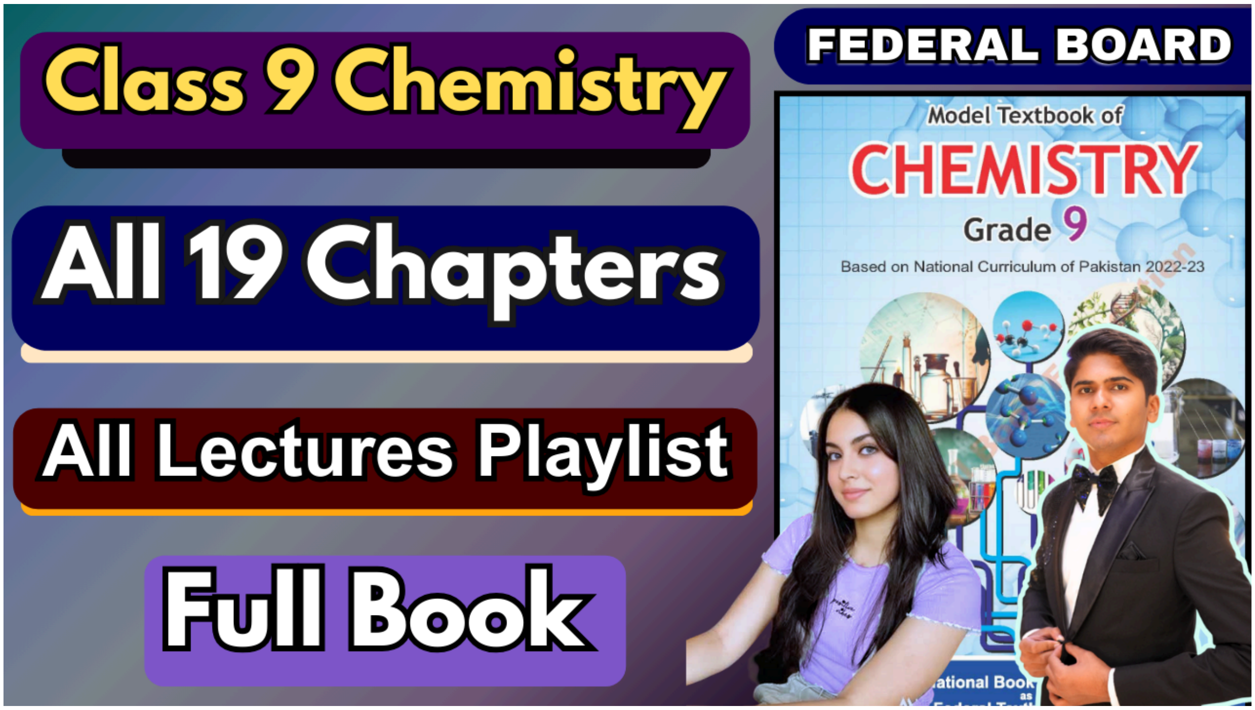
Chapter 6: Stoichiometry

All Lectures Uploaded on YouTube:

https://tinyurl.com/fkm9-chemistry



Introduction: What is Matter Made Of?

- Scientists have always wondered what makes up everything around us (wood, rocks, living things).
- In the 19th century, John Dalton proposed an **atomic theory**, which says that everything is made of tiny particles called **atoms**.
- This chapter will teach you the basic language and concepts to understand matter, which will be very helpful for future studies.

6.1. Empirical Formula and Molecular Formula

A chemical formula tells you which elements are in a compound and the number of their atoms. There are two main types:

6.1.1. Empirical Formula

What it is: The *simplest whole number ratio* of atoms in a compound.

Think of it like: A recipe reduced to its simplest form (e.g., a 2:4 ratio simplifies to 1:2). **Examples:**

- o **Hydrogen Peroxide:** The actual molecule is H_2O_2, but the simplest ratio of H to O is 1:1. So, its empirical formula is **HO**.

- o **Glucose:** The actual molecule is $C_6H_{12}O_6$. The simplest ratio of C:H:O is 1:2:1 (divide all by 6). So, its empirical formula is **CH_2O**.

EMPIRICAL FORMULA	MOLECULAR FORMULA
NO_2	N_2O_4
The ratio of nitrogen to oxygen atoms is 1:2.	There are 2 nitrogen atoms and 4 oxygen atoms.

6.1.2. Molecular Formula

What it is: Shows the *actual* number of each type of atom in a single molecule of the compound. **Think of it like:** The exact, detailed recipe.

Examples:

o **Hydrogen Peroxide:** The molecular formula is **H_2O_2** (2 hydrogen atoms and 2 oxygen atoms).

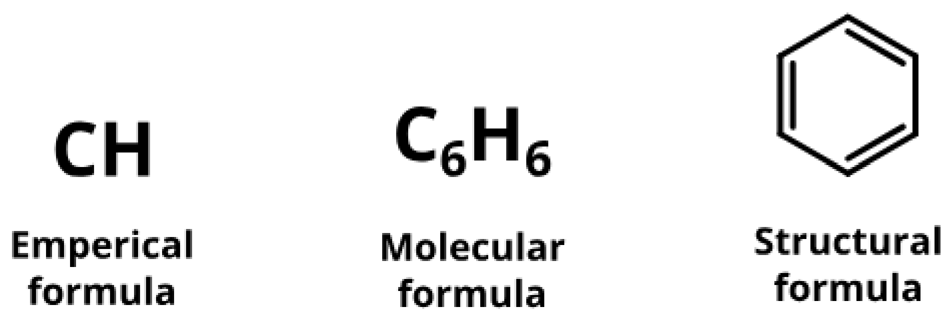
o **Glucose:** The molecular formula is **$C_6H_{12}O_6$** (6 carbon, 12 hydrogen, and 6 oxygen atoms).

o **Benzene:** The molecular formula is **C_6H_6**. Since the ratio of C to H is 1:1, its empirical formula is **CH**.

Key Relationship:

· The molecular formula is often a **multiple** of the empirical formula.

· For example, for glucose: (CH_2O) x 6 = $C_6H_{12}O_6$.



When are they the same?

For many simple compounds, the empirical formula *is* the molecular formula because the ratio cannot be simplified further.

· **Water (H_2O):** The H:O ratio is 2:1, which is already the simplest.

· **Carbon Dioxide (CO_2):** The C:O ratio is 1:2, which is already the simplest. · Others: Ammonia (NH_3), Methane (CH_4), Sulphur dioxide (SO_2).

6.2. Molecular Mass and Formula Mass

What it is: The mass of a molecule. It's the sum of the atomic masses of all the atoms in its **molecular formula**.

How to calculate it:

1. Look at the molecular formula.

2. Multiply the atomic mass of each element by the number of its atoms in the molecule. 3. Add them all together.

· **The unit is the atomic mass unit (amu)**.

Example: Calculating the Molecular Mass of Water (H_2O)

1. The formula is H_2O. So, it has **2 Hydrogen atoms** and **1 Oxygen atom**.

2. Atomic mass of Hydrogen (H) ≈ 1.008 amu

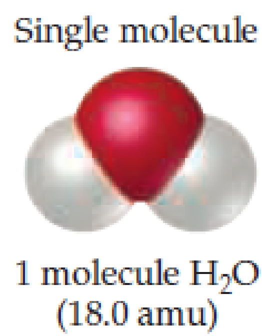
   Atomic mass of Oxygen (O) = 16.00 amu

3. Calculation:

   = (2 × 1.008 amu) + (1 × 16.00 amu)

   = 2.016 amu + 16.00 amu

   = **18.016 amu**



So, one molecule of water has a mass of about 18.016 amu.

Formula Mass (For Ionic Compounds)

Molecular Mass is the sum of atomic masses in a *molecule*. It's used for **molecular compounds** (like water H_2O or glucose $C_6H_{12}O_6$).

Formula Mass is used for **ionic compounds** (like table salt NaCl).

Why the difference? Ionic compounds (e.g., NaCl, MgO) don't exist as separate molecules. Instead, they form a giant 3D crystal lattice of positively charged ions (cations) and negatively charged ions (anions).

Formula Unit: We represent an ionic compound by its **formula unit**, which shows the simplest ratio of ions in the compound. For NaCl, the ratio of Na^+ to Cl^- ions is 1:1.

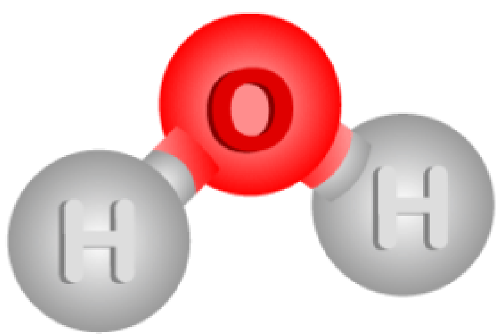
*The **Formula Mass** is the sum of the atomic masses of all the atoms in a compound's **formula unit**.*

Calculation: It's calculated the same way as molecular mass!

- *Example:* Formula mass of NaCl = Atomic mass of Na (23 amu) + Atomic mass of Cl (35.5 amu) = **58.5 amu**.

6.3. Naming and Writing Formulas for Binary Ionic Compounds

*A **binary ionic compound** is made from a metal (which forms a positive cation) and a non-metal (which forms a negative anion).*



a) How to NAME them:

1. **First,** write the name of the **metal cation** (exactly as the element's name).
2. **Second,** write the name of the **non-metal anion**, but change the ending to **"-ide"**.

- *Examples:*

NaCl: Sodium Chlor**ine** → Sodium Chlor**ide**

MgO: Magnesium Ox**ygen** → Magnesium Ox**ide**

AlN: Aluminium Nitro**gen** → Aluminium Nitr**ide**

b) How to WRITE their formulas (The Criss-Cross Method):

Step 1: Write the symbols for the cation and anion **with their charges**.

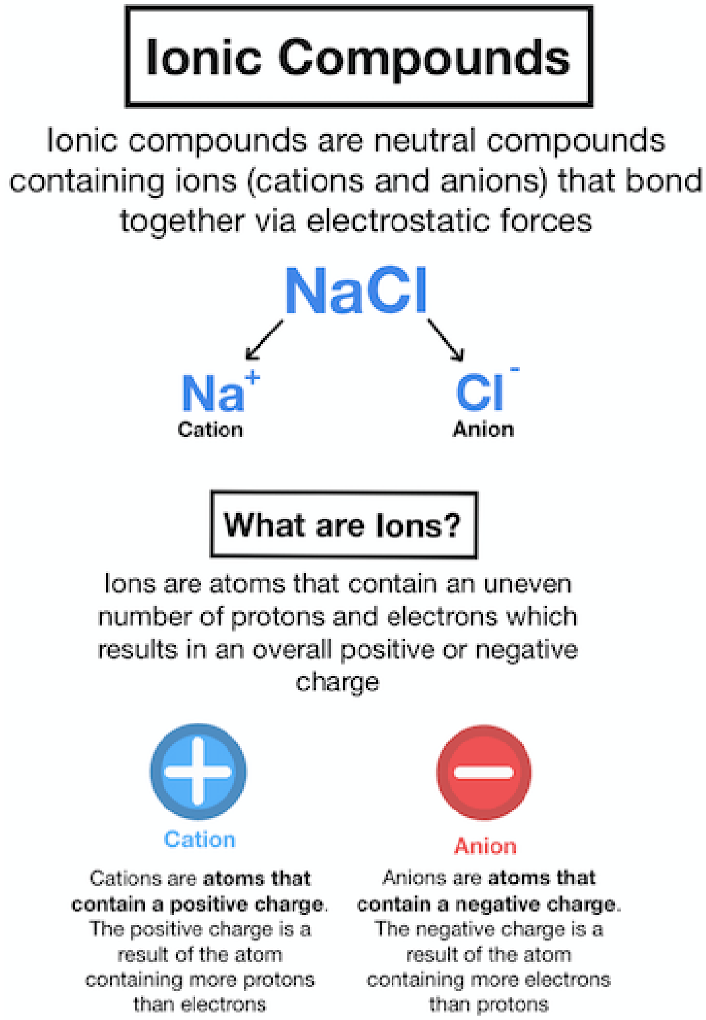
o *Example for Aluminium Oxide:* Al^{3+} and O^{2-}

Step 2: Criss-Cross the charges to become the subscripts for the other ion. Use the smallest whole numbers.

o The '3' from Al^{3+} becomes the subscript for O.

o The '2' from O^{2-} becomes the subscript for Al.

o This gives: Al_2O_3

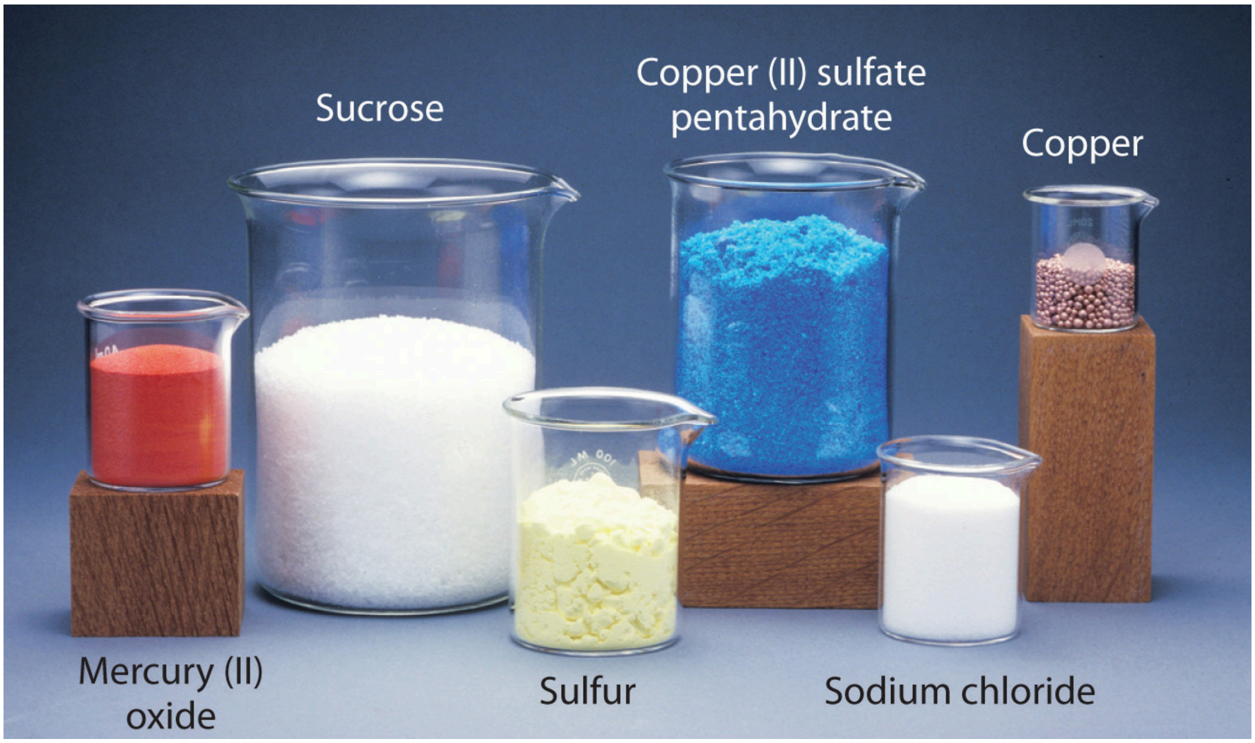


Step 3: Check for Neutrality: The total positive charge must equal the total negative charge. o Al_2O_3: (2 Al × +3) = +6 and (3 O × -2) = -6. The compound is neutral ✅

Important: If the subscripts have a common factor, simplify them. For example, Mg_2O_2 simplifies to MgO.

6.4. The Mole & Avogadro's Number

The Problem: Atoms and molecules are extremely small and tiny. We can't count them individually. We need a practical "counting unit" for the particle world.



The Solution: The **Mole (mol)**. Just like we use...

- o A **dozen** to count 12 things (eggs, oranges),
- o A **ream** to count 500 things (papers),
- o ...Chemists use a **mole** to count a very, very large number of particles.

Avogadro's Number (N_a): One mole of a substance contains **6.022×10^{23}** representative particles. This is a HUGE number: 602,200,000,000,000,000,000,000.

Examples:

▪ 1 mole of Carbon atoms = 6.022×10^{23} C atoms

▪ 1 mole of Sulphur atoms = 6.022×10^{23} S atoms

▪ 1 mole of Water molecules = 6.022×10^{23} H_2O molecules

▪ 1 mole of NaCl formula units = 6.022×10^{23} NaCl formula units

6.5. Molar Mass & Chemical Calculations

1. Molar Mass: The Grand Summary

The **mass of one mole** of any substance is called its **Molar Mass**. Its unit is grams per mole (g/mol). Think of it as the "weight" of a chemical "dozen" (where a dozen is 6.022×10^{23} particles).

This single concept has three specific names depending on what you're measuring: **a) Gram Atomic Mass (For Elements)**

What it is: The atomic mass from the periodic table, expressed in **grams**. **It contains: 1 mole of atoms** (6.022×10^{23} atoms) of an element.

Examples:

o Carbon (C): Atomic mass = 12 amu → **Gram Atomic Mass = 12 g/mol** o Sodium (Na): Atomic mass = 23 amu → **Gram Atomic Mass = 23 g/mol** o This means **12 grams of Carbon** contains exactly 6.022×10236.022×1023 carbon atoms.

b) Gram Molecular Mass (For Molecules)

The molecular mass of a compound, expressed in **grams**.

It contains: 1 mole of molecules (6.022×10^{23}molecules).

Examples:

o Water (H_2O): Molecular mass = 18.016 amu → **Gram Molecular Mass = 18.016**

g/mol o Glucose ($C_6H_{12}O_6$): Molecular mass = 180.096 amu → **Gram Molecular Mass = 180.096 g/mol**

o This means **18.016 grams of Water** contains exactly 6.022×10236.022×1023 H_2O molecules.

c) Gram Formula Mass (For Ionic Compounds)

What it is: The formula mass of an ionic compound, expressed in **grams**. **It contains: 1 mole of formula units** (6.022×10236.022×1023 formula units).
Examples:

o Sodium Chloride (NaCl): Formula mass = 58.5 amu → **Gram Formula Mass = 58.5 g/mol**

o Potassium Chloride (KCl): Formula mass = 74.5 amu → **Gram Formula Mass = 74.5 g/mol**

o This means **58.5 grams of NaCl** contains exactly 6.022×10236.022×1023 NaCl formula units.

The Unified Concept: Molar Mass

· All three terms (**Gram Atomic Mass, Gram Molecular Mass, and Gram Formula Mass**) are collectively known as **Molar Mass**.

· **Simple Definition of a Mole:** A mole is the **atomic mass, molecular mass, or formula mass expressed in grams**.

Introduction to Chemical Calculations

This is where everything comes together. The concepts of the mole, Avogadro's number, and molar mass form a powerful toolkit for solving chemical problems.

You can use these relationships as conversion factors:

Connecting Mass to Moles:

o **Moles = Mass (g) ÷ Molar Mass (g/mol)**

Connecting Moles to Number of Particles:

o **Number of Particles = Moles × (6.022×10236.022×1023 particles/mol)**

Connecting Mass directly to Number of Particles:

o **Number of Particles = [Mass (g) ÷ Molar Mass (g/mol)] × (6.022×10^{23})**

Mole-Mass Calculations & Chemical Equations

Part 1: Mole-Mass Calculations

This is the practical application of the mole concept. We use molar mass as a conversion factor. **The Core Formula:**

· **Mass (g) = Moles (mol) × Molar Mass (g/mol)**

· This formula lets you find the mass if you know the number of moles, or find the number of moles if you know the mass.

Type 1: Calculating Molar Mass

Rule:

· For a **metal** (like Na, Cu, K), the molar mass is simply its **atomic mass in grams** (Gram Atomic Mass).

· For a **non-metal that exists as a molecule** (like N_2, O_2, I_2), the molar mass is its **molecular mass in grams** (Gram Molecular Mass).

· For a **compound**, the molar mass is its **molecular or formula mass in grams**.

Example 6.3: Calculating Molar Mass

a) **Sodium (Na):** It's a metal.

Molar Mass = **23 g/mol**

b) **Nitrogen (N):** It exists as diatomic molecules (N_2).

Molecular mass of N_2 = 14 × 2 = **28 amu**

Molar Mass = **28 g/mol**

c) **Sucrose ($C_{12}H_{22}O_{11}$):** It's a molecular compound.

Molecular mass = (12×12) + (1×22) + (16×11) = 144 + 22 + 176 = **342 amu** Molar Mass = **342 g/mol**

Type 2: Calculating the Mass from Moles This uses the core formula directly: **Mass = Moles × Molar Mass Example 6.4: Mass of Ozone (O_3)**

Given: Moles of O_3 = 9.05 mol

Step 1: Find Molar Mass of O_3

- Molar Mass = 16 × 3 = **48 g/mol**

Step 2: Apply Formula

- Mass = Moles × Molar Mass
- Mass = 9.05 mol × 48 g/mol
- Mass = **434.4 g**

Answer: The mass of 9.05 moles of ozone is **434.4 grams**. **Example 6.5: Mass of Carbon Dioxide (CO_2)**

Given: Moles of CO_2 = 0.25 mol

Step 1: Find Molar Mass of CO_2

o Molar Mass = 12 + (16 × 2) = **44 g/mol**

Step 2: Apply Formula

o Mass = 0.25 mol × 44 g/mol

o Mass = **11 g**

· **Answer:** The mass of 0.25 moles of CO_2 is **11 grams**.

6.6. Chemical Equations and Balancing

A chemical equation is like a recipe for a chemical reaction. It shows what you start with and what you end up with.

Parts of a Chemical Equation:

· **Reactants:** The starting substances. They are written on the **left** side.

· **Products:** The new substances formed. They are written on the **right** side. · **Arrow (→):** Means "yields" or "produces" and shows the direction of the reaction. ·

Steps to Write a Chemical Equation:

Step 1: Write the Word Equation

· Identify the reactants and products.

· *Example (Burning Coal):* Coal + Oxygen → Carbon dioxide

Step 2: Write the Skeleton Equation with States

· Replace the names with correct chemical formulas.

- Add physical states in parentheses: (s)=solid, (l)=liquid, (g)=gas, (aq)=aqueous (dissolved in water).
- *Example (Burning Coal):*

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

- **Important:** This initial equation is often *unbalanced* (the number of atoms on the left doesn't equal the number on the right).

Step 3: Balance the Equation!

Law of Conservation of Mass: Matter cannot be created or destroyed. The number of each type of atom must be the same on both sides.

You balance by placing numbers (**coefficients**) in front of the formulas. **Never change the subscripts inside the formulas!**

Example (Burning Coal):

Skeleton: $C(s) + O_2(g) \rightarrow CO_2(g)$

Check Atoms: Left: 1 C, 2 O Right: 1 C, 2 O

It's already balanced!

Balancing Equations, Ionic Equations, and Formulas

Balancing Chemical Equations

The Golden Rule: Law of Conservation of Mass

- Atoms cannot be created or destroyed in a chemical reaction.

· Therefore, you must have the **same number of each type of atom** on the **reactant side** (left) and the **product side** (right).

A **balanced chemical equation** is one where this rule is followed.

How to Balance an Equation: A Step-by-Step Guide

· **Coefficients:** Large numbers you place *in front* of a chemical formula. This multiplies every atom in that formula. (e.g., $2H_2O$ means 2 molecules of water, totaling 4 H atoms and 2 O atoms).

· **NEVER change the subscripts** (the small numbers within the formula). This would change the identity of the compound itself!

Ionic Equations

Steps to Write an Ionic Equation:

Step 1: Write the Balanced Molecular Equation

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

Step 2: Write the Complete Ionic Equation

Split all soluble, ionic compounds (those with (aq)) into their individual ions. **Note:** Keep non-dissociating substances like liquids (l), gases (g), and solids (s) together. **Complete Ionic Equation:**

o $\mathbf{H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)}$

Step 3: Identify and Remove Spectator Ions

Spectator Ions: Ions that appear unchanged on both sides of the equation. They are present but do not participate in the actual chemical change.

In this equation: $\mathbf{Na^+(aq)}$ and $\mathbf{Cl^-(aq)}$ are spectator ions.

Step 4: Write the Net Ionic Equation

This shows only the particles that actually react.

Net Ionic Equation: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

This reveals the true essence of the reaction: an acid and a base combining to form water.

6.7. Molecular Formula vs. Structural Formula

Molecular Formula: Shows the **type and number** of atoms in a single molecule of a compound.

o *Example:* C_4H_{10}

Structural Formula: Shows **how the atoms are connected and bonded** to each other. o *Example:* CH_3-CH_2-CH_2-CH_3 (for n-Butane)

How to Write a Molecular Formula from a Structural Formula Example: For CH_3-CH_2-CH_2-OH

Identify the Elements: Carbon (C), Hydrogen (H), Oxygen (O).

Count the Atoms:

Carbon (C): There are 3 C atoms (one in each CH_3 and CH_2 group).

Hydrogen (H): Let's count carefully:

▪ CH_3- has 3 H

▪ -CH_2- has 2 H

▪ -CH_2- has 2 H

▪ The -OH has 1 H

Total H = 3 + 2 + 2 + 1 = 8

Oxygen (O): There is 1 O atom (in the -OH group).

Write the Formula: Write the symbols and the counts as subscripts.

- **Molecular Formula: C_3H_8O**

1. **Balance Equations:** Use coefficients to make atom counts equal on both sides. Never change subscripts.
2. **Write Ionic Equations:** Split (aq) compounds into ions, cancel spectators, and write the net reaction.
3. **Find Molecular Formula from Structure:** Identify all atoms and count them carefully to write the correct formula.

GET 95+%

IN FEDERAL BOARD EXAMS

FEDERAL KA MANJAN

BATCH 1.0 | Grade 9 & 10 FBISE

SUBJECTS OFFERED:

- Biology / Computer Science
- Chemistry
- Physics
- Math
- English

PROGRAM INCLUDES:

- ✓ Chapter Tests
- ✓ Monthly Tests
- ✓ Assignments
- ✓ MCQs Sheets
- ✓ Notes and Short Tricks
- ✓ Target / Guess Papers
- ✓ 24/7 Teacher Support
- ✓ Doubt Classes
- ✓ WhatsApp Group
- ✓ Mock Tests
- ✓ Live Grand Tests
- ✓ Most Important Exam Writing Method Sessions

AMAZING OFFER!

1 SUBJECT For Full One YEAR (12 Months)

~~Rs 12,000~~ **NOW ONLY Rs 3,000!**

(Per Subject for the Entire Year)

READY TO ACE YOUR EXAMS?

REGISTER ON WHATSAPP

0336-8079808