Summary and Notes for Exam Preparation: Moles and Equations

Key Terms and Concepts

1. **Relative Atomic Mass (Ar):**

- Definition: The weighted average mass of an atom of an element compared to 1/12th of the mass of a carbon-12 atom.

- Calculation:

- Importance: Used in chemical calculations to compare the masses of different atoms.

2. **Isotopic Mass**:

- Definition: The mass of a specific isotope of an element on a scale where carbon-12 has a mass of exactly 12 units.

- Example: The relative isotopic mass of carbon-13 is 13.00.

3. **Relative Molecular Mass (Mr):**

- Definition: The relative mass of one molecule of a compound compared to 1/12th of the mass of a carbon-12 atom.

- Calculation: Sum of the relative atomic masses of all atoms in the molecule.

- Example: Methane (CH_4)

4. **Relative Formula Mass:**

- Definition: Used for ionic compounds, calculated the same way as relative molecular mass.

- Example: Magnesium hydroxide $Mg(OH)_2$

5. **The Mole and Avogadro Constant:**

- Definition: A mole represents particles (atoms, molecules, ions, etc.), known as the Avogadro constant.

- Importance: Fundamental unit in chemistry for quantifying substances.

Analyzing Mass Spectra

- Mass Spectrometer: Instrument used to measure the mass of isotopes and their relative abundances.

- Interpreting Mass Spectra:

  - The mass spectrum shows relative abundance on the vertical axis and mass-to-charge ratio (m/e) on the horizontal axis.

  - Example: Lead isotopes

    - Isotopic masses: 204, 206, 207, 208

    - Relative abundances: 2%, 24%, 22%, 52%

  - Calculation: Relative atomic mass of lead

Calculating Empirical and Molecular Formulae

- Empirical Formula:

  - Simplest whole-number ratio of atoms in a compound.

  - Example: From combustion data or mass composition.

- Molecular Formula:

  - Actual number of atoms of each element in a molecule.

- Derived from empirical formula and molar mass.

Writing and Balancing Chemical Equations

- Balanced Equations:
  - Ensure the same number of each type of atom on both sides of the reaction.
  - Example: Combustion of methane

Calculations Using the Mole Concept

1. Reacting Masses:
   - From chemical formulae and balanced equations.
   - Example: Calculate the mass of CO_2 produced from 16g of CH_4.
2. Volumes of Gases:
   - Use the ideal gas law or molar volume at STP.
   - Example: Volume of CO_2 produced from burning hydrocarbons.
3. Volumes and Concentrations of Solutions:
   - Use molarity calculations.
   - Example: Calculate the concentration of a solution given the volume and moles of solute.

Stoichiometric Relationships

- Deducing Stoichiometry:
  - From reacting masses, gas volumes, and solution concentrations.

- Example: Determine the ratio of reactants and products in a reaction based on experimental data.

Practice Problems

1. Calculate the relative formula masses:

- **Calcium chloride ($CaCl_2$)**

Mr of $CaCl_2 = (1 \times 40.1) + (2 \times 35.5) = 111.1$

- **Copper(II) sulfate ($CuSO_4$)**

Mr of $CuSO_4 = (1 \times 63.5) + (1 \times 32.1) + (4 \times 16.0) = 159.6$

- **Ammonium sulfate $(NH_4)_2SO_4$**

Mr of $(NH_4)_2SO_4 = (2 \times (14.0 + 4 \times 1.0)) + (1 \times 32.1) + (4 \times 16.0) = 132.1$

- **Magnesium nitrate-6-water $Mg(NO_3)_2.6H_2O$**

Mr of $Mg(NO_3)_2 = (1 \times 24.3) + (2 \times (14.0 + 3 \times 16.0)) = 148.3$

Mr of $6H_2O = 6 \times (2 \times 1.0 + 16.0) = 108.0$

Mr of $Mg(NO_3)_2.6H_2O = 148.3 + 108.0 = 256.3$

Mass Spectrometry and Relative Atomic Mass (Ar)

Definition and Significance

Mass Spectrometry:

An analytical technique that ionizes chemical species and separates ions based on their mass-to-charge ratio (m/z).

Relative Atomic Mass (Ar): The weighted average mass of an atom compared to 1/12th the mass of carbon-12.

Significance: Essential for:

- Determining atomic masses accurately
- Identifying unknown compounds
- Analyzing isotopic composition
- Quality control in industry

Mass Spectrometer Components

- Ionization Chamber: Creates positively charged ions

- Acceleration Region: Accelerates ions using electric field
- Deflection Chamber: Separates ions based on mass/charge ratio
- Detector: Measures relative abundance of each ion

Detailed Calculations with Mass Spectra

Example 1:

Simple Two-Isotope System

Magnesium has three isotopes:

Mg-24 (abundance 78.99%)

Mg-25 (abundance 10.00%)

Mg-26 (abundance 11.01%)

Ar = (24 × 78.99 + 25 × 10.00 + 26 × 11.01) ÷ 100

= (1895.76 + 250.00 + 286.26) ÷ 100

= 24.32

Example 2:

Complex Isotope Pattern

Calculate Ar of silver:

Ag-107 (abundance 51.839%)

Ag-109 (abundance 48.161%)

Ar = (107 × 51.839 + 109 × 48.161) ÷ 100

= (5546.773 + 5249.549) ÷ 100

= 107.96

2. The Mole Concept and Avogadro Constant

Fundamental Concepts

The Mole: The amount of substance containing as many elementary entities as there are atoms in 12g of carbon-12

Avogadro Constant (NA): $6.02214076 \times 10^{23}$ mol^{-1}

Molar Mass: Mass of one mole of substance in grams

Significance

- Bridges microscopic (atoms) and macroscopic (visible quantities) worlds
- Essential for all chemical calculations
- Fundamental to stoichiometry
- Basis for concentration calculations

Detailed Examples

Example 3:

Multi-Step Mole Calculations

Calculate mass of NH_3 containing 1.5×10^{23} molecules

Step 1: Convert molecules to moles

Moles = $1.5 \times 10^{23} \div (6.02 \times 10^{23})$

= 0.249 moles

Step 2: Convert moles to mass

Mass = moles × molar mass

= 0.249 × 17

= 4.23g NH_3

Example 4:

Complex Particle Problems

Calculate total number of atoms in 5.0g of CH_4

Step 1: Find moles of CH_4

Moles = 5.0g ÷ 16g/mol = 0.3125 moles

Step 2: Calculate atoms per molecule

Atoms per CH_4 = 1(C) + 4(H) = 5 atoms

Step 3: Calculate total atoms

Total atoms = $0.3125 \times (6.02 \times 10^{23}) \times 5$

= 9.41×10^{23} atoms

3. Stoichiometry and Reacting Masses

Definition and Significance

Stoichiometry: The quantitative relationship between reactants and products

Importance:

- Essential for industry and manufacturing
- Critical for efficiency in chemical processes
- Key to environmental chemistry
- Fundamental to analytical chemistry

Advanced Stoichiometric Calculations

Example 5:

Limiting Reagent Problems

$2Al + 3CuO \rightarrow Al_2O_3 + 3Cu$

Given: 5.4g Al and 12g CuO

Step 1: Calculate moles of each reactant

Al: 5.4g ÷ 27g/mol = 0.2 moles

CuO: 12g ÷ 79.5g/mol = 0.151 moles

Step 2: Compare with equation ratios

Al needed: 2 mol

CuO needed: 3 mol

Ratio Al:CuO should be 2:3

Actual ratio = 0.2:0.151 = 2:1.51

Therefore, CuO is limiting reagent

Example 6:

Percentage Yield Problems

Calculate percentage yield when 15.0g Fe_2O_3 produces 9.8g Fe

Equation: $2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$

Step 1: Calculate theoretical yield

Moles Fe_2O_3 = 15.0g ÷ 159.7g/mol = 0.0939 mol

Moles Fe theoretical = 0.0939 × 4/2 = 0.1878 mol

Mass Fe theoretical = 0.1878 × 55.85 = 10.49g

Step 2: Calculate percentage yield

Yield = (9.8 ÷ 10.49) × 100 = 93.4%

4. Empirical and Molecular Formulae

Definitions

Empirical Formula: Simplest whole number ratio of atoms

Molecular Formula: Actual number of atoms in molecule

Relationship: Molecular formula = (Empirical formula) x n

Detailed Examples

Example 7:

Complex Empirical Formula

Compound contains:

C = 40.0%, H = 6.67%, O = 53.33%

Step 1: Assume 100g sample

C: 40.0g ÷ 12 = 3.33 mol

H: 6.67g ÷ 1 = 6.67 mol

O: 53.33g ÷ 16 = 3.33 mol

Step 2: Divide by smallest

C: 3.33 ÷ 3.33 = 1

H: 6.67 ÷ 3.33 = 2

O: 3.33 ÷ 3.33 = 1

Empirical formula = CH_2O

Example 8:

Molecular Formula Determination

Empirical formula = CH_2O

Molecular mass = 180g/mol

Empirical mass = 30g/mol

n = 180 ÷ 30 = 6

Molecular formula = $(CH_2O)_6 = C_6H_{12}O_6$

Practice Problems Set

Set 1: Mass Spectrometry

Calculate Ar for gallium:

Ga-69 (60.108%)

Ga-71 (39.892%)

A sample of lithium contains:

Li-6 (7.5%)

Li-7 (92.5%)

Calculate Ar and verify against periodic table value.

Set 2: Mole Calculations

Calculate:

a) Number of water molecules in 27g H_2O

b) Mass of 3.01×10^{23} CO_2 molecules

c) Number of atoms in 5.85g NaCl

For $CuSO_4 \cdot 5H_2O$:

a) Calculate molar mass

b) Find mass containing 6.02×10^{23} Cu^{2+} ions

c) Calculate water mass in 24.9g of hydrated salt

Set 3: Stoichiometry

In the reaction: $2KMnO_4 + 16HCl \rightarrow 2KCl + 2MnCl_2 + 8H_2O + 5Cl_2$

Calculate:

a) Mass of Cl_2 from 15.8g $KMnO_4$

b) Volume of Cl_2 at RTP (24dm³/mol)

c) Mass of HCl needed

For the reaction: $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

If 68g NH3 reacts with 160g O2:

a) Identify limiting reagent

b) Calculate maximum NO produced

c) Find mass of excess reagent remaining

- **Set 4: Empirical Formula**

A compound contains:

C = 54.55%, H = 9.09%, O = 36.36%

If molecular mass = 88g/mol, find:

a) Empirical formula

b) Molecular formula

Combustion of 2.34g organic compound produces:

6.82g CO2 and 2.79g H2O

Calculate empirical formula.

Chemical Formulae, Equations, and Solutions

1. Chemical Formulae and Deducing the Formula

Defining Chemical Formulae

- **Empirical Formula**: The simplest whole number ratio of atoms in a compound.
- **Molecular Formula**: The actual number of atoms of each element in a molecule.
- **Relationship**: Molecular formula = (Empirical formula)n, where n is an integer.

Determining Empirical Formulae

Example 1:

Finding Empirical Formula from Percentage Composition A compound contains 40.0% C, 6.67% H, and 53.33% O by mass.

Step 1: Assume 100 g sample

Moles of C = 40.0 g / 12.01 g/mol = 3.33 mol

Moles of H = 6.67 g / 1.008 g/mol = 6.67 mol

Moles of O = 53.33 g / 16.00 g/mol = 3.33 mol

Step 2: Divide moles by smallest value (3.33)

C: 3.33 / 3.33 = 1

H: 6.67 / 3.33 = 2

O: 3.33 / 3.33 = 1

Therefore, the empirical formula is CH_2O.

Example 2:

Determining Empirical Formula from Combustion Data A 2.34 g sample of an organic compound produced 6.82 g of CO_2 and 2.79 g of H_2O upon combustion.

Step 1:

Calculate moles of C and H

Moles of C = (6.82 g CO_2 × 12.01 g/mol C) / (44.01 g/mol CO_2) = 1.84 mol C

Moles of H = (2.79 g H_2O × 2.02 g/mol H) / (18.02 g/mol H_2O) = 3.08 mol H

Step 2: Find the simplest whole number ratio C: 1.84 mol H: 3.08 mol

Divide both by the smaller value (1.84) C: 1.00 H: 1.67 ≈ 2

Therefore, the empirical formula is CH_2.

Determining Molecular Formulae

Example 3:

Calculating Molecular Formula from Empirical Formula and Molar Mass The empirical formula of a compound is CH_2O, and its molar mass is 180 g/mol.

Step 1:

Calculate the empirical formula mass CH_2O has a mass of 30 g/mol

Step 2:

Determine the value of n Molecular mass / Empirical mass = n 180 g/mol / 30 g/mol = 6

Therefore, the molecular formula is $(CH_2O)_6$ = $C_6H_{12}O_6$.

2. Balancing Chemical Equations

Key Principles of Balancing Equations

1. The same number of atoms of each element must appear on both sides of the equation.
2. Coefficients are used to balance the equation, not subscripts.
3. Balanced equations must follow the law of conservation of mass.

Example 4:

Balancing a Simple Combustion Reaction Balanced the equation for the combustion of methane (CH_4) with oxygen (O_2) to form carbon dioxide (CO_2) and water (H_2O).

Unbalanced equation: $CH_4 + O_2 \rightarrow CO_2 + H_2O$

Step 1: Count atoms on each side Left side: 1 C, 4 H, unknown O Right side: 1 C, 2 H, 3 O

Step 2: Balance C and H atoms $CH_4 + O_2 \rightarrow CO_2 + 2H_2O$

Step 3: Balance O atoms $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

The balanced equation is: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Example 5:

Balancing a Redox Reaction Balance the equation for the reaction between potassium permanganate ($KMnO_4$) and hydrochloric acid (HCl) to form manganese(II) chloride ($MnCl_2$), potassium chloride (KCl), and chlorine gas (Cl_2).

Unbalanced equation: $KMnO_4 + HCl \rightarrow MnCl_2 + KCl + Cl_2$

Step 1: Identify oxidation and reduction half-reactions Oxidation: $2Cl^- \rightarrow Cl_2 + 2e^-$ Reduction: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

Step 2: Multiply half-reactions to balance electrons $10Cl^- \rightarrow 5Cl_2 + 10e^-$ $2MnO_4^- + 16H^+ + 10e^- \rightarrow 2Mn^{2+} + 8H_2O$

Step 3: Add half-reactions and simplify $2MnO_4^- + 16HCl \rightarrow 2MnCl_2 + 8H_2O + 5Cl_2$

Step 4: Add KCl to balance potassium $2KMnO_4 + 16HCl \rightarrow 2MnCl_2 + 2KCl + 8H_2O + 5Cl_2$

The balanced equation is: $2KMnO_4 + 16HCl \rightarrow 2MnCl_2 + 2KCl + 8H_2O + 5Cl_2$

3. Balancing Ionic Equations

Introduction to Ionic Equations

- Ionic equations represent the dissociation of ionic compounds in aqueous solutions.
- They focus on the participation of ions in the reaction rather than the overall compound.
- Balancing ionic equations follows the same principles as balancing molecular equations.

Example 6:

Balancing a Simple Ionic Equation Write the balanced ionic equation for the reaction between aqueous silver nitrate ($AgNO_3$) and aqueous sodium chloride (NaCl) to form solid silver chloride (AgCl) and aqueous sodium nitrate ($NaNO_3$).

Step 1: Write the unbalanced ionic equation: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

Step 2: Balance the equation: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

The balanced ionic equation is: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

Example 7:

Balancing a Complex Ionic Equation Write the balanced ionic equation for the reaction between aqueous lead(II) nitrate ($Pb(NO_3)_2$) and aqueous sodium hydroxide (NaOH) to form solid lead(II) hydroxide ($Pb(OH)_2$) and aqueous sodium nitrate ($NaNO_3$).

Step 1: Write the unbalanced ionic equation: $Pb^{2+}(aq) + NO_3^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Pb(OH)_2(s) + Na^+(aq) + NO_3^-(aq)$

Step 2: Balance the equation: $Pb^{2+}(aq) + 2NO_3^-(aq) + 2Na^+(aq) + 2OH^-(aq) \rightarrow Pb(OH)_2(s) + 2Na^+(aq) + 2NO_3^-(aq)$

The balanced ionic equation is: $Pb^{2+}(aq) + 2NO_3^-(aq) + 2OH^-(aq) \rightarrow Pb(OH)_2(s)$

4. Solutions and Concentration Calculations

Defining Solution Concentration

- **Concentration**: The amount of solute per unit volume of solution.
- **Molarity (M)**: The number of moles of solute per liter of solution.
- **Molality (m)**: The number of moles of solute per kilogram of solvent.

Example 8:

Calculating Molarity What is the molarity of a solution containing 12.4 g of NaOH dissolved in 250 mL of water?

Step 1: Convert volume to liters

250 mL = 0.250 L

Step 2: Calculate moles of NaOH

Moles of NaOH = 12.4 g / 40.00 g/mol = 0.310 mol

Step 3: Calculate molarity

Molarity = Moles of solute / Volume of solution in liters

Molarity = 0.310 mol / 0.250 L = 1.24 M

Example 9:

Calculating Concentration by Titration

A 25.0 mL sample of aqueous hydrochloric acid (HCl) was titrated with 0.100 M sodium hydroxide (NaOH) solution. It was found that 20.0 mL of the NaOH solution was required to reach the endpoint.

Calculate the concentration of the HCl solution.

Step 1: Write the balanced equation

$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

Step 2: Use the titration volumes to calculate moles of HCl

Moles of NaOH = 0.100 M × 0.0200 L = 0.00200 mol Moles of HCl = 0.00200 mol (since 1:1 mole ratio)

Step 3:

Calculate the concentration of HCl

Concentration of HCl = Moles of HCl / Volume of HCl sample

Concentration of HCl = 0.00200 mol / 0.0250 L = 0.0800 M

5. Gas Volumes and Stoichiometry

Molar Gas Volume

- At standard temperature and pressure (STP), 1 mole of any gas occupies 22.4 L.
- This is known as the molar gas volume.

Example 10:

Calculating Gas Volumes from Moles What volume of carbon dioxide (CO_2) is produced when 2.5 g of calcium carbonate ($CaCO_3$) reacts completely with excess hydrochloric acid (HCl)?

Step 1: Write the balanced equation

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

Step 2: Calculate moles of $CaCO_3$

Moles of $CaCO_3$ = 2.5 g / 100.09 g/mol = 0.0250 mol

Step 3: Use the balanced equation to find moles of CO_2

1 mol $CaCO_3$ → 1 mol CO_2

0.0250 mol $CaCO_3 \rightarrow 0.0250$ mol CO_2

Step 4: Calculate the volume of CO_2 at STP

Volume of CO_2 = Moles of CO_2 × Molar gas volume

Volume of CO_2 = 0.0250 mol × 22.4 L/mol = 0.560 L

Example 11:

Stoichiometry Calculations Involving Gases

In the reaction: $2Al(s) + 3HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$ If 5.40 g of Al reacts completely, calculate the volume of hydrogen gas produced at STP.

Step 1: Calculate moles of Al

Moles of Al = 5.40 g / 26.98 g/mol = 0.200 mol

Step 2: Use the balanced equation to find moles of H_2

2 mol Al → 3 mol H_2 0.200 mol Al → 0.300 mol H_2

Step 3: Calculate the volume of H_2 at STP

Volume of H_2 = Moles of H_2 × Molar gas volume

Volume of H_2 = 0.300 mol × 22.4 L/mol = 6.72 L

Practice Problems

1. A compound contains 40.0% C, 6.67% H, and 53.33% O by mass. Determine its empirical and molecular formulae if the molar mass is 180 g/mol.
2. Balance the following equations: a) $Na + Cl_2 \rightarrow NaCl$ b) $CH_4 + O_2 \rightarrow CO_2 + H_2O$ c) $Pb(NO_3)_2 + KI \rightarrow PbI_2 + KNO_3$
3. Write the balanced ionic equations for the following reactions: a) $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$ b) $Pb(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Pb(OH)_2(s) + 2NaNO_3(aq)$
4. Calculate the molarity of a solution containing 8.2 g of Na_2CO_3 dissolved in 250 mL of water.
5. In a titration, 35.0 mL of 0.120 M HCl was neutralized by 28.0 mL of NaOH solution. Calculate the concentration of the NaOH solution.
6. What volume of CO_2 is produced at STP when 10.0 g of $CaCO_3$ reacts completely with excess HCl?
7. In the reaction: $2Na + 2HCl \rightarrow 2NaCl + H_2$ Calculate the volume of H_2 gas produced at STP when 3.45 g of Na reacts completely.

Exam Technique Tips

1. Balancing Equations:

- Identify limiting reactants and ensure mass balance
- Use lowest whole number coefficients
- Check that atoms are balanced on both sides

2. Ionic Equations:
   - Focus on the dissociated ions, not the full compounds
   - Identify spectator ions and cancel them out
   - Ensure charges are balanced
3. Concentration Calculations:
   - Pay attention to units (mol/L vs mol/kg)
   - Show clear step-by-step workings
   - Check final units match the question
4. Gas Volume Problems:
   - Use the molar gas volume of 22.4 L/mol at STP
   - Relate moles of gas to volume using the balanced equation
   - Consider temperature and pressure if not at STP
5. General Tips:
   - Read the question carefully
   - Identify all given information
   - Plan your approach before solving
   - Check units and significant figures
   - Verify that your final answer is reasonable

6. Mass Spectra Questions:

   - Look for peak heights (relative abundance)
   - Check if data is in percentage form
   - Round final answer appropriately
   - Consider isotopic pattern

7. Mole Calculations:

   - Write all units
   - Check conversion factors
   - Show intermediate steps
   - Verify answer magnitude

8. Stoichiometry:

   - Always write balanced equation first
   - Identify limiting reagent if multiple reactants
   - Check units match
   - Consider percentage yield

9. Empirical Formula:

   - Assume 100g sample if percentages given
   - Show ratio simplification steps
   - Verify total percentage = 100%
   - Remember to find n for molecular formula

Common Student Mistakes

1. Forgetting to balance equations correctly
2. Mixing up empirical and molecular formulae
3. Confusing molarity and molality
4. Incorrect conversion of units (e.g., mL to L)
5. Failing to consider limiting reactants in stoichiometry
6. Using the wrong molar gas volume value

Essential Data and Equations

1. **Constants**:
   - Molar gas volume at STP = 22.4 L/mol
2. **Key Equations**:
   - Molarity (M) = moles of solute / volume of solution in liters
   - Molality (m) = moles of solute / mass of solvent in kg
   - Volume of gas = moles of gas × 22.4 L/mol (at STP)
   - Avogadro constant = 6.02×10^{23} mol^{-1}
   - Molar gas volume = 24 dm^3 (at RTP)
   - Key Equations:
   - n = m/M
   - n = V/Vm
   - n = N/NA
   - % composition = (mass of element/total mass) × 100

3. Unit Conversions:

- 1 dm^3 = 1000 cm^3
- 1 kg = 1000 g
- 1 dm^3 = 1 L