

# Chapter 6:

## Stoichiometry

### 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.

### 1. Chemical Formulas: The Blueprint of a Compound

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

#### a) 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:

- **Hydrogen Peroxide:** The actual molecule is  $\text{H}_2\text{O}_2$ , but the simplest ratio of H to O is 1:1. So, its empirical formula is **HO**.

- **Glucose:** The actual molecule is  $\text{C}_6\text{H}_{12}\text{O}_6$ . The simplest ratio of C:H:O is 1:2:1 (divide all by 6). So, its empirical formula is  **$\text{CH}_2\text{O}$** .

## b) 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:

- **Hydrogen Peroxide:** The molecular formula is  **$\text{H}_2\text{O}_2$**  (2 hydrogen atoms and 2 oxygen atoms).
- **Glucose:** The molecular formula is  **$\text{C}_6\text{H}_{12}\text{O}_6$**  (6 carbon, 12 hydrogen, and 6 oxygen atoms).
- **Benzene:** The molecular formula is  **$\text{C}_6\text{H}_6$** . Since the ratio of C to H is 1:1, its empirical formula is  **$\text{CH}$** .

### Key Relationship:

- The molecular formula is often a **multiple** of the empirical formula.
- For example, for glucose:  $(\text{CH}_2\text{O}) \times 6 = \text{C}_6\text{H}_{12}\text{O}_6$ .

### When are they the same?

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

- **Water ( $\text{H}_2\text{O}$ ):** The H:O ratio is 2:1, which is already the simplest.
- **Carbon Dioxide ( $\text{CO}_2$ ):** The C:O ratio is 1:2, which is already the simplest.
- Others: Ammonia ( $\text{NH}_3$ ), Methane ( $\text{CH}_4$ ), Sulphur dioxide ( $\text{SO}_2$ ).

## 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 atomic mass unit (amu).**

### Example: Calculating the Molecular Mass of Water (H<sub>2</sub>O)

1. The formula is H<sub>2</sub>O. So, it has **2 Hydrogen atoms** and **1 Oxygen atom**.
2. Atomic mass of Hydrogen (H)  $\approx$  1.008 amu  
Atomic mass of Oxygen (O) = 16.00 amu
3. Calculation:  
$$= (2 \times 1.008 \text{ amu}) + (1 \times 16.00 \text{ amu})$$
$$= 2.016 \text{ amu} + 16.00 \text{ amu}$$
$$= \mathbf{18.016 \text{ amu}}$$

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

### Quick Reference Table (From Your Book)

Common Substances	Formula (Most are Molecular Formulas)
-------------------	---------------------------------------

Hydrogen (gas)	H <sub>2</sub>
----------------	----------------

Common Substances	Formula (Most are Molecular Formulas)
-------------------	---------------------------------------

Oxygen (gas)	O <sub>2</sub>
--------------	----------------

Water	H <sub>2</sub> O
-------	------------------

Carbon Dioxide	CO <sub>2</sub>
----------------	-----------------

Hydrochloric Acid	HCl
-------------------	-----

## 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<sub>2</sub>O or glucose C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>).

**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<sup>+</sup> to Cl<sup>-</sup> 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**.

## 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 Oxy**gen** → Magnesium Oxi**de**

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**.

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

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

- The '3' from  $\text{Al}^{3+}$  becomes the subscript for O.
- The '2' from  $\text{O}^{2-}$  becomes the subscript for Al.
- This gives:  $\text{Al}_2\text{O}_3$

**Step 3: Check for Neutrality:** The total positive charge must equal the total negative charge.

- $\text{Al}_2\text{O}_3$ :  $(2 \text{ Al} \times +3) = +6$  and  $(3 \text{ O} \times -2) = -6$ . The compound is neutral ✓

**Important:** If the subscripts have a common factor, simplify them. For example,  $\text{Mg}_2\text{O}_2$  simplifies to  $\text{MgO}$ .

## 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...

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

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

**Examples:**

- 1 mole of Carbon atoms =  $6.022 \times 10^{23}$  C atoms
- 1 mole of Sulphur atoms =  $6.022 \times 10^{23}$  S atoms
- 1 mole of Water molecules =  $6.022 \times 10^{23}$   $\text{H}_2\text{O}$  molecules
- 1 mole of NaCl formula units =  $6.022 \times 10^{23}$  NaCl formula units

## Molar Mass: The Mass of One Mole

This is the most important connection between the microscopic (atoms) and macroscopic (grams) world.

The **mass of one mole** of any substance is equal to its **atomic mass, molecular mass, or formula mass** expressed in **grams (g)**.

This is called **Molar Mass**. The terms below are specific types of molar mass:

**a) Gram Atomic Mass (Molar Mass of an Element)**

The atomic mass of an element from the periodic table, expressed in grams, is the mass of **1 mole of atoms** of that element.

*Examples:*

- Atomic mass of Carbon (C) = **12 amu** → **Gram Atomic Mass = 12 g/mol** (1 mole of C atoms has a mass of 12 grams)
- Atomic mass of Sodium (Na) = **23 amu** → **Gram Atomic Mass = 23 g/mol**
- This is why 1 mole of Carbon (12 g) and 1 mole of Sulphur (32 g) have different masses—their individual atoms have different masses.

**b) Gram Molecular Mass (Molar Mass of a Molecule)**

The molecular mass of a compound, expressed in grams, is the mass of **1 mole of molecules** of that compound.

*Example:*

- Molecular mass of Water ( $\text{H}_2\text{O}$ ) = 18 amu → **Gram Molecular Mass = 18 g/mol** (1 mole of  $\text{H}_2\text{O}$  molecules has a mass of 18 grams)

**c) Gram Formula Mass (Molar Mass of an Ionic Compound)**

The formula mass of an ionic compound, expressed in grams, is the mass of **1 mole of formula units** of that compound.

*Example:*

- Formula mass of NaCl = 58.5 amu → **Gram Formula Mass = 58.5 g/mol** (1 mole of NaCl formula units has a mass of 58.5 grams)

## 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  $6.022 \times 10^{23}$  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** ( $6.022 \times 10^{23}$  atoms) of an element.

#### Examples:

- Carbon (C): Atomic mass = 12 amu → **Gram Atomic Mass = 12 g/mol**
- Sodium (Na): Atomic mass = 23 amu → **Gram Atomic Mass = 23 g/mol**
- This means **12 grams of Carbon** contains exactly  $6.022 \times 10^{23}$  carbon atoms.

#### b) Gram Molecular Mass (For Molecules)

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

**It contains: 1 mole of molecules** ( $6.022 \times 10^{23}$  molecules).

#### Examples:



- Water (H<sub>2</sub>O): Molecular mass = 18.016 amu → **Gram Molecular Mass = 18.016 g/mol**
- Glucose (C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>): Molecular mass = 180.096 amu → **Gram Molecular Mass = 180.096 g/mol**
- This means **18.016 grams of Water** contains exactly  $6.022 \times 10^{23}$  H<sub>2</sub>O 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** ( $6.022 \times 10^{23}$  formula units).

#### Examples:

- Sodium Chloride (NaCl): Formula mass = 58.5 amu → **Gram Formula Mass = 58.5 g/mol**
- Potassium Chloride (KCl): Formula mass = 74.5 amu → **Gram Formula Mass = 74.5 g/mol**
- This means **58.5 grams of NaCl** contains exactly  $6.022 \times 10^{23}$  NaCl formula units.

**Here's a simple table to clarify the differences:**

Term	What it Represents	What it Contains (1 mole of)	Example
<b>Gram Atomic Mass</b>	<b>An element</b>	<b>Atoms</b>	12 g of Carbon = $6.022 \times 10^{23}$ <b>C atoms</b>

Term	What it Represents	What it Contains (1 mole of)	Example
<b>Gram Molecular Mass</b>	A molecular compound/element	<b>Molecules</b>	18 g of Water $= 6.022 \times 10^{23} \times 6.022 \times 10^{23}$ $\text{H}_2\text{O}$ <b>molecules</b>
<b>Gram Formula Mass</b>	An ionic compound	<b>Formula Units</b>	58.5 g of Salt $= 6.022 \times 10^{23} \times 6.022 \times 10^{23}$ $\text{NaCl}$ <b>formula units</b>

### 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:

$$\text{Moles} = \text{Mass (g)} \div \text{Molar Mass (g/mol)}$$

### Connecting Moles to Number of Particles:

- **Number of Particles = Moles  $\times$  ( $6.022 \times 10^{23}$  particles/mol)**

### Connecting Mass directly to Number of Particles:

- **Number of Particles = [Mass (g)  $\div$  Molar Mass (g/mol)]  $\times$  ( $6.022 \times 10^{23}$ )**

## 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)  $\times$  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<sub>2</sub>, O<sub>2</sub>, I<sub>2</sub>), 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 ( $\text{N}_2$ ).

Molecular mass of  $\text{N}_2 = 14 \times 2 = \mathbf{28 \text{ amu}}$

Molar Mass = **28 g/mol**

c) **Sucrose ( $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ )**: It's a molecular compound.

Molecular mass =  $(12 \times 12) + (1 \times 22) + (16 \times 11) = 144 + 22 + 176 = \mathbf{342 \text{ amu}}$

Molar Mass = **342 g/mol**

## **Type 2: Calculating the Mass from Moles**

This uses the core formula directly: **Mass = Moles  $\times$  Molar Mass**

### **Example 6.4: Mass of Ozone ( $\text{O}_3$ )**

**Given:** Moles of  $\text{O}_3 = 9.05 \text{ mol}$

#### **Step 1: Find Molar Mass of $\text{O}_3$**

- Molar Mass =  $16 \times 3 = \mathbf{48 \text{ g/mol}}$

#### **Step 2: Apply Formula**

- Mass = Moles  $\times$  Molar Mass
- Mass =  $9.05 \text{ mol} \times 48 \text{ g/mol}$
- Mass = **434.4 g**

**Answer:** The mass of 9.05 moles of ozone is **434.4 grams**.

### **Example 6.5: Mass of Carbon Dioxide ( $\text{CO}_2$ )**

**Given:** Moles of  $\text{CO}_2 = 0.25 \text{ mol}$

#### **Step 1: Find Molar Mass of $\text{CO}_2$**

- Molar Mass =  $12 + (16 \times 2) = \mathbf{44 \text{ g/mol}}$

## Step 2: Apply Formula

- $\text{Mass} = 0.25 \text{ mol} \times 44 \text{ g/mol}$
- $\text{Mass} = \mathbf{11 \text{ g}}$
- **Answer:** The mass of 0.25 moles of  $\text{CO}_2$  is **11 grams**.

# 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 ( $\rightarrow$ ):** Means "yields" or "produces" and shows the direction of the reaction.
- 

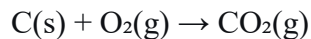
## Steps to Write a Chemical Equation:

### Step 1: Write the Word Equation

- Identify the reactants and products.
- *Example (Burning Coal):* Coal + Oxygen  $\rightarrow$  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):*



- **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:*  $\text{C(s)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}$

*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).

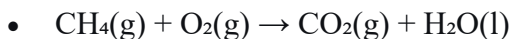
A **balanced chemical equation** is one where this rule is followed.

## 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<sub>2</sub>O 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!

### Let's Balance the Combustion of Methane (CH<sub>4</sub>):

#### Write the Skeleton (Unbalanced) Equation



#### Step 1: Count the Atoms on Each Side

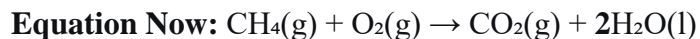
Atom	Reactant Side	Product Side	Balanced?
<b>C</b>	1	1	✓
<b>H</b>	4	2	✗
<b>O</b>	2	3 (2 from CO <sub>2</sub> + 1 from H <sub>2</sub> O)	✗

The equation is unbalanced for H and O.

#### Step 2: Balance One Element at a Time

Start with the most complex molecule (often the one with the most different atoms). Here, let's start with **H**.

There are 4 H on the left (in CH<sub>4</sub>) and only 2 on the right (in H<sub>2</sub>O). To get 4 H on the right, put a coefficient of **2** in front of H<sub>2</sub>O.



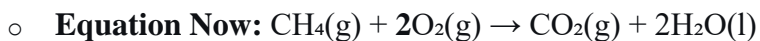
### Recount All Atoms:

Atom	Reactant Side	Product Side	Balanced?
<b>C</b>	1	1	✓
<b>H</b>	4	4 (2 molecules × 2 H each)	✓
<b>O</b>	2	4 (2 from CO <sub>2</sub> + 2 from 2H <sub>2</sub> O)	✗

Now Carbon and Hydrogen are balanced, but Oxygen is not.

### Step 3: Balance the Next Element

There are 4 O on the right but only 2 on the left (in O<sub>2</sub>). To get 4 O on the left, put a coefficient of **2** in front of O<sub>2</sub>.



### Step 4: Do a Final Check

Atom	Reactant Side	Product Side	Balanced?
<b>C</b>	1	1	✓
<b>H</b>	4	4	✓
<b>O</b>	4 (2 molecules × 2 O each)	4 (2 from CO <sub>2</sub> + 2 from 2H <sub>2</sub> O)	✓

**The equation is now balanced!**

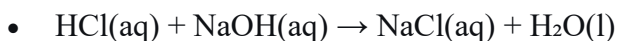
**Final Balanced Equation:**  $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$



# Ionic Equations

## .Steps to Write an Ionic Equation:

### Step 1: Write the Balanced Molecular Equation



### 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:



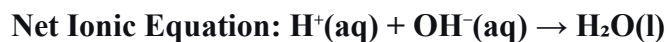
### Step 3: Identify and Remove Spectator Ions

**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:  $\text{Na}^+(\text{aq})$  and  $\text{Cl}^-(\text{aq})$  are spectator ions.

### Step 4: Write the Net Ionic Equation

This shows only the particles that actually react.



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

## Molecular Formula vs. Structural Formula

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

- *Example:* C<sub>4</sub>H<sub>10</sub>

**Structural Formula:** Shows **how the atoms are connected and bonded** to each other.

- *Example:* CH<sub>3</sub>-CH<sub>2</sub>-CH<sub>2</sub>-CH<sub>3</sub> (for n-Butane)

## How to Write a Molecular Formula from a Structural Formula

**Example: For CH<sub>3</sub>-CH<sub>2</sub>-CH<sub>2</sub>-OH**

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

**Count the Atoms:**

**Carbon (C):** There are 3 C atoms (one in each CH<sub>3</sub> and CH<sub>2</sub> group).

**Hydrogen (H):** Let's count carefully:

- CH<sub>3</sub>- has 3 H
- -CH<sub>2</sub>- has 2 H
- -CH<sub>2</sub>- has 2 H
- The -OH has 1 H

$$\text{Total H} = 3 + 2 + 2 + 1 = 8$$

**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<sub>3</sub>H<sub>8</sub>O**

**Concept Assessment 6.8 (Answers):**

1. **CH<sub>3</sub>-CH<sub>2</sub>-OH**

- C: 2, H: 6 (3+2+1), O: 1 → **C<sub>2</sub>H<sub>6</sub>O**

2. **CH<sub>3</sub>-CH<sub>2</sub>-NH<sub>2</sub>**

- C: 2, H: 7 (3+2+2), N: 1 → **C<sub>2</sub>H<sub>7</sub>N**

3. **CH<sub>3</sub>-CO-CH<sub>3</sub>**

- C: 3, H: 6 (3+0+3), O: 1 → **C<sub>3</sub>H<sub>6</sub>O**

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.