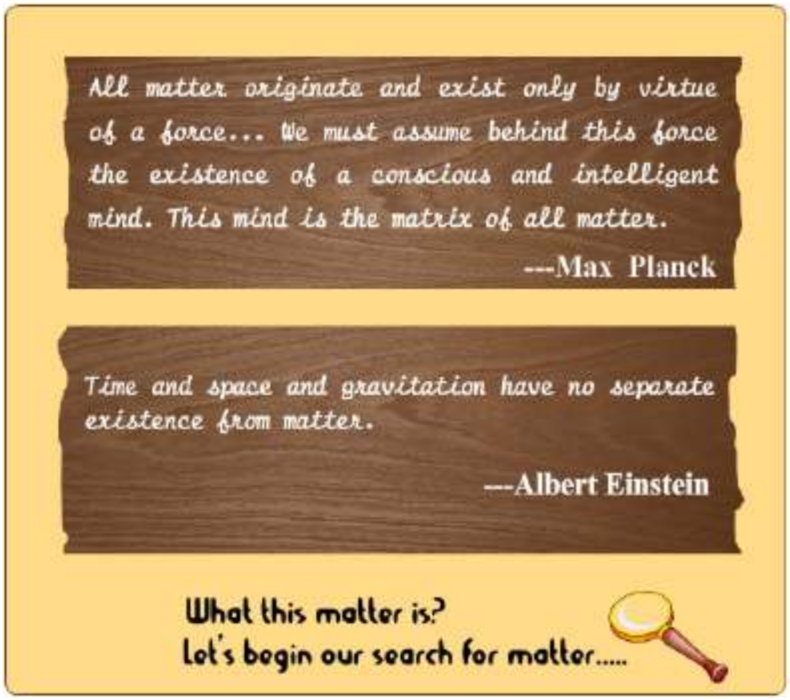
Some basic concepts

Nature of Matter



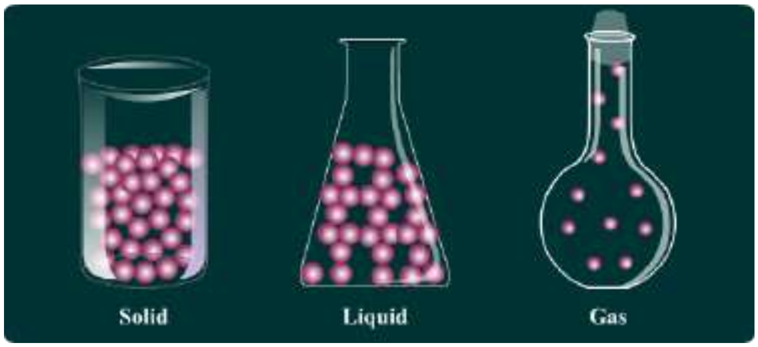
- Matter is anything that has mass and occupies space.
- **Three physical states of matter**

- **Characteristics of solid**
  - Definite volume
  - Definite shape

- **Characteristics of liquid**
  - Definite volume
  - Indefinite shape

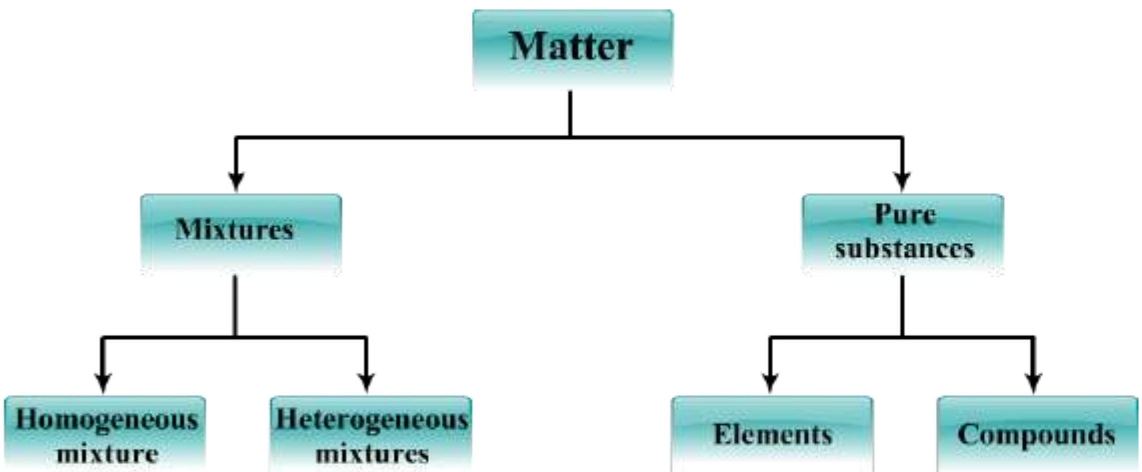
- **Characteristics of gas**
    - Indefinite volume
    - Indefinite shape

- Arrangement of particles in the three states



- **Inter-conversion between the three states**

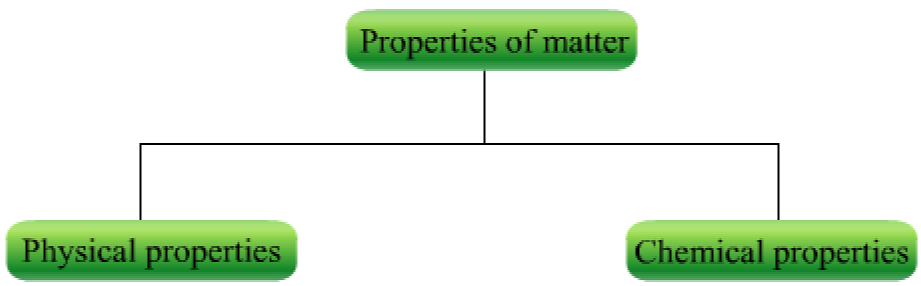
- **Classification of matter in macroscopic level**



- **Mixtures**
    - Components are present in any ratio.
    - Homogeneous mixture – Uniform composition throughout the mixture.
    - Heterogeneous mixture – Non-uniform composition throughout the mixture.
    - Components can be separated by physical methods such as hand picking, filtration, crystallization, distillation, etc.

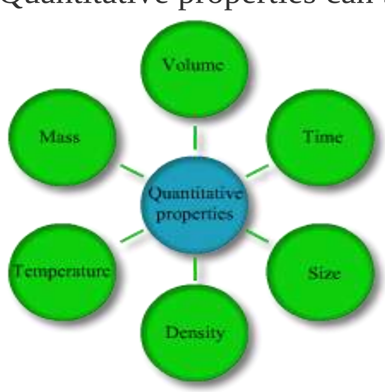
- **Pure substances**
  - Fixed composition
  - Constituents cannot be separated by simple physical methods.
  - Elements contain only one type of particles – atoms (example Na, K, Cu, C, Ag) or molecules [Example: H_2, N_2, O_2, F_2]
  - Compounds are formed by the combination of two or more atoms of different elements. [Example: water (H_2O), carbon dioxide (CO_2)]
  - Constituents of compounds cannot be separated by physical methods; they can be separated by chemical methods only.

Properties of Matter



- **Physical properties**
  - Properties which can be measured or observed without changing the identity or composition of the substance
  - Example – Colour, odour, melting point, boiling point, density, etc.

- **Chemical Properties**
  - Properties in which chemical change in the substance takes place
  - Examples – Characteristic reactions of different substances such as acidity, basicity, combustibility, reactions with other elements and compounds
  - Quantitative properties can be of the following types:



Laws of Chemical Combination

- **Law of Conservation of Mass:**

Matter can be neither created nor destroyed.

- **Law of Definite Proportions:**

A given compound always contains exactly the same proportion of elements by weight.

- **Law of Multiple Proportions:**

If two elements can combine to form more than one compound, then the masses of one element that combines with a fixed mass of the other element are in the ratio of small whole numbers.

- **Gay Lussac's Law of Gaseous Volumes:**

When gases combine or are produced in a chemical reaction, they do so in a simple ratio by volume, provided all the gases are at the same temperature and pressure.

- **Avogadro Law:**

Equal volumes of gases at the same temperature and pressure should contain equal number of molecules.

Dalton's Atomic Theory

- **Postulates:**

- All matter is made of very tiny indivisible particles called atoms.
- All the atoms of a given element are identical in mass and chemical properties whereas those of different elements have different masses and chemical properties.
- Atoms of different elements combine in a fixed whole number ratio to form compounds.
- Chemical reactions involve reorganization of atoms. Atoms are neither created nor destroyed in a chemical reaction.

- The laws of chemical combination could be explained by Dalton's atomic theory.

Atomic and Molecular Masses

- **Atomic mass:**

- The mass of an atom
- One atomic mass unit (1 amu) = Mass equal to one-twelfth of the mass of one carbon-12 atom

1 amu = 1.66056×10^{-24} g

- Nowadays, 'u' (unified mass) has replaced 'amu'.
- Average atomic mass = $\sum$(Mass of isotope × Relative abundance)

Example

The relative abundance of two isotopes of copper, having atomic masses 62.93 u and 64.94 u, are 69.09% and 30.91% respectively. Calculate the average atomic mass of copper.

Solution:

Average atomic mass of copper = $\left(62.93 \times \frac{69.09}{100}\right) + \left(64.94 \times \frac{30.91}{100}\right)$

= 63.55 u

- **Molecular Mass:**

- Sum of the atomic masses of all the elements present in a molecule
- Example − Molecular mass of CO_2 = 1 × Atomic mass of carbon + 2 × Atomic mass of oxygen

= (1 × 12.011 u) + (2 × 16.00 u)

= 12.011 u + 32.00 u

= 44.011 u

- **Formula Mass:**

- Sum of the masses of all the atoms present in a formula unit of a compound
- Used for compounds whose constituent particles are ions
- Example − Formula mass of sodium chloride (NaCl)

= Atomic mass of sodium + Atomic mass of chlorine

= 23.0 u + 35.5 u => 58.5 u

Mole Concept and Molar Masses

- 1 mole of any substance can be defined as:

- Amount of a substance that contains as many particles (atoms, molecules or ions) as there are atoms in 12 g of the ^{12}C isotope
- Avogadro number or Avogadro constant (N_A); equal to 6.022×10^{23} particles
- Example − 1 mole of oxygen atoms = 6.022×10^{23} atoms

1 mole of carbon dioxide molecules = 6.022×10^{23} molecules

1 mole of sodium chloride = 6.022×10^{23} formula units of sodium chloride

Molar mass of a substance can be defined as:

- Mass of one mole of a substance in grams
- Numerically equal to atomic/molecular/formula mass in u.
- Example − Molar mass of CO_2 = 44.011 g mol^{-1}

Molar mass of NaCl = 58.5 g mol^{-1}

Examples

1. What number of moles contains 3.011×10^{23} molecules of glucose?

Solution:

1 mole of glucose is equivalent to 6.022×10^{23} molecules of glucose.

Hence, 3.011×10^{23} molecules of glucose will be present in

$= \frac{1 \times 3.011 \times 10^{23}}{6.022 \times 10^{23}}$ mol = 0.5 mol (of glucose)

Thus, 0.5 mole of glucose contains 3.011×10^{23} molecules of glucose.

2. What is the mass of a mole of fluorine molecule?

Solution:

1 mole of fluorine molecule contains 6.022×10^{23} molecules and weighs 38 g.

Therefore, mass of a fluorine molecule = $\frac{38}{6.022 \times 10^{23}}$ g

= 6.31×10^{-23} g

Percentage Composition

Mass percent of an element = $\frac{\text{Mass of that element in the compound} \times 100\%}{\text{Molar mass of the compound}}$

Example

What is the mass percent of oxygen in potassium nitrate? (Atomic mass of K = 39.10 u, atomic mass of N = 14.007 u, atomic mass of O = 16.00 u)

Solution:

Atomic mass of K = 39.10 u (Given)

Atomic mass of N = 14.007 u (Given)

Atomic mass of O = 16.00 u (Given)

Therefore, molar mass of potassium nitrate (KNO_3)

= 39.10 + 14.007 + 3(16.00)

= 101.107 g

Therefore, mass percent of oxygen in KNO_3

$$= \frac{\text{Mass of oxygen in } KNO_3 \times 100\%}{\text{Molar mass of } KNO_3}$$

$$= \frac{3 \times 16.00}{101.107} \times 100\%$$

= 47.47% (approx)

- Empirical formula and molecular formula:

Empirical formula	Molecular formula
Represents the simplest whole number ratio of various atoms present in a compound	Represents the exact number of different types of atoms present in a molecule of a compound

- Empirical formula is determined if mass % of various elements are known.
- Molecular formula is determined from empirical formula if molar mass is known.

Example

A compound contains 92.26% carbon and 7.74% hydrogen. If the molar mass of the compound is 26.038 g mol^{-1}, then what are its empirical and molecular formulae?

Solution:

Mass percent of carbon (C) = 92.26% (Given)

Mass percent of hydrogen (H) = 7.74% (Given)

Number of moles of carbon present in the compound = $\frac{92.26}{12.011}$

= 7.68 mol

Number of moles of hydrogen present in the compound $= \frac{7.74}{1.008}$

= 7.68 mol

Thus, in the given compound, carbon and hydrogen are present in the ratio C : H = 7.68 : 7.68

= 1 : 1

Therefore, the empirical formula of the compound is CH.

Empirical formula mass of CH = (12.011 + 1.008)g

= 13.019 g

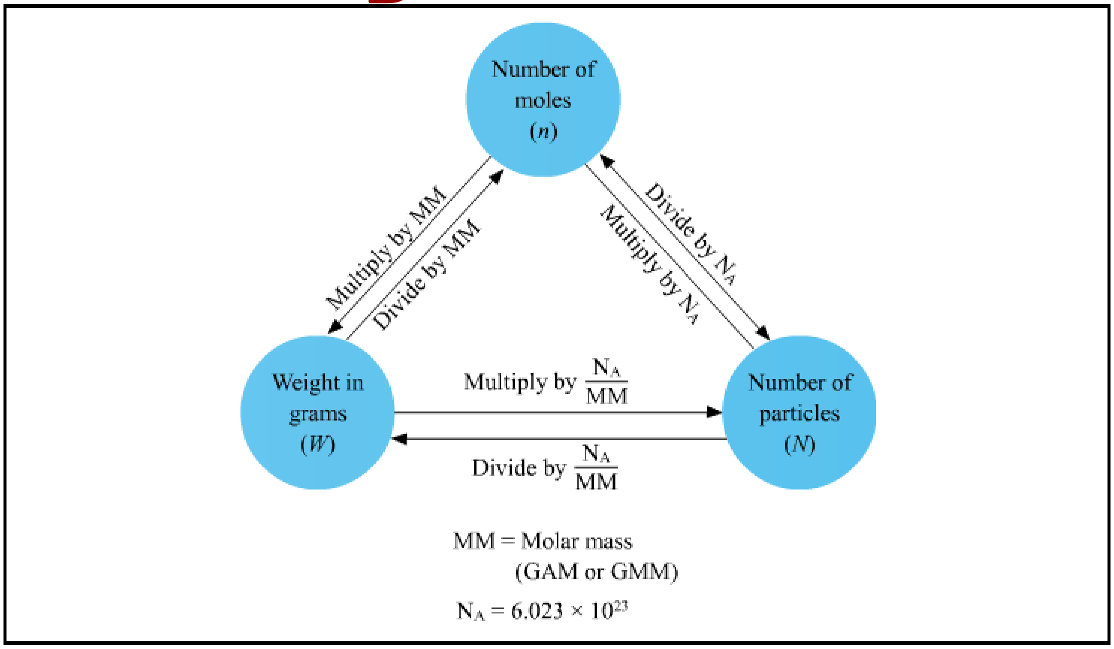
Molar mass of the compound = 26.038 g (Given)

Therefore, $n = \frac{\text{Molar Mass}}{\text{Empirical formula mass}}$

$= \frac{26.038 \text{ g}}{13.019 \text{ g}}$

= 2Hence, the molecular mass of the compound is $(CH)_n$, i.e., $(CH)_2$ or C_2H_2.

Interconversion among number of moles, mass and number of molecules



Stoichiometric Calculations in Balanced Chemical Equations

- An example of a balanced chemical equation is given below.

$$C_3H_{8(g)} + 5O_{2(g)} \longrightarrow 3CO_{2(g)} + 4H_2O_{(l)}$$

From the above balanced chemical equation, the following information is obtained:

- One mole of $C_3H_{8(g)}$ reacts with five moles of $O_{2(g)}$ to give three moles of $CO_{2(g)}$ and four moles of $H_2O_{(l)}$.
- One molecule of $C_3H_{8(g)}$ reacts with five molecules of $O_{2(g)}$ to give three molecules of $CO_{2(g)}$ and four molecules of $H_2O_{(l)}$.
- 44 g of $C_3H_{8(g)}$ reacts with (5 × 32 = 160) g of $O_{2(g)}$ to give (3 × 44 = 132) g of CO_2 and (4 × 18 = 72) g of H_2O.
- 22.4 L of $C_3H_{8(g)}$ reacts with (5 × 22.4) L of $O_{2(g)}$ to give (3 × 22.4) L of O_2 and (4 × 22.4) L of H_2O.

Example

Nitric acid (HNO_3) is commercially manufactured by reacting nitrogen dioxide (NO_2) with water (H_2O). The balanced chemical equation is represented as follows:

$$3NO_{2(g)} + H_2O_{(l)} \longrightarrow 2HNO_{3(aq)} + NO_{(g)}$$

Calculate the mass of NO_2 required for producing 5 moles of HNO_3.

Solution:

According to the given balanced chemical equation, 3 moles of NO_2 will produce 2 moles of HNO_3.

Therefore, 2 moles of HNO_3 require 3 moles of NO_2.

Hence, 5 moles of HNO_3 require $= \frac{3}{2} \times 5$ moles of NO_2

= 7.5 moles of NO_2

Molar mass of NO_2 = (14 + 2 × 16) g mol^{-1}

= 46 g mol^{-1}

Thus, required mass of NO_2 = (7.5 × 46) g mol^{-1}

= 345 g mol^{-1}

- Limiting reagent or limiting reactant:

- Reactant which gets completely consumed when a reaction goes to completion
- So called because its concentration limits the amount of the product formed

Example

Lead nitrate reacts with sodium iodide to give lead iodide and sodium nitrate in the following manner:

$$Pb(NO_3)_2 + 2NaI \longrightarrow PbI_2 + 2NaNO_3$$

What amount of sodium nitrate is obtained when 30 g of lead nitrate reacts with 30g of sodium iodide?

Solution:

Molar mass of $Pb(NO_3)_2 = 207 + \left[\{14 + (16 \times 3)\} \times 2\right]$

= 331 g mol^{-1}

Molar mass of NaI = (23 + 127) = 150 g mol^{-1}

According to the given equation, 1 mole of $Pb(NO_3)_2$ reacts with 2 moles of NaI, i.e.,

331 g of $Pb(NO_3)_2$ reacts with 300 g of NaI to give PbI_2 and $NaNO_3$

Thus, $Pb(NO_3)_2$ is the limiting reagent.

Therefore, 30 g of Pb $(NO_3)_2$ $= \dfrac{30}{331} = 0.09$ mole

According to the equation, 0.09 mole of $Pb(NO_3)_2$ will give (2 × 0.09) mole of $NaNO_3$ = 0.18 mole of $NaNO_3$.

- **Reactions in solutions:**

Ways for expressing the concentration of a solution −

- Mass per cent or weight per cent (w/w%)

Mass per cent $= \dfrac{\text{Mass of solute}}{\text{Mass of solution}} \times 100\%$

Example

4.4 g of oxalic acid is dissolved in 200 mL of a solution. What is the mass per cent of oxalic acid in the solution? (Density of the solution = 1.1 g mL^{-1})

Solution:

Density of the solution = 1.1 g mL^{-1}

So the mass of the solution = (200 mL) × (1.1 g mL^{-1})

= 220 g

Mass of oxalic acid = 4.4 g

Therefore, mass per cent of oxalic acid in the solution

$$= \frac{\text{Mass of oxalic acid}}{\text{Mass of the solution}} \times 100\%$$
$$= \frac{4.4\text{g}}{220\text{g}} \times 100\%$$
$$= 2\%$$

- Mole fraction:

If a substance 'A' dissolves in a substance 'B', then mole fraction of A $= \frac{\text{Number of moles of A}}{\text{Number of moles of the solution}}$

$$= \frac{n_A}{n_A + n_B}$$

Mole fraction of B $= \frac{\text{Number of moles of B}}{\text{Number of moles of the solution}}$

$$= \frac{n_B}{n_A + n_B}$$

n_A – Number of moles of A & n_B – Number of moles of B

Example

A solution is prepared by dissolving 45 g of a substance **X** (molar mass = 25 g mol^{-1}) in 235 g of a substance **Y**(molar mass = 18 g mol^{-1}). Calculate the mole fractions of **X** and **Y**.

Solution:

Moles of X, n_X = $\frac{45}{25}$

= 1.8 mol

Moles of Y, n_Y = $\frac{235}{18}$

= 13.06 mol

Therefore, mole fraction of X, n_X $= \frac{1.8}{1.8 + 13.06}$

$= \frac{1.8}{14.86}$

$= 0.121$

And, mole fraction of Y, $n_Y = 1 - n_X$

= 1 − 0.121

= 0.879

- Molarity:

Number of moles of a solute in 1 L of a solution

Molarity (M) = $\frac{\text{Number of moles of solute}}{\text{Volume of solution in litres}}$

Molarity equation:

$M_1V_1 = M_2V_2$

M_1 = Molarity of a solution when its volume is V_1

M_2 = Molarity of the same solution when its volume is V_2

Examples

1. 10g of HCl is dissolved in enough water to form 500 mL of the solution. Calculate the molarity of the solution.

Solution:

Molar mass of HCl = 36.5 g mol^{-1}

So the moles of HCl = $\frac{10}{36.5}$ mol

= 0.274 mol

Volume of the solution = 500 mL = 0.5 L

Therefore, molarity = $\frac{\text{Number of moles of HCl}}{\text{Volume of solution in litres}}$

$= \frac{0.274 \text{ mol}}{0.5 \text{ L}}$

= 0.548 M

2. Commercially available concentrated HCl contains 38% HCl by mass. What volume of concentrated HCl is required to make 2.5 L of 0.2 M HCl? (Density of the solution = 1.19 g mL^{-1})

Solution:

38% HCl by mass means that 38g of HCl is present in 100 g of the solution.

Moles of HCl = $\frac{38}{36.5} = 1.04 \text{ mol}$

Volume of the solution $= \frac{\text{Mass}}{\text{Density}}$

$= \frac{100\text{g}}{1.19\text{g mL}^{-1}}$

= 84.03 mL

= 0.08403L

Therefore, molarity of the solution = $\frac{1.04}{0.08403\text{ L}}$

= 12.38 M

According to molarity equation,

$M_1V_1 = M_2V_2$

Here,

M_1 = 12.38 M

M_2 = 0.2 M

V_2 = 2.5 L

Now, $M_1V_1 = M_2V_2$

$$\Rightarrow V_1 = \frac{M_2V_2}{M_1}$$
$$= \frac{0.2 \times 2.5}{12.38}$$
$$= 0.0404 \text{ L} \quad (\text{approx})$$

Hence, required volume of HCl = 0.0404 L

- Molality:

Number of moles of solute present in 1 kg of solvent

Molality (m) = $\dfrac{\text{Number of moles of solute}}{\text{Mass of solvent in kg}}$

Example

What is the molality of a solution of glucose in water, which is labelled as 15% (w/w)?

Solution:

15% (w/w) solution means that 15 g of glucose is present in 100 g of the solution, i.e., (100 − 15) g = 85 g of water = 0.085 kg of water

Moles of glucose = $\dfrac{15\text{g}}{180\text{ g mol}^{-1}}$

= 0.083 mol

Therefore, molality of the solution $= \dfrac{0.083\text{ mol}}{0.085\text{ kg}}$

= 0.976 m

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