Balance Remaining Elements:
   Adjust hydrogen by adding a coefficient: $2H_2 + O_2 \rightarrow 2H_2O$.

5. Verify Atom Counts:
   Reactants: H = 4, O = 2. Products: H = 4, O = 2.

Examples of Balanced Equations

1. Combustion of methane: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$.
2. Formation of water: $2H_2 + O_2 \rightarrow 2H_2O$.
3. Reaction of aluminum with oxygen: $4Al + 3O_2 \rightarrow 2Al_2O_3$.

MCQS

6.1: Empirical Formula and Molecular Formula

1. **What does the empirical formula represent?**
   a) Actual number of atoms in a molecule
   b) Simplest whole-number ratio of atoms ☑
   c) Mass of the compound
   d) Number of moles in a reaction

2. **Which formula represents the exact composition of a compound?**
   a) Empirical formula
   b) Molecular formula ☑
   c) Structural formula
   d) Chemical formula

3. **What is the empirical formula of glucose ($C_6H_{12}O_6$)?**
   a) $C_6H_{12}O_6$
   b) CHO
   c) CH_2O ☑
   d) $C_3H_6O_3$

4. **How is the molecular formula related to the empirical formula?**
   a) It is a multiple of the empirical formula ☑
   b) It is always simpler
   c) It is unrelated
   d) It contains fewer atoms

6.2: Molecular Mass and Formula Mass

5. **What is molecular mass?**
   a) Mass of one atom
   b) Weighted average of isotopes
   c) Sum of atomic masses in a molecule ☑
   d) Mass of ions in a compound

6. **What is formula mass used for?**
   a) Calculating the mass of ionic compounds ☑
   b) Determining atomic numbers
   c) Measuring molecular bonds
   d) Representing nonpolar compounds

7. **What is the molecular mass of H_2O?**
   a) 18 u
   b) 16 u

c) 18 u (2 × 1 + 16) ✅

d) 10 u

8. **Which term is used for covalent compounds?**

a) Formula mass

b) Molecular mass ✅

c) Atomic mass

d) Mass number

6.3: Chemical Formula and Name of Binary Ionic Compounds

9. **What is a binary ionic compound?**

a) A compound with two nonmetals

b) A compound with three elements

c) A compound with one metal and one non-metal ✅

d) A compound with multiple ions

10. **What is the formula for sodium chloride?**

a) $NaCl$ ✅

b) Na_2Cl

c) $NaCl_2$

d) $ClNa_2$

11. **What is the name of MgO?**

a) Magnesium dioxide

b) Magnesium oxide ✅

c) Magnesium hydroxide

d) Magnesium chloride

12. **How are binary ionic compounds named?**

a) Nonmetal first, metal second

b) Use "ide" for the metal

c) Prefix for the metal's charge

d) Metal name first, nonmetal with "ide" ✅

6.4: Avogadro's Number and Mole

13. **What is Avogadro's number?**

a) 3.14×10^{23}

b) 6.022×10^{23} ✅

c) 9.81×10^{23}

d) 1.66×10^{23}

14. **How many particles are in one mole of a substance?**

a) 1×10^{23}

b) 6.022×10^{23} ✅

c) 5.55×10^{23}

d) 3.14×10^{23}

15. **What is the molar mass of CO_2?**
    a) 28 g/mol
    b) 44 g/mol (12 + 16 × 2) ☑
    c) 32 g/mol
    d) 18 g/mol

16. **How is a mole defined?**
    a) Number of particles equal to Avogadro's number ☑
    b) Volume of a gas at STP
    c) Weight of one molecule
    d) Ratio of protons to neutrons

6.5: Chemical Calculations

17. **What is the formula to calculate moles?**
    a) Moles = Mass × Molar Mass
    b) Moles = Mass × Volume
    c) Moles = Mass ÷ Molar Mass ☑
    d) Moles = Volume ÷ Molar Mass

18. **How many moles are in 36 g of H_2O?**
    a) 1 mole
    b) 2 moles (36 ÷ 18) ☑
    c) 3 moles
    d) 4 moles

19. **What is percentage composition?**
    a) Total mass of a molecule
    b) Percentage by mass of each element in a compound ☑
    c) Ratio of molecules in a reaction
    d) Number of electrons in a compound

20. **What is the percentage composition of H in H_2O?**
    a) 11.1% ☑
    b) 22.2%
    c) 50%
    d) 88.9%

6.6: Chemical Equations and Balancing

21. **Why must chemical equations be balanced?**
    a) To equalize reactants and products
    b) To match empirical formulas
    c) To obey the law of conservation of mass ☑
    d) To determine molar mass

22. **What is the balanced equation for water formation?**
    a) $2H_2 + O_2 \rightarrow 2H_2O$ ☑

b) $H_2 + O_2 \rightarrow H_2O$
c) $2H_2 + 2O_2 \rightarrow H_2O$
d) $H_2 + 2O_2 \rightarrow H_2O_2$

23. **How do coefficients balance equations?**
a) By adding atoms
b) By adjusting atom counts on both sides ☑
c) By rearranging molecules
d) By removing elements

24. **What type of reaction is $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$?**
a) Combustion ☑
b) Decomposition
c) Synthesis
d) Displacement

6.7: Molecular and Structural Formula

25. **What is a molecular formula?**
a) Exact number of atoms in a molecule ☑
b) Simplest ratio of elements
c) Structural arrangement of atoms
d) Weight of a molecule

26. **What is the molecular formula of ethane?**
a) CH_4
b) C_2H_6 ☑
c) C_2H_4
d) C_3H_8

27. **What is a structural formula?**
a) Number of atoms in a compound
b) Simplest atomic ratio
c) Arrangement of atoms in a molecule ☑
d) Mass of atoms in a molecule

28. **How does a structural formula differ from a molecular formula?**
a) It is simpler
b) It shows how atoms are arranged ☑
c) It uses more atoms
d) It includes the empirical ratio

29. **What is the molecular formula of glucose?**
a) $C_6H_{12}O_6$ ☑
b) CH_2O
c) C_2H_6O
d) C_6H_6

30. **What is the structural formula of ethane?**

a) C=C

b) H_3C-CH_3 ✓

c) $H_2C=CH_2$

d) C_3H_8