1. SOME BASIC CONCEPTS IN CHEMISTRY

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CHEMISTRY

The branch of science which deals with the composition, structure and properties of matter and the changes which matter undergoes.

BRANCHES OF CHEMISTRY

- ✔ Inorganic Chemistry
- ✔ Organic Chemistry
- ✔ Physical Chemistry
- ✔ Analytical Chemistry
- ✔ Polymer Chemistry
- ✔ Biochemistry
- ✔ Medicinal Chemistry
- ✔ Industrial Chemistry
- ✔ Hydrochemistry
- ✔ Electrochemistry
- ✔ Green Chemistry
- ✔ Pharmaceutical Chemistry

MATTER

Anything which occupies space and has mass and can be perceived by our senses.

Eg: Air, Water, Stone etc.

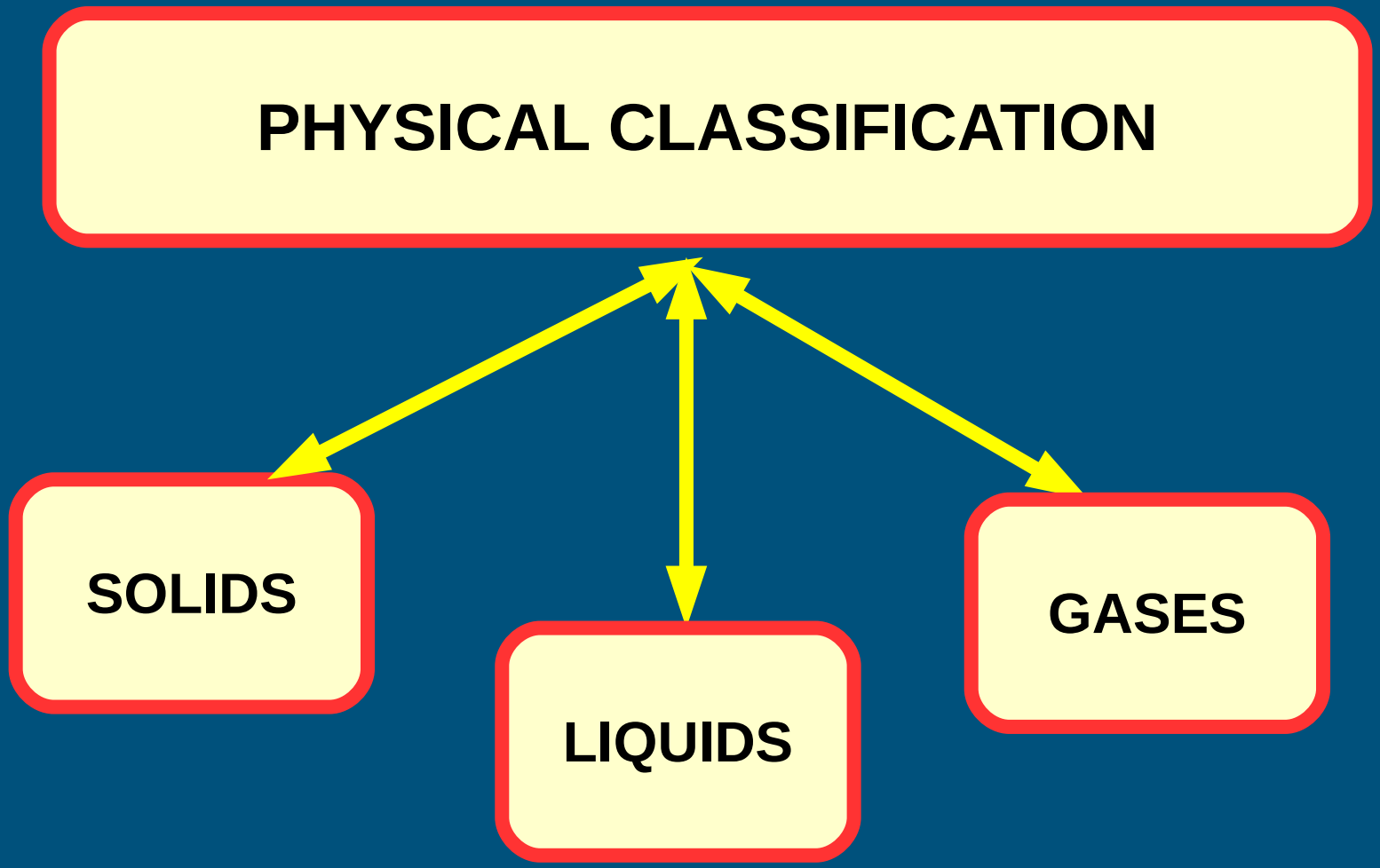
STATES OF MATTER

- Solid state
- Liquid state
- Gaseous state
- Plasma state
- Bose-Einstein condensate
- Fermionic condensate

CLASSIFICATION OF MATTER

1. PHYSICAL CLASSIFICATION

2. CHEMICAL CLASSIFICATION


PHYSICAL CLASSIFICATION
SOLIDS
LIQUIDS
GASES

SOLIDS

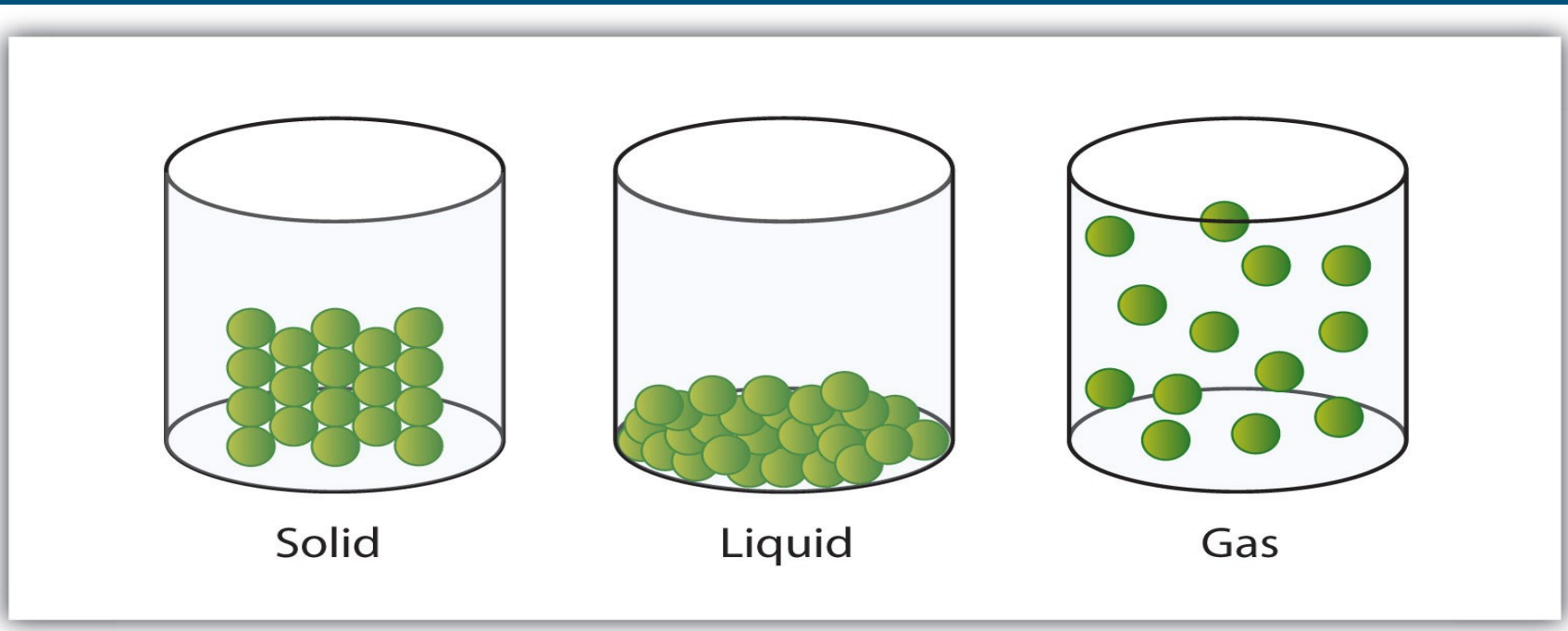
- Have definite mass, volume and shape.
- The constituent particles are tightly packed
- Intermolecular distances are short.
- Intermolecular forces are strong.
- Constituent particles have fixed positions.
- They can only oscillate about their mean positions.
- Incompressible and rigid.

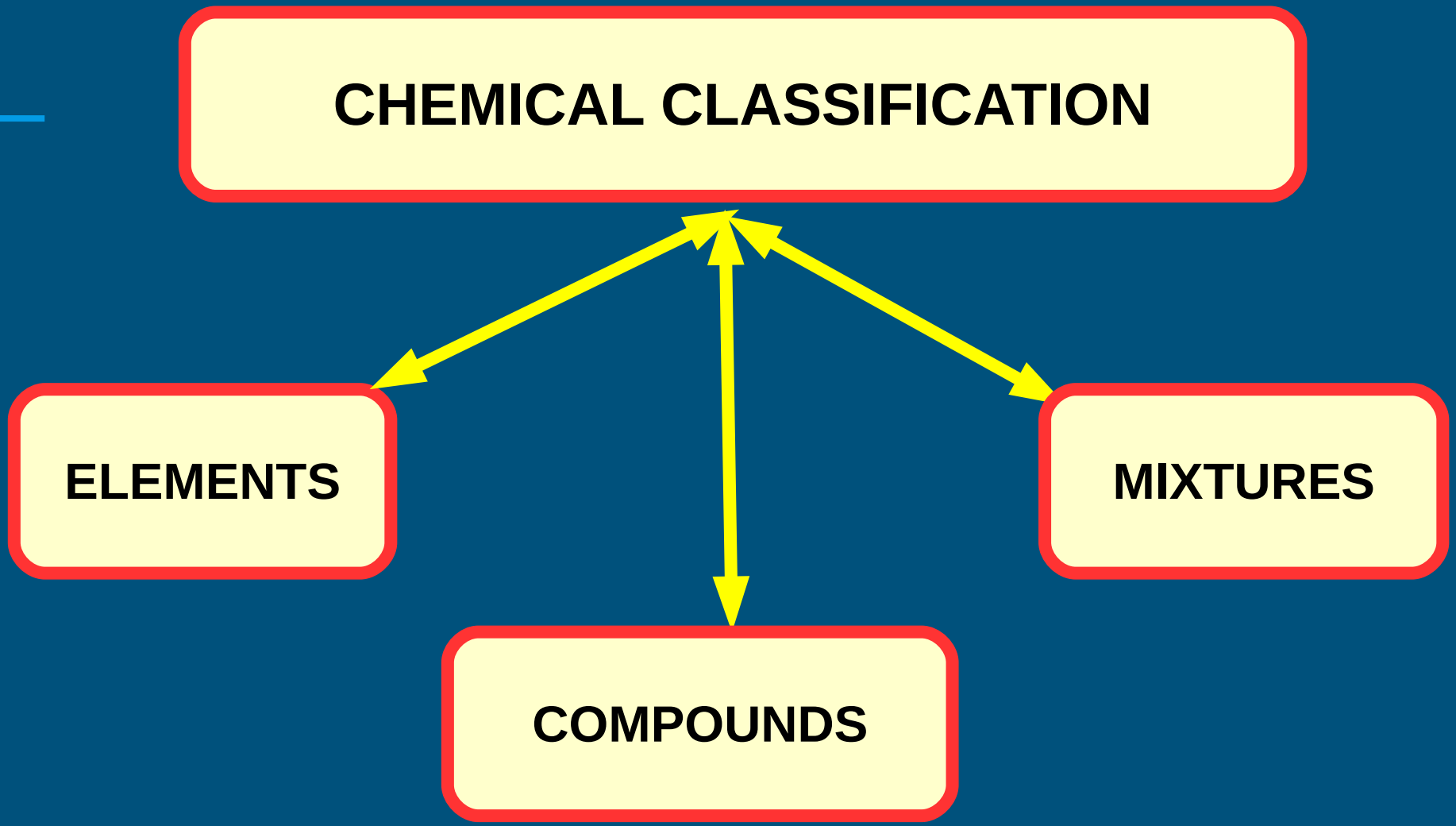
LIQUIDS

- Liquids are not rigid, but have definite volume.
- They do not have a definite shape.
- They take the shape of the container in which they are placed.
- The constituent particles are loosely packed.
- The intermolecular distances will be large as compared to that of solids.
- Intermolecular forces are weak compared to solids.
- The rate of diffusion will be greater than solids.

GASES

- Gases have neither definite shape nor definite volume.
- The particles are far apart from one another.
- The intermolecular distances are very large.
- The intermolecular forces of attraction are very weak.
- Gases are highly compressible.
- Gases exert pressure equally in all directions.
- Gases have much lower density than solids and liquids.
- Gases mix evenly and completely in all proportions without any mechanical aid.


Solid
Liquid
Gas


CHEMICAL CLASSIFICATION
ELEMENTS
COMPOUNDS
MIXTURES

ELEMENTS

- An element is a pure substance.
- It cannot be broken down into two or more simpler substances by physical and chemical means.
- It is made up of only one kind of atom.
- Eg: Copper, Silver, Gold, Iron etc.


Cr
Mg
Fe
Se
Zn
Ca
Na
F
Cu

COMPOUNDS

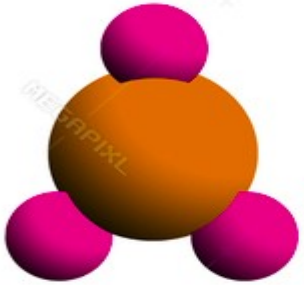
- A compound is a pure substance.
- It is made up of two or more elements combined chemically in fixed proportion.
- Eg: H_2O, NH_3 , CH_4 etc.

Common Chemical Compounds

Water

H_2O

Ammonia



NH_3

Carbon Dioxide

CO_2

Hydrogen Peroxide

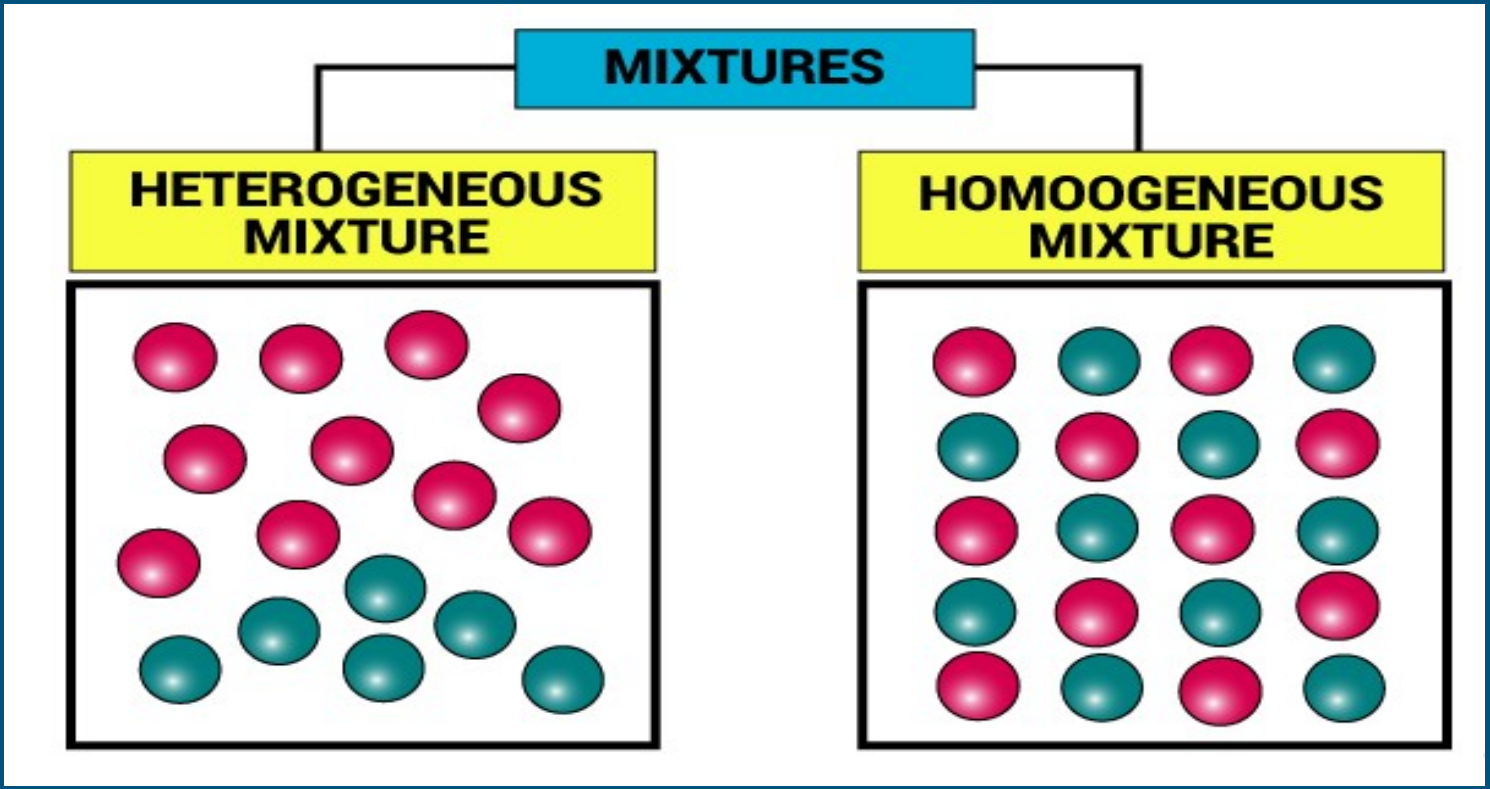
H_2O_2

MIXTURES

A mixture is made up of two or more elements or compounds or both.

Eg: Brass is a mixture of Cu and Zn.

Bronze is a mixture of Cu and Sn.


MIXTURES
HETEROGENEOUS MIXTURE
HOMOOGENEOUS MIXTURE

CLASSIFICATION OF MIXTURES

HOMOGENEOUS MIXTURE

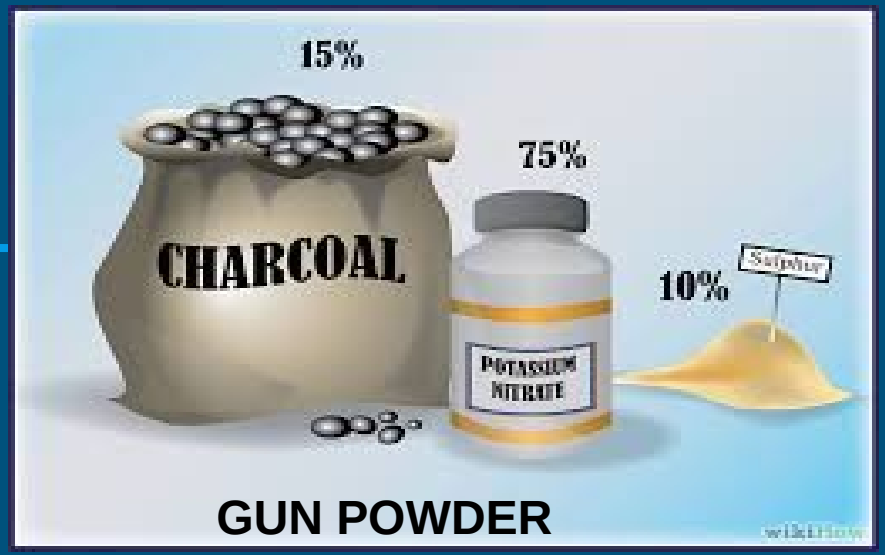
- In a homogeneous mixture, the components completely mix with each other.
- Its composition is uniform throughout the mixture.
- Eg: Ethanol – Water mixture

  Salt solution, sugar solution etc.


Water
Sugar
Sugar solution
Dissolving of Sugar in Water

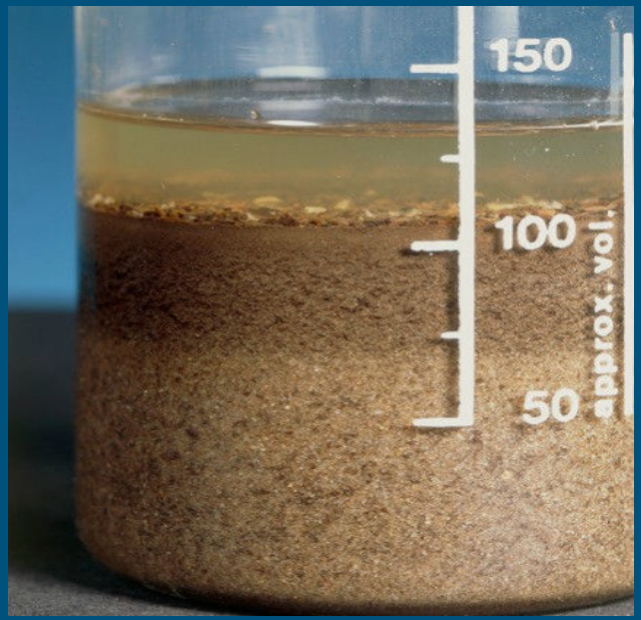
HETEROGENEOUS MIXTURE

- The constituent particles are not uniformly mixed.
- The properties and composition vary throughout the mixture.
- Eg: Gun powder is a mixture of charcoal, sulphur and nitre.
- Muddy water, Chalk powder in water etc.



GUN POWDER

CHALK POWDER IN WATER



SAND IN WATER

PROPERTIES OF MATTER

PHYSICAL PROPERTIES

Properties which can be measured or observed without changing the identity or the composition of the substance.

Eg: Colour, odour, melting point, boiling point, density etc.

CHEMICAL PROPERTIES

- Properties in which a chemical change occur.
- Eg: Chemical reactions of different substances.

MASS

- Mass of a substance is the amount of matter present in it.
- The mass of a substance is constant.
- The SI unit of mass is kg.

WEIGHT

- Weight is the force exerted by gravity on an object.
- Weight may vary from one place to another due to change in gravity.

DENSITY

- **Density of a substance is the amount of mass per unit volume.**

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

- **The SI unit of density is kg/m^3**

TEMPERATURE

- Temperature is the degree of hotness of a body.
- Three common scales are used to measure temperature.
- Degree Celsius (^{0}C)
- Degree Fahrenheit (^{0}F)
- Kelvin (K)

- The Celsius scale is represented between zero degree to hundred degree Celsius.
- Fahrenheit scale is represented between 32^0 to 212^0 F.
- The Fahrenheit scale is related to Celsius scale as

Celsius To Fahrenheit	Fahrenheit To Celsius
$F = \frac{9}{5}C + 32$	$C = \frac{5}{9}(F - 32)$

- The Kelvin scale is related to Celsius scale as follows $K = {}^0C + 273.15$

ACCURACY AND PRECISION

- **ACCURACY** is the agreement of a particular value to the true value of the result.
- **PRECISION** is the closeness of various measurements for the same quantity.
- Eg: If the true value for a result is 2.00 g and a student ‘A’ takes two measurements and reports the result as 1.95 g and 1.93 g.
- These values are precise but are not accurate.

SIGNIFICANT FIGURES

Significant figures are meaningful digits which are known with certainty.

RULES FOR SIGNIFICANT FIGURES

1. All non zero digits are significant.

   Eg: 235 has 3 significant figures.

2. Zero's to the left of the first non zero digit are not significant.

   Eg: 0.002 has only one significant figure.

3. Zero's between non zero digits are significant.

   Eg: 3.02 has 3 significant figures.

4. Zero's to the right of the decimal point are significant.

   Eg: 2.00 has 3 significant figures.

LAWS OF CHEMICAL COMBINATION

LAW OF CONSERVATION OF MASS

The Law States that Matter can neither be created nor destroyed.

Eg: 12 g carbon combines with 32 g oxygen to form 44 g CO_2. Here the total mass of the reactants is equal to the total mass of the products.

Antoine Lavoisier

LAW OF DEFINITE PROPORTION

The same compound always contain the same elements combined in the same fixed proportion by mass.

Joseph Proust

Eg: NaCl may be obtained from sea water. It is also prepared by chemical reactions between NaOH and HCl. These samples on analysis are found to contain Na and Cl in the ratio 23:35.5 by mass.

LAW OF MULTIPLE PROPORTION

when two elements combines to form more than one compound, the different masses of one of the elements which combines with a fixed mass of the other element are in the ratio of simple whole numbers.

John Dalton

Eg: Carbon combines with oxygen to form two different oxides CO and CO_2 under different conditions. In CO_2, 12 g of carbon combines with 32 g of oxygen and in carbon monoxide 12 g carbon combines with 16g of oxygen. In these two cases, the mass of oxygen combining with the fixed mass of carbon are in the ratio 16:32 or 1:2.

GAY LUSSAC'S LAW

when gases combine to form gaseous products, a simple ratio exists between the volumes of the reactants and the products at constant temperature and pressure.

Eg: When H and Cl combine to form HCl, a simple ratio by volume exists between the gases H_2, Cl_2 and HCl at constant temperature and pressure.

John Louis Gay Lussac

AVOGADRO'S LAW

Equal volume of all gases under the same conditions of temperature and pressure contains the same number of molecules.

Avogadro

DALTON'S ATOMIC THEORY

- Matter is made up of small indivisible particles called atoms.
- Atoms of the same element are identical in mass and other properties.
- Atoms of different elements are different in mass and other properties.
- Atoms can neither be created nor be destroyed.
- Since atoms are indivisible, they combine in small whole numbers to form compound atoms called molecules.

ATOMS AND MOLECULES

- **An atom is the smallest possible unit of matter that exhibits all the properties of matter that may or may not have independent existence.**

- **A molecule is the smallest unit of matter which exhibits all the properties of that kind of matter and is capable of independent existence.**

TYPES OF MOLECULES

- A molecule containing only one type of atom is called homonuclear molecule.
- Eg. H_2, O_2, N_2, O_3 etc.
- Molecules containing different types of atoms are called Heteronuclear molecules
- E.g. CO_2, H_2O, $C_6H_{12}O_6$, NH_3 etc.

TYPES OF MOLECULES BASED ON NUMBER OF ATOMS

- Monoatomic molecules contain only one atom.
- Eg. All metals, noble gases like He, Ne, Ar etc.
- Diatomic molecules contain 2 atoms.
- Eg. H_2, O_2, N_2, halogens (F_2, Cl_2, Br_2 and I_2)
- Polyatomic molecules contain more than two atoms.
- Eg. ozone (O_3), Phosphorus (P_4), Sulphur (S_8) etc.

ATOMIC MASS

Atomic mass of an element is defined as a number which expresses how many times the mass of one atom of the element is greater than $\frac{1}{12}^{th}$ mass of a carbon-12 atom.

$$\text{Atomic mass} = \frac{\text{Mass of one atom of the element}}{\frac{1}{12} \text{ x mass of one C-12 atom}}$$

ATOMIC MASS UNIT

One atomic mass unit is defined as mass exactly equal to $\frac{1}{12}^{th}$ the mass of one carbon-12 atom.

amu is also known as unified mass (u).

1 amu = 1.66056×10^{-24}

Eg: Mass of an atom of hydrogen = 1.6736×10^{-24}

The mass of hydrogen atom in terms of amu = $\frac{1.6736 \times 10^{-24} g}{1.66056 \times 10^{-24} g}$

= 1.0078 amu

= 1.008 amu

MOLECULAR MASS

- Molecular mass is the sum of atomic masses of the elements present in a molecule.
- It is obtained by multiplying the atomic mass of each element by the number of its atoms and adding them together.

GRAM ATOMIC MASS OR GRAM ATOM

Atomic mass expressed in grams is called gram atomic mass or simply gram atom.

$$\text{No. of gram atoms} = \frac{\text{Mass in grams}}{\text{Gram atomic mass}}$$

GRAM MOLECULAR MASS OR GRAM MOLE

The molecular mass expressed in grams is called gram molecular mass or simply gram mole.

$$\text{No. of gram mole} = \frac{\text{Mass in grams}}{\text{Gram molecular mass}}$$

FORMULA MASS

- The formula mass of a molecule is the sum of the atomic weights of the atoms in the empirical formula of the compound.
- Eg: The formula mass of NaCl =

  Atomic mass of sodium + Atomic mass of Cl

  = 23 + 35.5 = 58.5

MOLE CONCEPT

- Mole is the unit used for counting very large number of particles like atoms, molecules or ions.
- This is equal to 6.023×10^{23}.
- Mole is represented by the symbol ‘mol’.

MOLAR VOLUME

- The volume occupied by one mole of a gas is called molar volume or gram molar volume.
- One mole of all gases occupies 22.4 L at STP.

PERCENTAGE COMPOSITION

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

Eg: Molar mass of water = 18g

$$\text{Mass percentage of Hydrogen} = \frac{1x2}{18} x100 = 11.1\%$$

$$\text{Mass percentage of oxygen} = \frac{16}{18} x100 = 88.9\%$$

EMPIRICAL FORMULA

- An empirical formula represents the simplest whole number ratio of various atoms present in a compound.
- Eg: Empirical formula of Benzene is CH.

MOLECULAR FORMULA

Molecular formula gives the exact number of different types of atoms present in a molecule of a compound.

Eg: Molecular formula of Benzene is C_6H_6.

RELATION BETWEEN EMPIRICAL FORMULA AND MOLECULAR FORMULA

Molecular formula = n x Empirical Formula

$$n = \frac{\text{Molecular Mass}}{\text{Empirical Formula Mass}}$$

STOICHIOMETRY

- The word Stoichiometry is defined from two Greek words.
- ‘stoichion’ meaning element and ‘metron’ meaning measure.
- Stoichiometry deals with the calculation of masses of the reactants and the products involved in a chemical reaction.

EXAMPLE

$$CH_4\ (g) + 2O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (g)$$

- Here 1 mol of CH_4 reacts with 2 moles of O_2 to give one mole of CO_2 and 2 moles of H_2O

OR

- 16g of CH_4 reacts with 64g of O_2 to give 44g of CO_2 and 36 g of H_2O

OR

- 22.4L of CH_4 reacts with 44.8L of oxygen to give 22.4L of CO_2 and 44.8L of H_2O.

REACTIONS IN SOLUTIONS

MASS PERCENTAGE

Mass percentage of a component in a given solution is the mass of the compound per 100g of the solution.

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

MOLE FRACTION

It is the ratio of number of moles of one component to the total number of moles of the solution.

$$\text{Mole fraction of a component} = \frac{\text{No. of moles of the component}}{\text{Total No. of moles of all the components}}$$

If a substance 'A' dissolves in a substance 'B' and their number of moles are n_A and n_B respectively. Then the mole fractions of A and B are given as

$$\text{Mole fraction of A, } X_A = \frac{n_A}{n_A + n_B}$$

$$\text{Mole fraction of B, } X_B = \frac{n_B}{n_A + n_B}$$

$$X_A + X_B = \frac{n_A}{n_A + n_B} + \frac{n_B}{n_A + n_B} = 1$$

MOLARITY

- It is defined as the number of moles of the solute in one litre of the solution.
- It is denoted by 'M'

$$\text{Molarity (M)} = \frac{\text{No. of moles of the solute}}{\text{Volume of the solution in litre}}$$

$$\text{Molarity (M)} = \frac{\text{Mass per litre}}{\text{Molar Mass}}$$

MOLARITY

- **The number of moles of solute present in 1 kg of solvent.**
- **It is denoted by ‘m’.**

$$\text{Molality (m)} = \frac{\text{No. of moles of solute}}{\text{Mass of solvent in kg}}$$

NORMALITY

- It is defined as the number of gram equivalents of the solute dissolved in one litre of the solution.
- It is denoted by 'N'

$$\text{Normality (N)} = \frac{\text{No. of gram equivalents of solute}}{\text{Volume of the solution in litre}}$$

$$\text{Normality (N)} = \frac{\text{Mass per litre}}{\text{Equivalent Mass}}$$

LIMITING REACTANT OR LIMITING REAGENT

The reagent which is completely consumed in a chemical reaction is called limiting reactant or limiting reagent.


Thank You