2.4 MOLE CONCEPT

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Learning outcomes

After the completion of the chapter, the students should be able to:

- Define a mole and explain its relation with mass, volume, and number of particles.
- Interpret a balanced chemical equation in terms of interacting moles, representative particles, masses, and volume of gases (at STP).
- Perform stoichiometric calculations involving the above quantities.

Introduction

Imagine trying to count every grain of rice in a sack by picking them up one at a time. Impossible, right? Shopkeepers solve this by weighing rice instead of counting grains. Chemists face the same problem — atoms and molecules are far too small and far too numerous to count one by one. Their solution is a special counting unit called the mole.

Why do we even need a 'mole'?

A single atom of hydrogen is so light that even a whole gram of hydrogen gas contains more atoms than there are grains of sand on Earth. We cannot weigh a single atom on a lab balance, so chemists needed a bridge between the invisible world of atoms and the visible world of grams that we can actually measure. That bridge is the mole.

Before defining the mole formally, get comfortable with these two building blocks:

Gram Atomic Mass and Gram Molecular Mass

Gram Atomic Mass (Gram Atom)

The atomic mass of an element expressed in grams is called its gram atomic mass or gram atom.

Examples

- Atomic mass of sodium (Na) = 23 amu
  Therefore, 23 g of Na = 1 gram atom of Na
  46 g of Na = 2 gram atoms of Na
- Atomic mass of calcium (Ca) = 40 amu
  Therefore, 40 g of Ca = 1 gram atom of Ca

Gram molecular mass (gram molecule)

The molecular mass of a substance expressed in grams is called its gram molecular mass or gram molecule.

Examples

- Molecular mass of CO_2 = 44 amu
  Therefore, 44 g of CO_2 = 1 gram molecule of CO_2
- Molecular mass of H_2 = 2 amu
  Therefore, 4 g of H_2 = 2 gram molecules of H_2
- Molecular mass of (water) H_2O = 1 x 2 +16 = 18 amu
  18 g water = 1 gram molecule of water
  72 g water = 4 gram molecules of water

Mole Concept

Scientists discovered that:

- 1 gram atom of every element contains the same number of atoms.
- 1 gram molecule of every substance contains the same number of molecules.
  This fixed number is called **Avogadro's number**.

Definition of Mole

A **mole** is the amount of substance that contains **6.022×10^{23} particles**.

Depending on the substance, a particle can be:

Atom:	e.g., 1 mole of Na = 6.022×10^{23} sodium atoms
Molecule:	e.g., 1 mole of H_2O = 6.022×10^{23} water molecules
Ion:	e.g., 1 mole of Na^+ = 6.022×10^{23} sodium ions
Electron:	e.g., 1 mole of e^- = 6.022×10^{23} electrons
Formula unit:	e.g., 1 mole of NaCl = 6.022×10^{23} formula units

Think of it like a chemist's 'dozen' — just as a dozen always means 12, a mole always means 6.022×10^{23} particles.

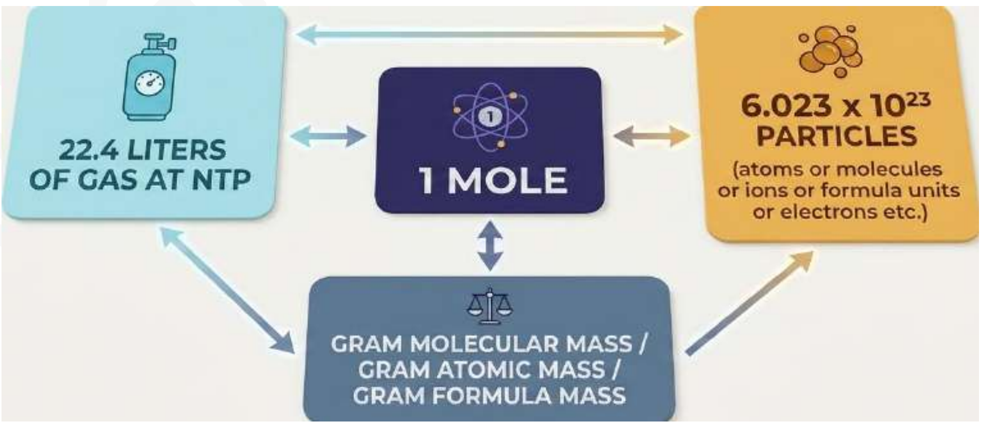
Did You Know?

6.022×10^{23} is an unimaginably huge number! If every student in your class counted one particle every second without stopping, it would take far longer than the age of the universe to count one mole. That's how big Avogadro's number is!

Another definition states:

One mole is the amount of substance containing as many particles as there are atoms in exactly 12 g of the carbon-12 isotope.

1 mole = 1 gram atom/ 1 gram molecule
= gram atomic mass/ gram molecular mass/ gram formula mass
= 6.022×10^{23} particles (atoms/molecules/ions/formula units, electrons, etc.)
= 22.4 litres of gas at NTP



Or,

$$\text{no. of moles} = \frac{\text{mass in gram}}{\text{atomic or molecular mass}} = \frac{\text{volume of gas at NTP}}{22.4} = \frac{\text{number of particles}}{6.023 \times 10^{23}}$$

Worked Example

1. How many molecules are contained in 0.35 mole of N_2? 1

Solution:

1 mole of N_2 molecule contains 6.022×10^{23} molecules

0.35 moles of N_2 molecule contains $0.35 \times 6.023 \times 10^{23}$ molecules

$= 2.1 \times 10^{23}$ molecules of N_2

1. The absolute mass of 1 amu is

a. 1g *b. 1 mg* *c. 1.66×10^{-24} g* *d. 1.66×10^{-27} g*

Solution:

We know that 1 amu is the mass of 1/12th of the C^{12} isotope.

We also know that, 6.023×10^{23} C^{12} isotopes = 12 g

$$1\ C^{12} \text{ isotope} = \frac{12}{6.023 \times 10^{23}} = 1.992 \times 10^{-23} \text{ g}$$

$$1/12\text{th of } C^{12} \text{ isotope} = \frac{1.992 \times 10^{-23}}{12} = 1.660 \times 10^{-24} \text{ g}$$

Therefore, 1 amu = 1.660×10^{-24} g.

Key Points to Remember

- Mole is the SI unit of amount of substance.
- One mole contains 6.022×10^{23} particles.
- One mole of gas occupies 22.4 L at STP.
- Molar mass is numerically equal to atomic or molecular mass but expressed in grams.
- Avogadro's number connects microscopic particles with measurable laboratory quantities.

Summary

The mole concept provides a bridge between the microscopic world of atoms and molecules and the macroscopic world of laboratory measurements. It allows chemists to calculate masses, volumes, and numbers of particles accurately. Understanding the mole concept is essential for solving numerical problems in chemistry and for studying chemical reactions quantitatively.

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