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 H_2O_2, but the simplest ratio of H to O is 1:1. So, its empirical formula is **HO**.

- **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**.

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 **H_2O_2** (2 hydrogen atoms and 2 oxygen atoms).
- **Glucose:** The molecular formula is **$C_6H_{12}O_6$** (6 carbon, 12 hydrogen, and 6 oxygen atoms).
- **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: $(CH_2O) \times 6 = C_6H_{12}O_6$.

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).

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_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.

Quick Reference Table (From Your Book)

Common Substances	Formula (Most are Molecular Formulas)
Hydrogen (gas)	H_2

Common Substances	Formula (Most are Molecular Formulas)
Oxygen (gas)	O_2
Water	H_2O
Carbon Dioxide	CO_2
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_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**.

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**.

- *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.

- The '3' from Al^{3+} becomes the subscript for O.
- The '2' from O^{2-} becomes the subscript for Al.
- This gives: Al_2O_3

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

- 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.

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 $\mathbf{6.022 \times 10^{23}}$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

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 (H_2O) = 18 amu → **Gram Molecular Mass** = **18 g/mol** (1 mole of H_2O 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 6.022×10^{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:

- 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 6.022×10236.022×10^{23} carbon atoms.

b) Gram Molecular Mass (For Molecules)

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

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

Examples:

- Water (H_2O): Molecular mass = 18.016 amu → **Gram Molecular Mass = 18.016 g/mol**
- Glucose ($C_6H_{12}O_6$): Molecular mass = 180.096 amu → **Gram Molecular Mass = 180.096 g/mol**
- This means **18.016 grams of Water** contains exactly $6.022\times10236.022\times1023$ 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 ($6.022\times10236.022\times1023$ 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 $6.022\times10236.022\times1023$ NaCl formula units.

Here's a simple table to clarify the differences:

Term	What it Represents	What it Contains (1 mole of)	Example
Gram Atomic Mass	An **element**	**Atoms**	12 g of Carbon = 6.022×10^{23}C **atoms**

Term	What it Represents	What it Contains (1 mole of)	Example
Gram Molecular Mass	A **molecular compound/ element**	**Molecules**	18 g of Water = $6.022\times10^{23}6.022\times10^{23}$ H_2O **molecules**
Gram Formula Mass	An **ionic compound**	**Formula Units**	58.5 g of Salt = $6.022\times10^{23}6.022\times10^{23}$ 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:

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

Connecting Moles to Number of Particles:

- **Number of Particles = Moles × ($6.022\times10236.022\times10_{23}$ particles/mol)**

Connecting Mass directly to Number of Particles:

- **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

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

Step 2: Apply Formula

- Mass = 0.25 mol × 44 g/mol
- Mass = **11 g**
- **Answer:** The mass of 0.25 moles of 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 (→):** 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!

Let's Balance the Combustion of Methane (CH_4):

Write the Skeleton (Unbalanced) Equation

- $CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$

Step 1: Count the Atoms on Each Side

Atom	Reactant Side	Product Side	Balanced?
C	1	1	✓
H	4	2	✕
O	2	3 (2 from CO_2 + 1 from H_2O)	✕

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_4) and only 2 on the right (in H_2O). To get 4 H on the right, put a coefficient of **2** in front of H_2O.

Equation Now: $CH_4(g) + O_2(g) \rightarrow CO_2(g) + \mathbf{2}H_2O(l)$

Recount All Atoms:

Atom	Reactant Side	Product Side	Balanced?
C	1	1	✓
H	4	4 (2 molecules × 2 H each)	✓
O	2	4 (2 from CO_2 + 2 from $2H_2O$)	✗

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_2). To get 4 O on the left, put a coefficient of **2** in front of O_2.

- **Equation Now:** $CH_4(g) + \mathbf{2}O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

Step 4: Do a Final Check

Atom	Reactant Side	Product Side	Balanced?
C	1	1	✓
H	4	4	✓
O	4 (2 molecules × 2 O each)	4 (2 from CO_2 + 2 from $2H_2O$)	✓

The equation is now balanced!
Final Balanced Equation: $CH_4\,(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

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:

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

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

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

Molecular Formula vs. Structural Formula

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

- *Example:* C_4H_{10}

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

- *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**

Concept Assessment 6.8 (Answers):

1. **CH_3-CH_2-OH**

- C: 2, H: 6 (3+2+1), O: 1 → **C_2H_6O**

2. **CH_3-CH_2-NH_2**

- C: 2, H: 7 (3+2+2), N: 1 → **C_2H_7N**

3. **CH_3-CO-CH_3**

- C: 3, H: 6 (3+0+3), O: 1 → **C_3H_6O**

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.