🧪 SOME BASIC CONCEPTS OF CHEMISTRY

Class 11 Chemistry - Chapter 1

CBSE Board 2025-26 Syllabus

Complete Notes | Tips & Tricks | Case Studies | Expected Questions

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★ ★ ★ COMPREHENSIVE STUDY MATERIAL FOR CLASS 11 CHEMISTRY ★ ★ ★

📌 CHAPTER OVERVIEW

🎯 What You Will Learn:

- **Nature of Matter:** Classification into solid, liquid, and gas; Elements, compounds, and mixtures
- **Physical & Chemical Properties:** Measurement and quantification
- **SI Units:** International System of Units and conversions
- **Uncertainty in Measurement:** Scientific notation, significant figures, precision vs. accuracy
- **Laws of Chemical Combination:** Conservation of mass, definite proportions, multiple proportions, Gay-Lussac's law, Avogadro's law
- **Dalton's Atomic Theory:** Fundamental postulates
- **Atomic & Molecular Masses:** Calculations and average atomic mass
- **Mole Concept:** Avogadro's number, molar mass
- **Percentage Composition:** Empirical and molecular formulas
- **Stoichiometry:** Limiting reagent, stoichiometric calculations
- **Concentration Terms:** Molarity, molality, mole fraction, mass percent

💡 Why This Chapter is Important:

This is the **foundation chapter** of Chemistry! The concepts learned here are used throughout Class 11 and Class 12, and are crucial for:

- ✅ CBSE Board Exams (3-5 marks typically)
- ✅ JEE Main & Advanced
- ✅ NEET
- ✅ Understanding all future chemistry chapters

Key Focus Areas: Mole concept, Stoichiometry, Limiting reagent, Concentration terms

🔬 DETAILED CONCEPTS

1. NATURE OF MATTER

States of Matter

Property	Solid	Liquid	Gas
Shape	Definite shape	No definite shape (takes container shape)	No definite shape
Volume	Definite volume	Definite volume	No definite volume
Particle Arrangement	Closely packed, orderly	Close but can move around	Far apart, random motion
Compressibility	Negligible	Very small	High

Classification of Matter

Matter Classification Tree

```
MATTER
   |
   |-------- PURE SUBSTANCES
   |               |
   |               |---- ELEMENTS (e.g., Na, Cu, O2, H2)
   |               |
   |               |---- COMPOUNDS (e.g., H2O, CO2, NaCl)
   |
   |-------- MIXTURES
                   |
                   |---- HOMOGENEOUS (e.g., sugar solution, air)
                   |
                   |---- HETEROGENEOUS (e.g., sand + salt, oil + water)
```

🎯 Quick Memory Trick:

"ELEM-COM-MIX"

- **ELEM**ents → Single type of atoms (Na, Fe)
- **COM**pounds → Fixed ratio combination (H_2O = always 2:1)
- **MIX**tures → Variable composition, can be separated physically

2. SI UNITS & MEASUREMENT

7 Base SI Units (REMEMBER: "MATCH IS")

Metre (length) - **A**mpere (current) - **T**emperature (Kelvin)
Candela (luminous intensity) - **H**our? NO! Second (time)
Inertia? NO! Kilogram (mass) - **S**ubstance (mole)

Physical Quantity	SI Unit	Symbol
Length	metre	m
Mass	kilogram	kg
Time	second	s
Electric Current	ampere	A
Temperature	kelvin	K
Amount of Substance	mole	mol
Luminous Intensity	candela	cd

Important Prefixes

Prefix	Symbol	Multiplier	Example
nano	n	10^{-9}	1 nm = 10^{-9} m
micro	μ	10^{-6}	1 μm = 10^{-6} m
milli	m	10^{-3}	1 mm = 10^{-3} m
centi	c	10^{-2}	1 cm = 10^{-2} m
kilo	k	10^{3}	1 km = 10^{3} m
mega	M	10^{6}	1 Mm = 10^{6} m

🔥 **Temperature Conversion Tricks:**

$$K = {}^\circ C + 273.15$$

$$^\circ F = (9/5) \times {}^\circ C + 32$$

$$^\circ C = (5/9) \times ({}^\circ F - 32)$$

Quick Trick: Remember "9/5 for F, 5/9 from F"

3. UNCERTAINTY IN MEASUREMENT

Scientific Notation

$$N \times 10^n$$

Where: $1 \leq N < 10$ and n is an integer

Examples:

- $232.508 = 2.32508 \times 10^2$
- $0.00016 = 1.6 \times 10^{-4}$
- $602{,}200{,}000{,}000{,}000{,}000{,}000{,}000 = 6.022 \times 10^{23}$

Significant Figures - Rules

🎯 Golden Rules of Significant Figures:

Rule 1: All non-zero digits are significant

Example: 285 → 3 significant figures

Rule 2: Leading zeros are NOT significant

Example: 0.0052 → 2 significant figures (only 5 and 2)

Rule 3: Zeros between non-zero digits ARE significant

Example: 2005 → 4 significant figures

Rule 4: Trailing zeros AFTER decimal are significant

Example: 0.200 → 3 significant figures

Rule 5: Trailing zeros WITHOUT decimal are NOT significant (unless specified)

Example: 100 → 1 significant figure (but 100. or 1.00×10^2 → 3 significant figures)

Rounding Off Rules

If digit to be removed is > 5: Increase preceding digit by 1

Example: 1.386 → 1.39

If digit to be removed is < 5: Keep preceding digit unchanged

Example: 4.334 → 4.33

If digit is exactly 5:

- **Preceding digit ODD → Increase by 1 (6.35 → 6.4)**
- **Preceding digit EVEN → Keep unchanged (6.25 → 6.2)**

⚠️ Common Mistakes to Avoid:

- ❌ Counting leading zeros as significant (0.003 has only 1 significant figure, not 4!)
- ❌ Not considering trailing zeros after decimal (2.00 has 3 sig figs, not 1!)
- ❌ Wrong rounding when result has more sig figs than least precise measurement
- ❌ Forgetting that exact numbers (counting) have infinite sig figs

⚖️ LAWS OF CHEMICAL COMBINATION

1. Law of Conservation of Mass (Lavoisier, 1789)

"Matter can neither be created nor destroyed"

Mass of reactants = Mass of products

Example:

$2Na + Cl_2 \rightarrow 2NaCl$

46g + 71g = 117g

Total mass remains constant!

2. Law of Definite Proportions (Proust)

"A compound always contains the same elements in the same proportion by mass"

Example: Water (H_2O)

H : O = 2 : 16 = 1 : 8 (always!)

Whether from river, rain, or laboratory → same ratio

3. Law of Multiple Proportions (Dalton, 1803)

"When two elements form more than one compound, masses of one element that combine with fixed mass of other are in simple whole number ratio"

Example: CO and CO_2

In CO: C(12g) + O(16g)

In CO_2: C(12g) + O(32g)

Oxygen masses: 16:32 = 1:2 ✓ (Simple ratio!)

🎯 **Memory Trick for Laws:**

"CDMGA"

- **C**onservation - Total mass constant
- **D**efinite - Same compound, same ratio
- **M**ultiple - Different compounds, simple ratio
- **G**ay-Lussac - Gas volumes in simple ratio
- **A**vogadro - Equal volumes, equal molecules

4. Gay-Lussac's Law of Gaseous Volumes (1808)

"Gases combine in simple volume ratios at same T and P"

Example:

$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

100 mL + 50 mL → 100 mL

Volume ratio: 2:1:2 ✓

5. Avogadro's Law (1811)

"Equal volumes of gases at same T and P contain equal number of molecules"

⚛️ DALTON'S ATOMIC THEORY

📜 John Dalton (1808)

Main Postulates:

1. Matter consists of indivisible atoms
2. All atoms of same element are identical in mass and properties
3. Atoms of different elements differ in mass
4. Compounds are formed when atoms combine in fixed ratios
5. Chemical reactions involve reorganization of atoms (not creation/destruction)

💡 Modern Modifications:

- Atoms ARE divisible (electrons, protons, neutrons)
- Isotopes exist → same element, different masses
- Nuclear reactions CAN convert atoms

⚖️ ATOMIC & MOLECULAR MASSES

Atomic Mass Unit (u)

1 u = 1/12 × mass of one C-12 atom

$1\ u = 1.66056 \times 10^{-24}$ g

🔥 **Quick Calculation Trick:**

Mass of 1 atom of any element = **Atomic mass / (6.022×10^{23}) grams**

Average Atomic Mass

Average Atomic Mass = Σ (Isotope mass × Relative abundance)

Example: Chlorine

^{35}Cl: 34.97 u (75.77%)

^{37}Cl: 36.97 u (24.23%)

Average = (34.97 × 0.7577) + (36.97 × 0.2423) = **35.45 u**

Molecular Mass

Molecular Mass = Σ (Atomic mass × Number of atoms)

Examples:

H_2O: (2 × 1) + (1 × 16) = 18 u

CO_2: (1 × 12) + (2 × 16) = 44 u

CH_4: (1 × 12) + (4 × 1) = 16 u

$C_6H_{12}O_6$: (6 × 12) + (12 × 1) + (6 × 16) = 180 u

🧮 MOLE CONCEPT - THE HEART OF CHEMISTRY!

⭐ THE MOST IMPORTANT CONCEPT! ⭐

1 mole = 6.022 × 10^{23} entities (Avogadro's Number)

🎯 "MOLE TRIANGLE" - Master This!

```
        MOLES (n)
           /\
          /  \
         /    \
        /______\
    MASS    NUMBER OF
   (m)     PARTICLES (N)
```

Key Formulas:

$n = m/M$ (mass/molar mass)

$n = N/N_a$ (particles/Avogadro's number)

$n = V/22.4$ (volume in L at STP/22.4)

The Golden Mole Concept Formulas

1. Number of Moles (n)

n = Given mass (g) / Molar mass (g/mol)

2. Number of Particles (N)

$$N = n \times N_a = n \times 6.022 \times 10^{23}$$

3. For Gases at STP

n = Volume (L) / 22.4 L

(Note: 1 mole of any gas = 22.4 L at STP)

4. Mass Calculation

Mass = n × Molar mass

🔥 SUPER IMPORTANT Examples:

Q1: How many molecules in 36g of H_2O?

Solution:

Step 1: n = 36/18 = 2 moles

Step 2: N = 2 × 6.022 × 10^{23} = **1.2044 × 10^{24} molecules**

Q2: What is mass of 3.011 × 10^{23} atoms of Cu? (Atomic mass = 63.5)

Solution:

Step 1: n = N/N_a = 3.011×10^{23} / 6.022×10^{23} = 0.5 mole

Step 2: Mass = 0.5 × 63.5 = **31.75 g**

Q3: How many moles in 11.2 L of CO_2 at STP?

Solution:

n = 11.2/22.4 = **0.5 mole**

🚀 Super Fast Mole Tricks:

Trick 1: "Dozen Analogy"

1 dozen = 12 items

1 mole = 6.022×10^{23} items

Trick 2: Quick Molecular Count

Molecules in compound = Moles × N_a

Total atoms = Molecules × Atomicity

Example: 1 mole H_2O → 6.022×10^{23} molecules → $3 \times 6.022 \times 10^{23}$ atoms!

Trick 3: Remember "6-2-2"

6.022 × **10^{23}** (Avogadro)

22.4 L (Molar volume at STP)

📊 PERCENTAGE COMPOSITION & FORMULAS

Mass Percent

Mass % of element = (Mass of element in compound / Molar mass of compound) × 100

Example: H_2O (Molar mass = 18 g/mol)

Mass % of H = (2/18) × 100 = **11.11%**

Mass % of O = (16/18) × 100 = **88.89%**

Check: 11.11 + 88.89 = 100% ✓

Empirical vs Molecular Formula

Empirical Formula	Molecular Formula
Simplest whole number ratio	Actual number of atoms
Example: CH_2O	Example: $C_6H_{12}O_6$ (glucose)
Lower formula mass	Higher (actual) molecular mass

Molecular Formula = $(\text{Empirical Formula})_n$

Where: n = Molecular mass / Empirical formula mass

🎯 5-STEP METHOD for Empirical Formula:

Step 1: Convert mass % to grams (assume 100g sample)

Step 2: Convert grams to moles (divide by atomic mass)

Step 3: Divide all moles by smallest mole value

Step 4: If not whole numbers, multiply by 2, 3, 4... to get whole numbers

Step 5: Write empirical formula

🔥 Complete Example:

Q: A compound contains 4.07% H, 24.27% C, 71.65% Cl. Molar mass = 98.96 g. Find empirical and molecular formula.

Solution:

Step 1: In 100g: H = 4.07g, C = 24.27g, Cl = 71.65g

Step 2: Moles:

- H: 4.07/1 = 4.07 mol
- C: 24.27/12 = 2.02 mol
- Cl: 71.65/35.5 = 2.02 mol

Step 3: Divide by smallest (2.02):

- H: 4.07/2.02 ≈ 2
- C: 2.02/2.02 = 1
- Cl: 2.02/2.02 = 1

Step 4 & 5: Empirical formula = **CH_2Cl**

Empirical mass = 12 + 2 + 35.5 = 49.5 g

n = 98.96/49.5 ≈ 2

Molecular formula = $C_2H_4Cl_2$

⚗️ STOICHIOMETRY - CALCULATIONS IN CHEMISTRY

🎯 **Stoichiometry = "Measurement of Elements"**

Key Principle: Balanced equation coefficients = Mole ratios!

Steps for Stoichiometric Calculations

📝 **Universal 4-Step Method:**

1. **Write balanced equation**
2. **Convert given quantity to moles**
3. **Use mole ratio from equation**
4. **Convert moles to required units**

🔥 **MASTER EXAMPLE:**

Q: How much CO_2 is produced when 16g CH_4 burns completely?

Step 1: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Step 2: Moles of CH_4 = 16/16 = 1 mole

Step 3: From equation: 1 mol CH_4 → 1 mol CO_2

So, 1 mol CH_4 gives 1 mol CO_2

Step 4: Mass of CO_2 = 1 × 44 = **44 g**

Limiting Reagent Concept

Limiting Reagent = Reactant that gets consumed first

(Determines maximum product formation)

🎯 How to Find Limiting Reagent:

1. Calculate moles of all reactants
2. Divide each by its stoichiometric coefficient
3. Smallest value → That's your limiting reagent!

🔥 CRITICAL EXAMPLE for Exams:

Q: $N_2 + 3H_2 \rightarrow 2NH_3$

If 50 kg N_2 reacts with 10 kg H_2, find limiting reagent and mass of NH_3 formed.

Solution:

Moles of N_2 = 50,000/28 = 1785.7 mol

Moles of H_2 = 10,000/2 = 5000 mol

From equation: 1 mol N_2 needs 3 mol H_2

So, 1785.7 mol N_2 needs: 1785.7 × 3 = 5357.1 mol H_2

But we have only 5000 mol H_2

∴ H_2 is the limiting reagent!

From H_2: 3 mol H_2 → 2 mol NH_3

5000 mol H_2 → (5000 × 2)/3 = 3333.3 mol NH_3

Mass = 3333.3 × 17 = **56,666 g = 56.67 kg**

⚠️ Common Stoichiometry Mistakes:

- ❌ Forgetting to balance equation first!
- ❌ Using mass directly instead of moles
- ❌ Wrong mole ratio from unbalanced equation
- ❌ Not identifying limiting reagent when both reactants given
- ❌ Calculation errors in molar mass

🧪 CONCENTRATION TERMS

1. Mass Percent (w/w%)

Mass % = (Mass of solute / Mass of solution) × 100

2. Mole Fraction (χ)

$$\chi_a = n_a / (n_a + n_b)$$

Note: $\chi_a + \chi_b = 1$ (for binary solution)

3. Molarity (M) ⭐ MOST IMPORTANT!

Molarity (M) = Moles of solute / Volume of solution (L)

OR: M = (Mass of solute × 1000) / (Molar mass × Volume in mL)

🔥 **Molarity Dilution Formula:**

$$M_1V_1 = M_2V_2$$

(Works because moles remain constant!)

4. Molality (m)

Molality (m) = Moles of solute / Mass of solvent (kg)

Key Difference: Independent of temperature!

Property	Molarity (M)	Molality (m)
Depends on	Volume of solution	Mass of solvent
Temperature effect	Changes with T	Independent of T
Unit	mol/L or M	mol/kg or m
When to use	General calculations	Colligative properties

🔥 SUPER IMPORTANT Calculation:

Q: Calculate molarity of solution with 5.85g NaCl in 500 mL solution.

Solution:

Molar mass of NaCl = 23 + 35.5 = 58.5 g/mol

Moles = 5.85/58.5 = 0.1 mol

Volume = 500 mL = 0.5 L

Molarity = 0.1/0.5 = 0.2 M

Q: How to prepare 250 mL of 0.5M NaOH from 2M stock solution?

Solution:

$M_1V_1 = M_2V_2$

$2 \times V_1 = 0.5 \times 250$

V_1 = 125/2 = **62.5 mL**

Take 62.5 mL of 2M solution and dilute to 250 mL

🎯 Quick Memory Trick:

"MoLaR-iTy" has T → Temperature dependent

"MoLaL-iTy" has L → Links to soLvent mass (temperature independent)

💡 TIPS, TRICKS & SHORTCUTS

🚀 Speed Calculation Tricks for Exams:

Trick 1: Mole-Mass Quick Convert

If you know one, remember:

- 1 mole H_2O = 18 g = 6.022×10^{23} molecules = 22.4 L (gas at STP)
- 1 mole CO_2 = 44 g = 6.022×10^{23} molecules = 22.4 L (gas at STP)
- 1 mole O_2 = 32 g = 6.022×10^{23} molecules = 22.4 L (gas at STP)

Trick 2: Percentage Composition Shortcut

For compound A_xB_Y:

% of A = (x × At.mass of A / Mol.mass of compound) × 100

Then: % of B = 100 - % of A

Trick 3: Limiting Reagent Quick Test

Moles/Coefficient method:

Calculate (moles/stoichiometric coefficient) for each reactant

Smallest value = Limiting reagent

Trick 4: Empirical Formula Pattern

Common ratios to memorize:

- If ratio is 1:1.5 → multiply by 2 (gets 2:3)
- If ratio is 1:1.33 → multiply by 3 (gets 3:4)
- If ratio is 1:1.25 → multiply by 4 (gets 4:5)

Trick 5: Molarity-Molality Relationship (Approximate)

For dilute aqueous solutions (density ≈ 1 g/mL):

M ≈ m (approximately equal!)

⚡ Must Memorize Values:

- **Avogadro's Number:** 6.022×10^{23}
- **1 amu:** 1.66×10^{-24} g
- **STP:** 273 K, 1 atm
- **Molar volume at STP:** 22.4 L
- **Atomic masses:** H=1, C=12, N=14, O=16, Na=23, Cl=35.5, Ca=40

📋 CASE-BASED QUESTIONS (Latest CBSE Pattern)

Case Study 1: Pharmaceutical Analysis

Passage: A pharmaceutical company is analyzing a sample of aspirin (acetylsalicylic acid, $C_9H_8O_4$). The quality control team needs to verify the purity of a 500g batch. The molecular mass of aspirin is 180 g/mol. The sample is dissolved in an appropriate solvent to prepare solutions for testing.

Q1. How many moles of aspirin are present in the 500g sample?

(a) 2.5 mol

(b) 2.78 mol ✓

(c) 3.0 mol

(d) 2.0 mol

Answer: (b) Moles = 500/180 = 2.78 mol

Q2. How many molecules of aspirin are in this sample?

(a) 1.67×10^{24} ✓

(b) 3.01×10^{23}

(c) 6.02×10^{23}

(d) 1.20×10^{24}

Answer: (a) $N = 2.78 \times 6.022\times10^{23} = 1.67\times10^{24}$

Q3. What is the mass percentage of carbon in aspirin? (C=12, H=1, O=16)

(a) 50%

(b) 55%

(c) 60% ✓

(d) 65%

Answer: (c) Mass of C in $C_9H_8O_4$ = 9×12 = 108g

Mass % = (108/180) × 100 = 60%

Q4. If 100 mL of 0.5M aspirin solution is diluted to 500 mL, what is the new molarity?

(a) 0.05 M

(b) 0.1 M ✓

(c) 0.2 M

(d) 0.25 M

Answer: (b) $M_1V_1 = M_2V_2 \rightarrow 0.5\times100 = M_2\times500 \rightarrow M_2 = 0.1M$

Case Study 2: Environmental Chemistry

Passage: Carbon dioxide (CO_2) levels in the atmosphere have been rising due to industrial emissions. A monitoring station measures CO_2 concentration in parts per million (ppm). In a controlled experiment, 88g of CO_2 gas is collected at STP to study its properties. (Molar mass of CO_2 = 44 g/mol)

Q1. How many moles of CO_2 are in the sample?

(a) 1 mol

(b) 2 mol ✓

(c) 3 mol

(d) 4 mol

Answer: (b) n = 88/44 = 2 mol

Q2. What volume does this CO_2 occupy at STP?

(a) 22.4 L

(b) 33.6 L

(c) 44.8 L ✓

(d) 11.2 L

Answer: (c) V = n × 22.4 = 2 × 22.4 = 44.8 L

Q3. How many oxygen atoms are present in this sample?

(a) 6.022×10^{23}

(b) 1.204×10^{24}

(c) 2.408×10^{24} ✓

(d) 3.011×10^{23}

Answer: (c) Molecules = 2 × 6.022×10^{23} = 1.204×10^{24}

Oxygen atoms = 2 × 1.204×10^{24} = 2.408×10^{24}

Q4. What is the mass percentage of oxygen in CO_2?

(a) 36.36%

(b) 50%

(c) 63.64%

(d) 72.73% ✓

Answer: (d) Mass of O in CO_2 = 2×16 = 32g

Mass % = (32/44) × 100 = 72.73%

Case Study 3: Industrial Chemistry

Passage: Ammonia (NH_3) is produced industrially by the Haber process: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$. A chemical plant reacts 28 kg of N_2 with 9 kg of H_2 to produce ammonia. (N=14, H=1)

Q1. What is the limiting reagent?

(a) N_2

(b) H_2 ✓

(c) Both

(d) Neither

Solution:

Moles of N_2 = 28,000/28 = 1000 mol

Moles of H_2 = 9,000/2 = 4500 mol

For 1000 mol N_2, need 3000 mol H_2

But we have 4500 mol H_2, so N_2 should be limiting?

Wait! Check properly: 1000 mol N_2 needs 3000 mol H_2 (we have 4500) ✓

Actually, check from H_2: 4500 mol H_2 needs 4500/3 = 1500 mol N_2

But we have only 1000 mol N_2!

Answer: (a) N_2 is limiting (Corrected!)

Q2. Maximum mass of NH_3 produced:

(a) 17 kg

(b) 34 kg ✓

(c) 51 kg

(d) 25.5 kg

Answer: (b) From N_2: 1 mol N_2 → 2 mol NH_3

1000 mol N_2 → 2000 mol NH_3

Mass = 2000 × 17 = 34,000 g = 34 kg

Q3. Mass of unreacted reagent:

(a) 3 kg ✓

(b) 4.5 kg

(c) 6 kg

(d) 0 kg

Answer: (a) H_2 used = 1000×3 = 3000 mol

H_2 remaining = 4500-3000 = 1500 mol

Mass = 1500 × 2 = 3000 g = 3 kg

🎯 MOST EXPECTED QUESTIONS FOR EXAMS

Section A: MCQs (1 mark each)

Q1. One mole of any substance contains:

(a) 6.023×10^{22}

(b) 6.022×10^{23} ✓

(c) 6.022×10^{22}

(d) 3.011×10^{23}

Q2. The SI unit of amount of substance is:

(a) kilogram

(b) candela

(c) mole ✓

(d) metre

Q3. Which has maximum number of atoms?

(a) 1g of Li ✓

(b) 1g of Na

(c) 1g of Fe

(d) 1g of Ag

Explanation: Lighter element → more atoms per gram!

Q4. Number of significant figures in 0.00250:

(a) 2

(b) 3 ✓

(c) 4

(d) 5

Q5. Empirical formula of $C_6H_{12}O_6$:

(a) CHO

(b) CH_2O ✓

(c) $C_2H_4O_2$

(d) $C_3H_6O_3$

Section B: Short Answer (2-3 marks)

Q6. Define and distinguish between molarity and molality.

Answer:

Molarity (M): Number of moles of solute per litre of solution. Temperature dependent.

Formula: M = n/V(L)

Molality (m): Number of moles of solute per kg of solvent. Temperature independent.

Formula: m = n/mass of solvent(kg)

Q7. Calculate the number of atoms in 52g of He.

Answer:

Atomic mass of He = 4g/mol

Moles = 52/4 = 13 mol

Number of atoms = 13 × 6.022×10^{23} = **7.83×10^{24} atoms**

Q8. State the law of multiple proportions with example.

Answer:

When two elements combine to form two or more compounds, the masses of one element combining with fixed mass of other are in simple whole number ratio.

Example: CO and CO_2

In CO: 12g C combines with 16g O

In CO_2: 12g C combines with 32g O

Ratio of oxygen masses = 16:32 = 1:2 (simple ratio)

Section C: Long Answer (5 marks)

Q9. (a) What is limiting reagent? How do you identify it? (3 marks)

(b) 50g $CaCO_3$ is treated with 50g HCl. Find limiting reagent and mass of CO_2 produced. (2 marks)

Answer:

(a) The reactant which is completely consumed first and limits the amount of product formed is called limiting reagent.

Identification Method:

1. Calculate moles of all reactants
2. Divide each by stoichiometric coefficient
3. Smallest value indicates limiting reagent

(b) $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$

Moles of $CaCO_3$ = 50/100 = 0.5 mol

Moles of HCl = 50/36.5 = 1.37 mol

From equation: 1 mol $CaCO_3$ needs 2 mol HCl

So, 0.5 mol $CaCO_3$ needs 1 mol HCl

We have 1.37 mol HCl (excess!)

$CaCO_3$ is limiting reagent

From equation: 1 mol $CaCO_3$ → 1 mol CO_2

0.5 mol $CaCO_3$ → 0.5 mol CO_2

Mass of CO_2 = 0.5 × 44 = 22g

Q10. (a) Define significant figures. (1 mark)

(b) State rules for determining significant figures. (3 marks)

(c) Round off 2.865 to 3 significant figures. (1 mark)

Answer:

(a) Significant figures are meaningful digits in a number that indicate the precision of measurement.

(b) Rules:

1. All non-zero digits are significant
2. Zeros between non-zero digits are significant
3. Leading zeros are NOT significant
4. Trailing zeros after decimal are significant
5. Trailing zeros without decimal are NOT significant

(c) 2.865 → 2.87 (since 5 is followed by odd number 6, round up) **or 2.86** (if we consider only the 5)

Correct answer using standard rule: **2.86** (5 rounds to nearest even)

🏆 EXAM PREPARATION STRATEGY

📊 Chapter Weightage: 3-5 marks

Typical question pattern:

- 1-2 MCQs (1 mark each)
- 1 Short answer (2 marks)
- 1 Numerical problem (3 marks)
- Sometimes: 1 Case study (4-5 marks)

✅ Must Practice Topics (High Probability):

1. **Mole Concept:**
   - Interconversion: mass ↔ moles ↔ molecules ↔ volume
   - At least 15 numerical problems
2. **Stoichiometry:**
   - Product/reactant mass calculations
   - Limiting reagent problems (Very Important!)
   - At least 10 problems
3. **Significant Figures:**
   - Identification and rounding
   - At least 20 practice questions
4. **Empirical & Molecular Formula:**
   - From percentage composition
   - At least 10 problems
5. **Concentration Terms:**
   - Molarity calculations and dilution
   - Molality problems
   - At least 10 problems
6. **Laws of Chemical Combination:**
   - Theory questions (2-3 marks)
   - Numerical based on laws

⚠️ Last Minute Checklist (1 Day Before Exam):

- ✅ Revise all formulas (write them down once!)
- ✅ Memorize Avogadro's number, molar volume, common atomic masses
- ✅ Practice 5 mole concept problems
- ✅ Practice 3 limiting reagent problems
- ✅ Review significant figures rules
- ✅ Practice molarity calculations
- ✅ Revise all laws with one example each
- ✅ Check SI units table

🎯 Exam Writing Tips:

1. **Always write units** in numerical answers
2. **Show all steps** in calculations (partial marks!)
3. **Use proper significant figures** in final answer
4. **Balance equations first** in stoichiometry problems
5. **Underline/highlight** final answers
6. **Write formulas clearly** before substituting values
7. **Check dimensional consistency** (units should match!)

📝 Formula Sheet - Write This on Last Page of Answer Sheet!

Concept	Formula
Moles	$n = m/M = N/N_a = V/22.4$ (STP)
Molecular Mass	Σ(Atomic mass × No. of atoms)
Mass %	(Mass of element/Molar mass) × 100
Molarity	$M = n/V(L)$; $M_1V_1 = M_2V_2$
Molality	$m = n/\text{mass of solvent(kg)}$
Mole Fraction	$\chi_a = n_a/(n_a + n\_b)$
Avogadro's Number	$N_a = 6.022 \times 10^{23}$
STP Molar Volume	22.4 L

📚 QUICK REVISION NOTES

🔑 Key Definitions (Must Know Word-to-Word)

1. Mole: Amount of substance containing 6.022×10^{23} elementary entities.

2. Avogadro's Number: 6.022×10^{23} (number of entities in 1 mole)

3. Molar Mass: Mass of 1 mole of substance (numerically equal to molecular mass in grams)

4. Empirical Formula: Simplest whole number ratio of atoms in compound

5. Molecular Formula: Actual number of atoms of each element in molecule

6. Limiting Reagent: Reactant completely consumed first, limiting product formation

7. Stoichiometry: Quantitative study of reactants and products in chemical reactions

8. Molarity: Number of moles of solute per litre of solution

9. Molality: Number of moles of solute per kg of solvent

10. Significant Figures: Meaningful digits in measurement indicating precision

⚡ One-Line Facts (Speed Revision)

- 1 u = 1.66×10^{-24} g
- 1 mole of any gas at STP = 22.4 L
- K = °C + 273.15
- °F = (9/5)°C + 32
- Density = Mass/Volume (SI unit: kg/m^3)
- Mass ≠ Weight (Weight = mg)
- 1 L = 1000 mL = 1000 cm^3 = 1 dm^3
- Precision ≠ Accuracy
- Scientific notation: $N \times 10^n$ ($1 \leq N < 10$)
- All exact numbers have infinite sig figs

🎯 Top 10 Numerical Types (Practice These!)

1. Mole ↔ Mass conversion
2. Mole ↔ Number of particles
3. Mole ↔ Volume (for gases)
4. Percentage composition calculation
5. Empirical formula from % composition
6. Molecular formula from empirical formula
7. Stoichiometric calculations (mass of product)
8. Limiting reagent identification
9. Molarity calculations
10. Dilution problems ($M_1V_1 = M_2V_2$)

📚 Study Material Information

This comprehensive study material on **Some Basic Concepts of Chemistry** has been meticulously prepared following the latest CBSE curriculum and examination pattern for Class 11 Chemistry (2025-26 session). The content includes detailed explanations of all topics, properties of matter, SI units, uncertainty in measurement, laws of chemical combination, Dalton's atomic theory, atomic and molecular masses, mole concept (the foundation of chemistry!), percentage composition, stoichiometry, limiting reagent, concentration terms, complete NCERT-aligned content, tips and tricks for fast calculations, case-based questions as per new CBSE pattern, and most expected questions for board exams.

Key Features of This Material:

- ✅ Complete Chapter 1 coverage with crystal-clear concepts
- ✅ Nature of matter: states, classification (elements, compounds, mixtures)
- ✅ SI units and measurement systems
- ✅ Scientific notation and significant figures with rules
- ✅ All five laws of chemical combination with examples
- ✅ Dalton's atomic theory and modern modifications
- ✅ Atomic mass, molecular mass, and average atomic mass
- ✅ **MOLE CONCEPT** - explained with multiple examples and tricks
- ✅ Percentage composition and formula determination
- ✅ Stoichiometry and limiting reagent (exam favorites!)
- ✅ All concentration terms: molarity, molality, mole fraction, mass %
- ✅ Memory tricks and shortcuts for faster calculations
- ✅ 3 complete case studies as per latest CBSE pattern
- ✅ Most expected questions with detailed solutions
- ✅ Common mistakes to avoid
- ✅ Exam preparation strategy and last-minute checklist
- ✅ Quick revision notes and formula sheet

🎯 Why This Chapter is THE Foundation:

Chapter 1 is not just another chapter - it's the **foundation of all chemistry!** The mole concept, stoichiometry, and concentration terms are used in:

- ✓ Solutions (Class 12)
- ✓ Chemical Equilibrium
- ✓ Thermodynamics
- ✓ Electrochemistry
- ✓ Chemical Kinetics
- ✓ JEE & NEET (direct 3-4 questions every year!)

Master this chapter = 40% of chemistry becomes easy!

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