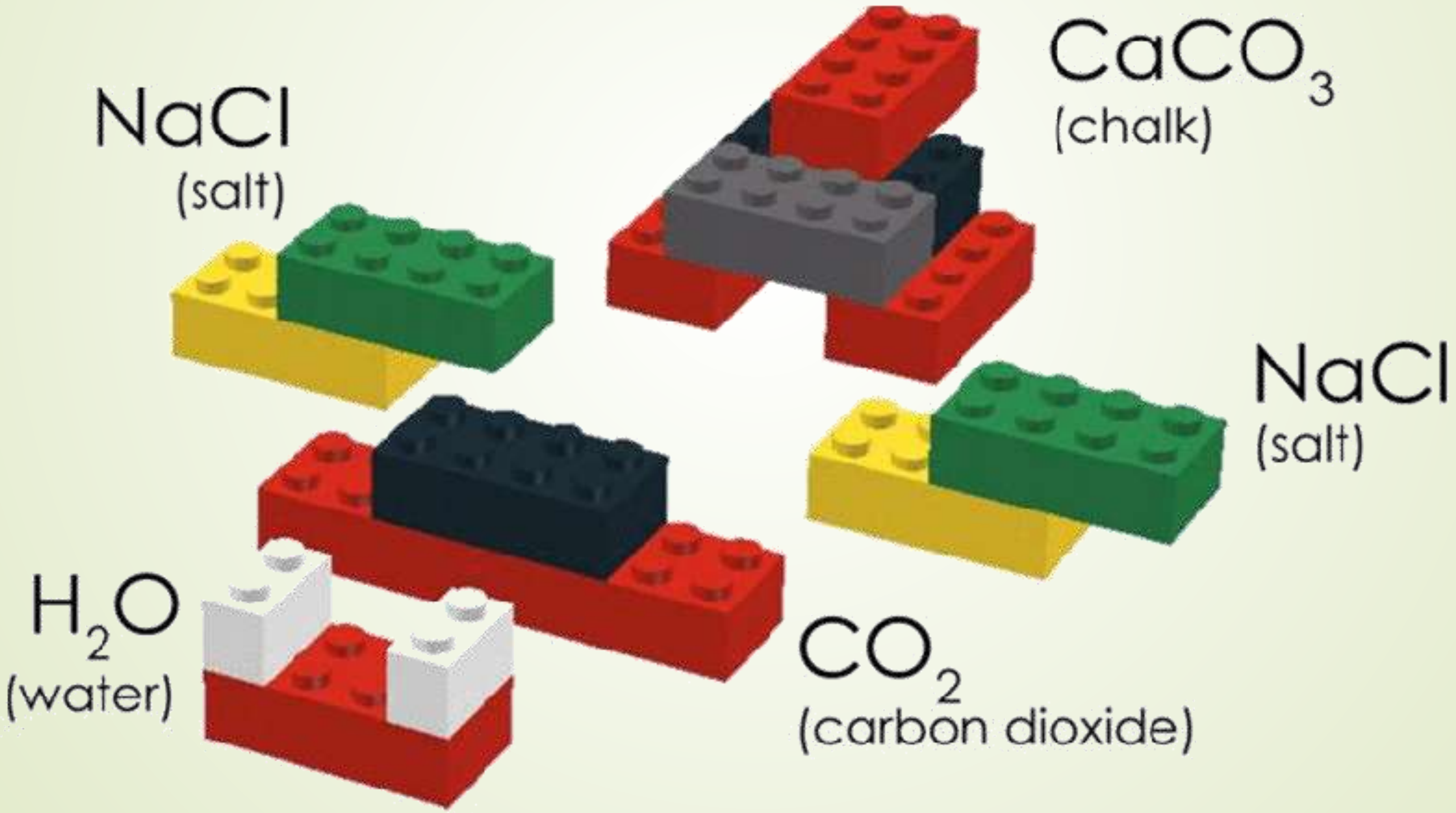
Some basics

1. Some basic concepts of chemistry

- Understand and appreciate the role of chemistry in different spheres of life
- Explain the characteristics of three states of matter
- Classify different substances into elements, compounds and mixtures
- Define SI base units
- Use scientific notations
- Convert physical quantities from one system of units to another
- Explain various laws of chemical combination
- Appreciate significance of atomic mass, average atomic mass, molecular mass and formula mass
- Describe the terms – mole and molar mass
- Determine empirical formula and molecular formula for a compound from the given experimental data
- Perform the stoichiometric calculations

What is Chemistry -

- Chemistry is the science of atoms, molecules and their transformations.
- Chemistry deals with the composition, structure, properties & interaction of matter.

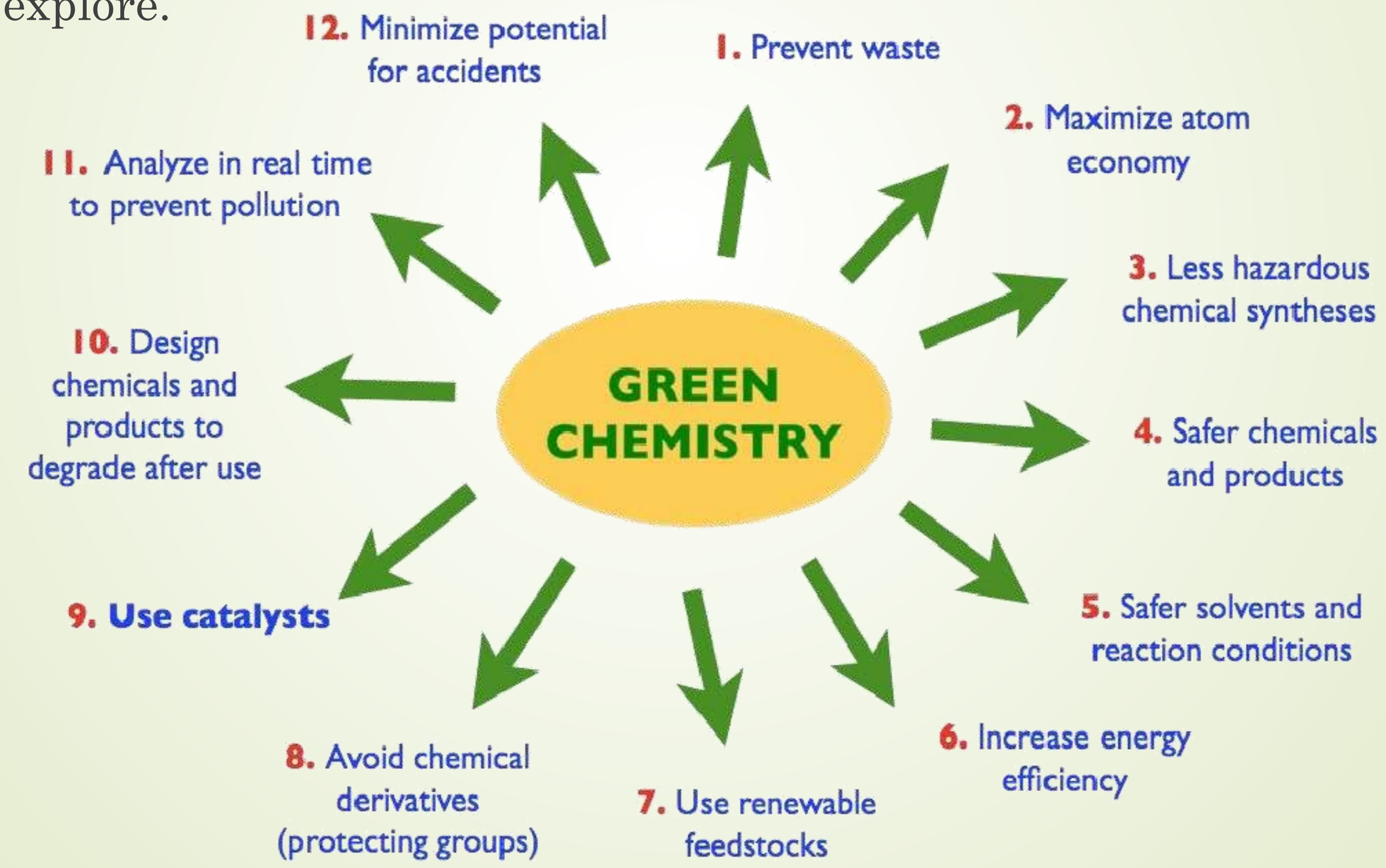


Why study Chemistry -

Why study Chemistry -

Advantages -

- With a better understanding of chemical principles it has now become possible to design and synthesize new materials having specific magnetic, electric and optical properties.
- Safer alternatives to environmentally hazardous refrigerants like CFCs (chlorofluorocarbons), responsible for ozone depletion in the stratosphere, have been successfully synthesized.
- Still a lot to explore.



Matter -

Anything which has mass and occupies space is called **matter**.

Physical states of matter(internal order) –

- In **solids**, these particles are held very close to each other in an orderly fashion and there is not much freedom of movement.
- In **liquids**, the particles are close to each other but they can move around.
- In **gases**, the particles are far apart as compared to those present in solid or liquid states and their movement is easy and fast.

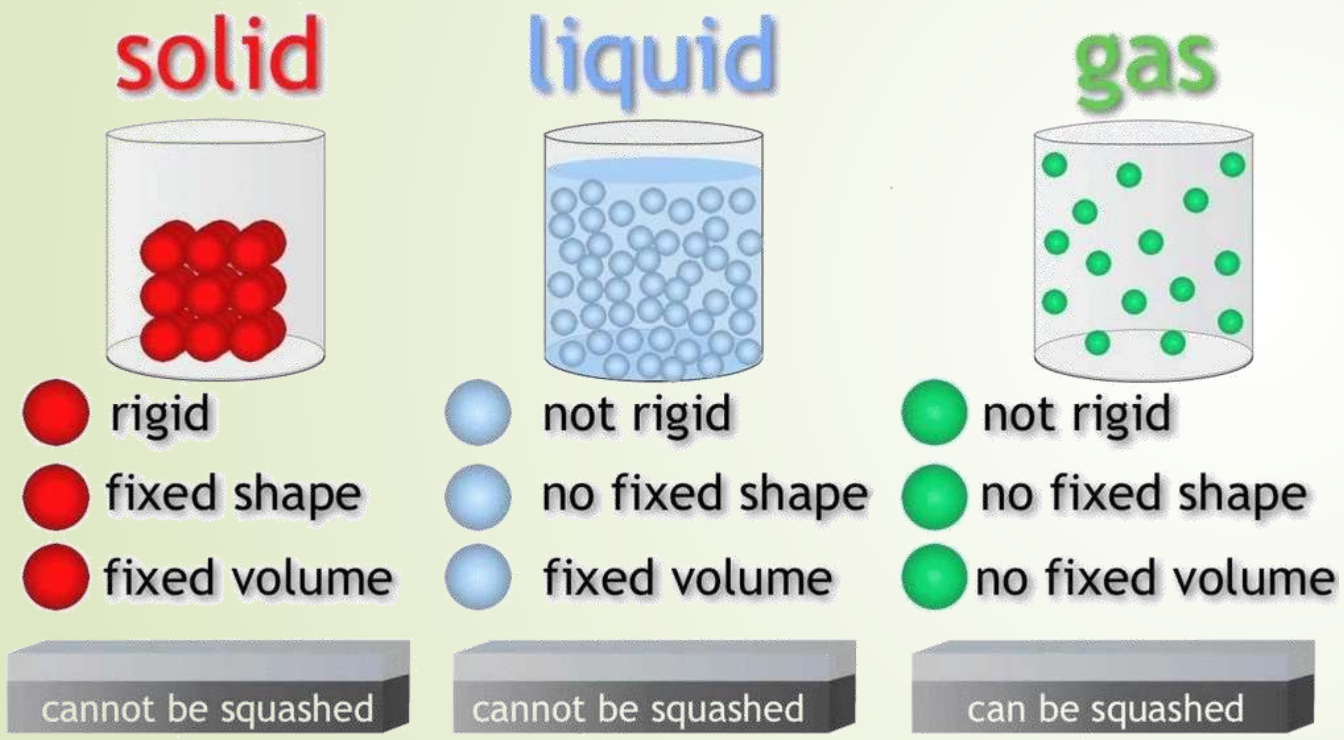
Hence –

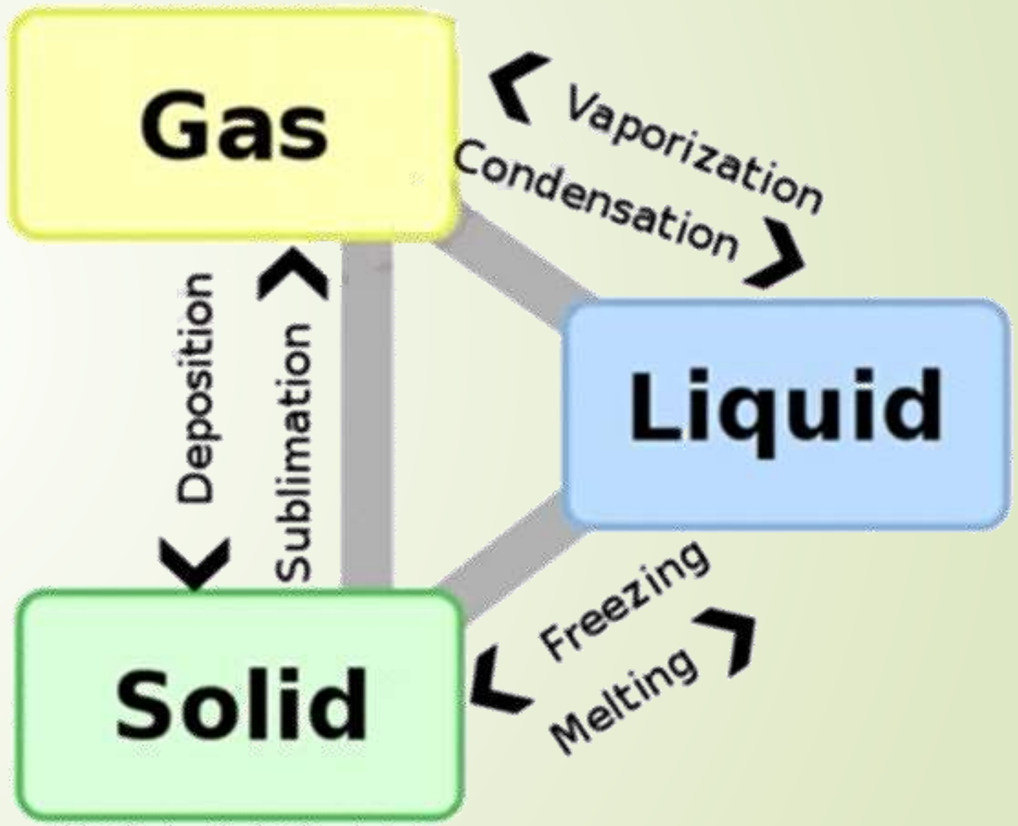
(i) Solids have definite volume and definite shape.

(ii) Liquids have definite volume but not the definite shape. They take the shape of the container in which they are placed.

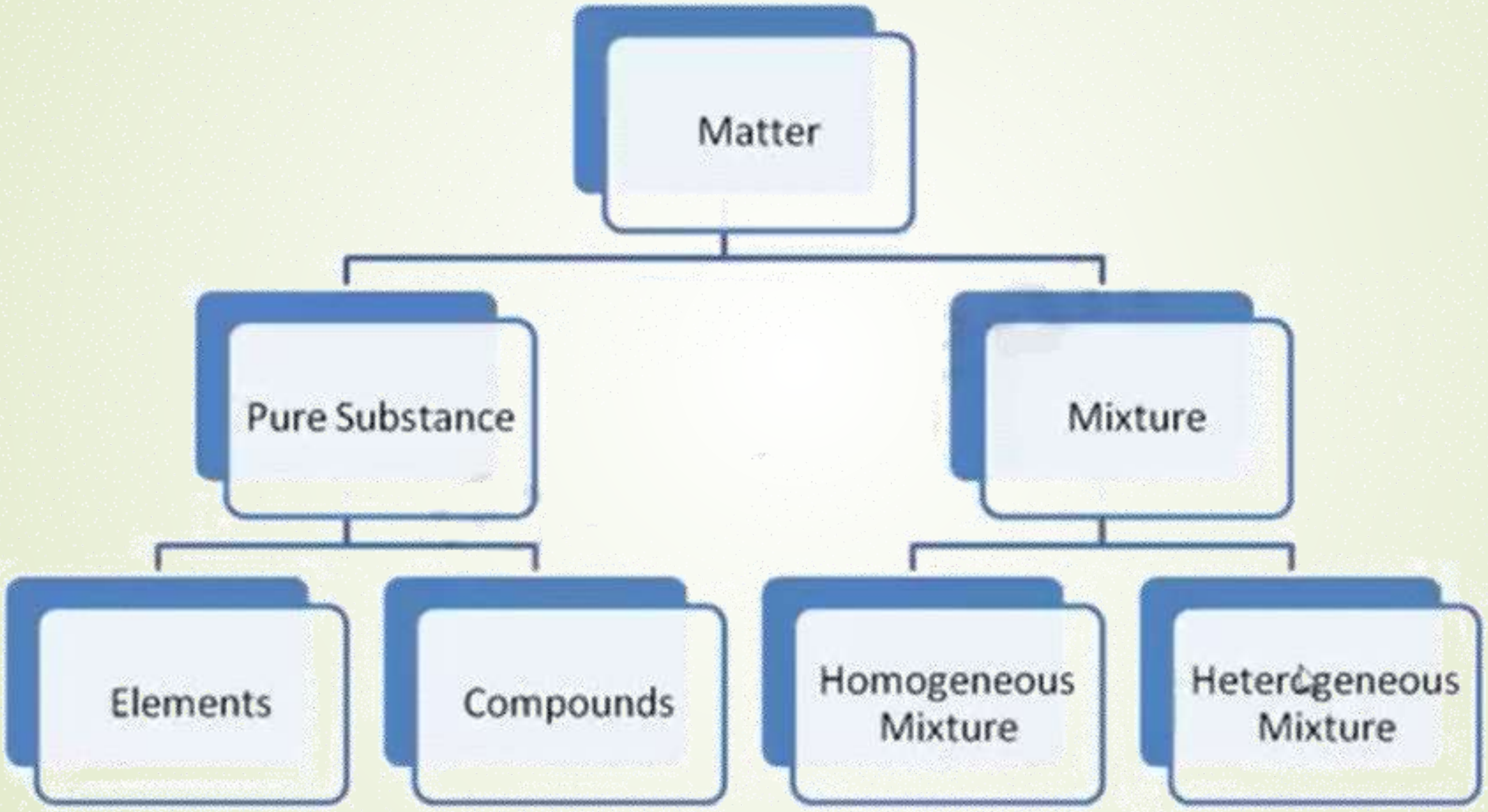
(iii) Gases have neither definite volume nor definite shape. They completely occupy the container in which they are placed.

Interconversion of States -





Matter at macroscopic level -



Pure Substances -

- They have fixed composition.
- The constituents of pure substances cannot be separated by simple physical methods.
- Pure substances can be further classified into **elements** and **compounds**.

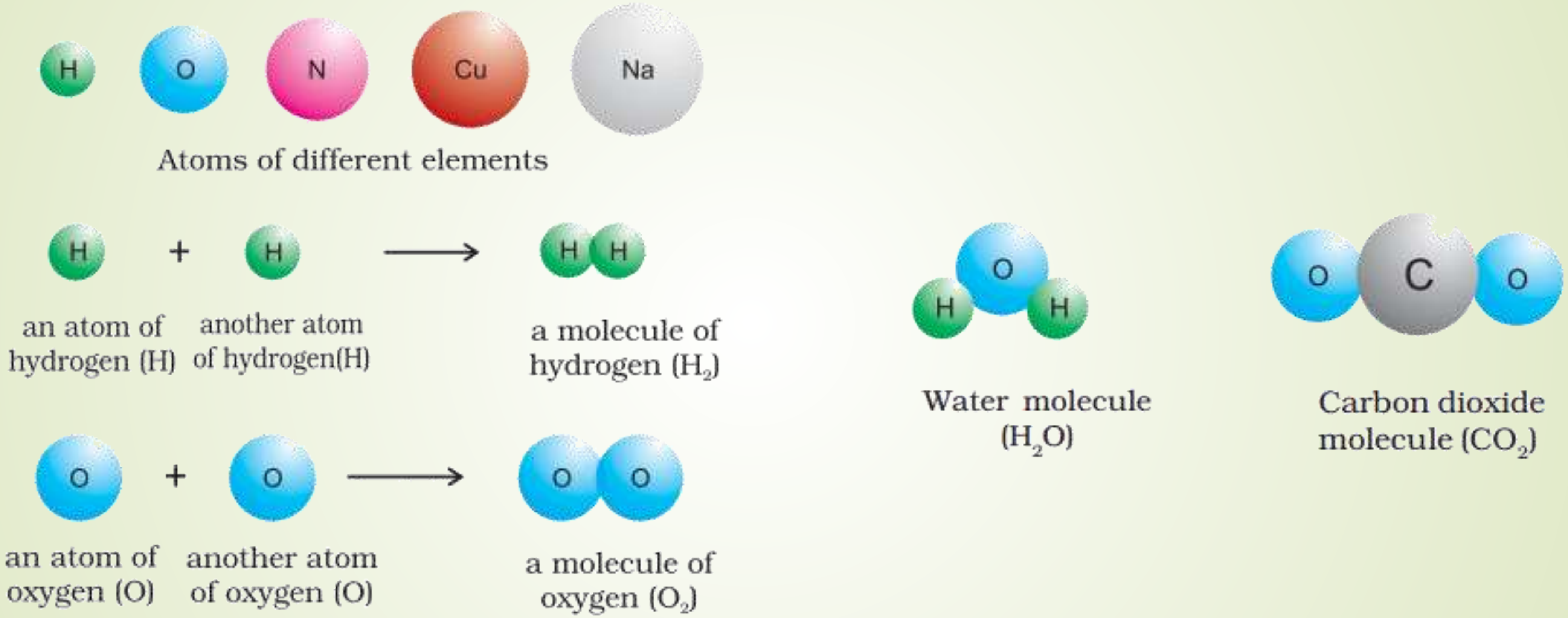
Element-

-Consists of only one type of particles.

-These particles may be **atoms** or **molecules.**

Compound-

-When two or more atoms of different elements combine, the molecule of a compound is obtained.


H
O
N
Cu
Na
Atoms of different elements
an atom of hydrogen (H)
another atom of hydrogen(H)
a molecule of hydrogen (H_2)
an atom of oxygen (O)
another atom of oxygen (O)
a molecule of oxygen (O_2)
Water molecule (H_2O)
Carbon dioxide molecule (CO_2)
C

Compounds: Cont. -

- The atoms of different elements are present in a compound in a fixed and definite ratio and this ratio is characteristic of a particular compound.
- The constituents of a compound cannot be separated into simpler substances by physical methods.
- They **can be separated by chemical methods**.
- The properties of a compound are different from those of its constituent elements.

E.g. hydrogen burns with a pop sound and oxygen is a supporter of combustion, but water is used as a fire extinguisher.

Mixture -

- A mixture contains 2 or more substances present in any ratio.

  E.g. Air, pulse & stones

- **Imp** – Components of a Mixture can be separated using

  physical methods such as hand picking, filtration, distillation etc.

- A mixture can be homogenous or heterogeneous.

Homogenous mixture –

- The components completely mix with each other and its composition is uniform throughout. E.g. Air, sugar

Heterogeneous mixture -

- Composition is not uniform throughout and sometimes the different components can be observed. E.g. gems, chalk water, smog

Properties of matter -

- Physical properties
- Chemical properties

Physical properties -

- Those properties which can be measured or observed without changing the identity or the composition of the substance.
- Some examples of physical properties are color, odour, melting point, boiling point, density etc.

Chemical properties –

- The measurement or observation of chemical properties require a chemical change to occur.
- Some examples - acidity or basicity, combustibility etc.

Quantitative property -

- Many properties of matter such as length, area, volume, etc., are quantitative in nature.
- Any quantitative observation or measurement is represented by a number followed by units in which it is measured.

For example - Length of a room can be represented as 6 m; here 6 is the number and 'm' denotes *metre* – the unit in which the length is measured.

Two different systems of measurement –

- **English System**
- **Metric System ;** The metric system which originated in France in late eighteenth century, was more convenient as it was based on the decimal system.

The International System of Units(SI) -

- The International System of Units (in French Le Systeme International d'Unités – abbreviated as SI) was established by the 11th General Conference on Weights and Measures.
- The SI system has seven base units and they are listed in Table. These units pertain to the seven fundamental scientific quantities.
- The other physical quantities such as speed, volume, density etc. can be derived from these quantities.

Table 1.1 Base Physical Quantities and their Units

Base Physical Quantity	Symbol for Quantity	Name of SI Unit	Symbol for SI Unit
Length	l	metre	m
Mass	m	kilogram	kg
Time	t	second	s
Electric current	I	ampere	A
Thermodynamic temperature	T	kelvin	K
Amount of substance	n	mole	mol
Luminous intensity	I_v	candela	cd

Definitions of base quantities -

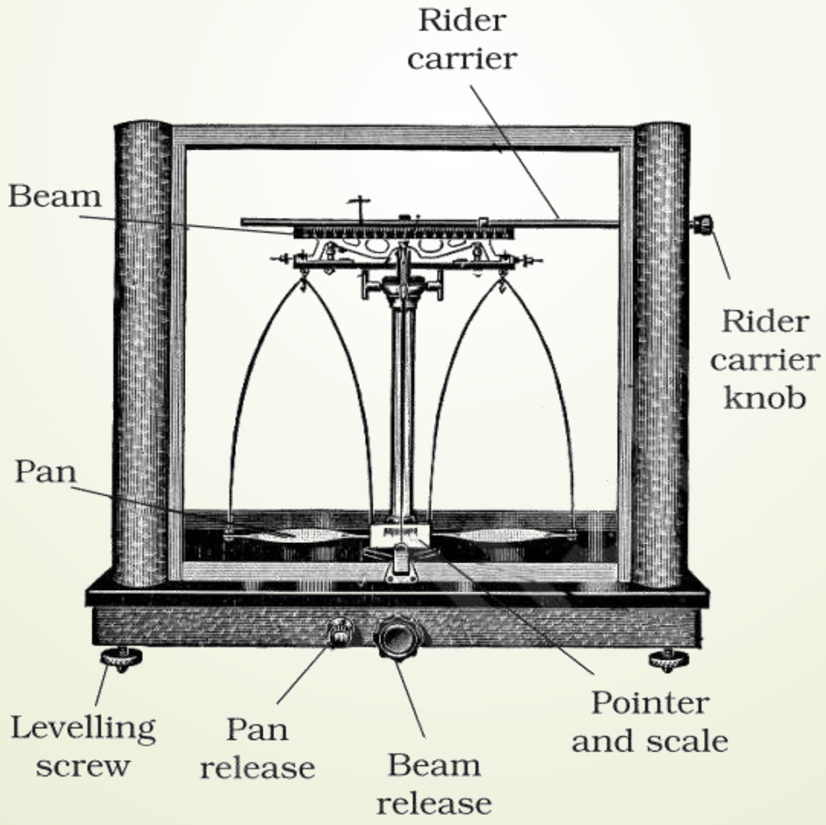
- Length → Metre (m): The metre is the length of the path travelled by light in vacuum during a time interval of 1/299,792,458 of a second. (1983)
- Mass → kilogram (kg) : The kilogram is equal to the mass of the international prototype of the kilogram (a platinum-iridium alloy cylinder) kept international Bureau of Weights and Measures, at Sevres, near Paris, France. (1889)
- Time → Second(s) : The second is the duration of 9,192,631,770 periods of the radiation corresponding to the transition between the two hyperfine levels of the ground state of the cesium-133 atom(1967)
- Electric Current → Ampere (A) : The ampere is that constant current which, if maintained in two straight parallel conductors of infinite length, of negligible circular cross-section, and placed 1 metre apart in vacuum would produce between these conductors a force equal to $2x10^{-7}$ newton per metre of length. (1948)
- Temperature → Kelvin(K) : The kelvin, is the fraction 1/273.16 of the thermodynamic temperature of the triple point of water. (1967)
- Amount of substance → Mole(mol) : The mole is the amount of substance of a system, which contain as many elementary entities as there are atoms in 0.01 kilogram of carbon - 12. (1971)
- Luminous intensity → Candela(cd) : The candela is the luminous intensity, in a given direction, of a source that emits monochromatic radiation frequency 540 $x10^{12}$ hertz and that has a radiant intensity in that direction of 1/683 watt per steradian. (1979)

Prefixes used in the SI System -

Multiple	Prefix	Symbol
10^{-24}	yocto	y
10^{-21}	zepto	z
10^{-18}	atto	a
10^{-15}	femto	f
10^{-12}	pico	p
10^{-9}	nano	n
10^{-6}	micro	µ
10^{-3}	milli	m
10^{-2}	centi	c
10^{-1}	deci	d
10	deca	da
10^{2}	hecto	h
10^{3}	kilo	k
10^{6}	mega	M
10^{9}	giga	G
10^{12}	tera	T
10^{15}	peta	P
10^{18}	exa	E
10^{21}	zeta	Z
10^{24}	yotta	Y

Mass and Weight

- Mass of a substance is the amount of matter present in it.
- Weight is the force exerted by gravity on an object.
- The mass of a substance is constant whereas its weight may vary from one place to another due to change in gravity.
- The mass of a substance can be determined very accurately in the laboratory by using an analytical balance.

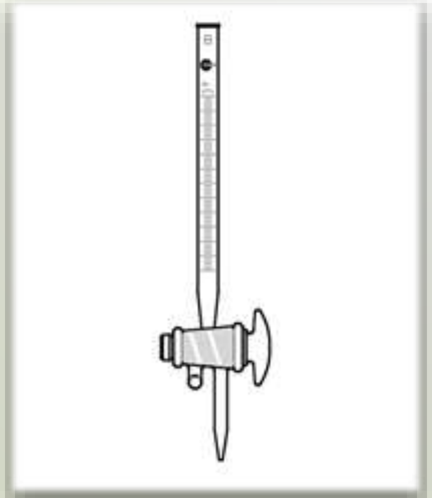


Volume

- Volume has the units of $(length)^3$. So in SI system, volume has units of m^3.
- A common unit, litre (L) which is not an SI unit, is used for measurement of volume of liquids.

  $$1\ L = 1000\ mL\ ,\ 1000\ cm^3 = 1\ dm^3$$

- In the laboratory, volume of liquids or solutions can be measured by graduated cylinder, burette, pipette etc. A volumetric flask is used to prepare a known volume of a solution.



Density

Density of a substance is its amount of mass per unit volume. So SI units of density can be obtained as follows:

$$D = M/V$$

SI unit of density = SI unit of mass / SI unit of volume

$= kg/m^3$ or $kg\ m^{-3}$

This unit is quite large and a chemist often expresses density in $g\ cm^{-3}$.

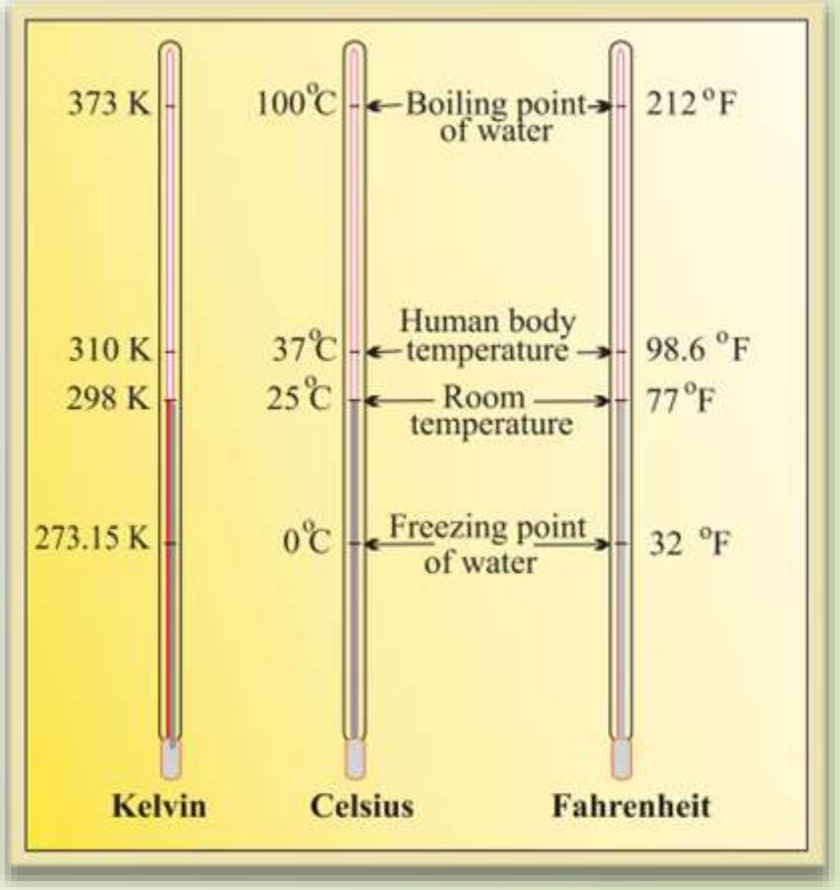
Temperature

- There are three common scales to measure temperature — °C (degree celsius), °F (degree fahrenheit) and K (kelvin). Here, K is the SI unit.
- The fahrenheit scale is represented between 32° to 212°.
  The temperatures on two scales are related to each other by the following relationship:

$$F = 9/5\ C + 32$$

$$K = °C + 273.15$$

- In Kelvin scale, negative temperature is not possible.



Uncertainty In Measurement

- As chemistry is the study of atoms and molecules which have extremely low masses and are present in extremely large numbers, a chemist has to deal with numbers as large as 602,200,000,000,000,000,000,000 for the molecules of 2 g of hydrogen gas or as small as 0.00000000000000000000000166 g mass of a H atom.
- This problem is solved by using scientific notation for such numbers, *i.e.*, exponential notation in which any number can be represented in the form N x 10^n where n is an exponent having positive or negative values and N varies between 1.000... and 9.999…

Precision & Accuracy

Every experimental measurement has some amount of uncertainty associated with it. However, one would always like the results to be precise and accurate.

Precision refers to the closeness of various measurements for the same quantity.

Accuracy is the agreement of a particular value to the true value of the result.

Student	Exp1 output	Exp2 output	Mean	Is Precise	Is Accurate
Ramesh	2.80	3	2.9	No	No
Suresh	2.91	2.90	2.905	Yes	No
Rohit	3.01	2.99	3	Yes	Yes
Rakesh	2.90	3.10	3	No	Yes

Significant figures

- The uncertainty in the experimental or the calculated values is indicated by mentioning the number of significant figures.
- Significant figures are meaningful digits which are known with certainty.

Rules for significant figures:

- All non-zero digits are significant.
- Zeros preceding to first non-zero digit are not significant.
- Zeros between two non-zero digits are significant.
- Zeros at the end or right of a number are significant provided they are on the right side of the decimal point.
- Exact numbers have infinite significant figures.

Addition and Subtraction of Significant Figures

The result cannot have more digits to the right of the decimal point than either of the original numbers.

Multiplication and Division of Significant Figures

In these operations, the result must be reported with no more significant figures as are there in the measurement with the few significant figures.

Dimensional Analysis

- Often while calculating, there is a need to convert units from one system to other.
- The method used to accomplish this is called **factor label method** or **unit factor method** or **dimensional analysis**. E.g.

Laws Of Chemical Combinations

The combination of elements to form compounds is governed by the following five basic laws.

i. Law of Conservation of Mass
ii. Law of Definite Proportions
iii. Law of Multiple Proportions
iv. Gay Lussac's Law of Gaseous Volumes
v. Avogadro Law

Law of Conservation of Mass

- It states that matter can neither be created nor destroyed.
- This law was put forth by Antoine Lavoisier in 1789.
- He performed careful experimental studies for combustion reactions for reaching to the above conclusion.

Law of Definite Proportions

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

This law was given by, a French chemist, Joseph Proust.

Proust worked with two samples of cupric carbonate:

- one of which was of natural origin

- other was synthetic one.

He found that the composition of elements present in it was same for both the samples.

	% of copper	% of oxygen	% of carbon
Natural Sample	51.35	9.74	38.91
Synthetic Sample	51.35	9.74	38.91

Thus, irrespective of the source, a given compound always contains same elements in the same proportion.

Law of Multiple Proportions

- According to this law, if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers.
- This law was proposed by Dalton in 1803.
- For example, hydrogen combines with oxygen to form two compounds, namely, water and hydrogen peroxide.

Hydrogen + Oxygen → Water
2g 16g 18g

Hydrogen + Oxygen → Hydrogen Peroxide
2g 32g 34g

- Here, the masses of oxygen (i.e. 16 g and 32 g) which combine with a fixed mass of hydrogen (2g) bear a simple ratio, i.e. 16:32 or 1: 2.

Gay Lussac's Law of Gaseous Volumes

- He observed that when gases combine or are produced in a chemical reaction they do so in a simple ratio by volume provided all gases are at same temperature and pressure.

- This law was given by Gay Lussac in 1808.

- Thus, 100 mL of hydrogen combine with 50 mL of oxygen to give 100 mL of water vapour.

$$\begin{array}{ccccc} \text{Hydrogen} & + & \text{Oxygen} & \rightarrow & \text{Water} \\ 100\text{ mL} & & 50\text{ mL} & & 100\text{ mL} \end{array}$$

- The volumes of hydrogen and oxygen which combine together (i.e. 100 mL and 50 mL) bear a simple ratio of 2:1.

Avogadro Law

- In 1811, Avogadro proposed that equal volumes of gases at the same temperature and pressure should contain equal number of molecules.
- Avogadro made a distinction between **atoms** and **molecules** .

Dalton’s atomic theory

In 1808, Dalton published ‘A New System of Chemical Philosophy’ in which he proposed the following :

i. Matter consists of indivisible atoms.

ii. All the atoms of a given element have identical properties including identical mass. Atoms of different elements differ in mass.

iii. Compounds are formed when atoms of different elements combine in a fixed ratio.

iv. Chemical reactions involve reorganization of atoms. These are neither created nor destroyed in a chemical reaction.

Dalton’s theory could explain the laws of chemical combination.

Atomic mass: History

- The first among scientists to determine atomic mass was John Dalton in 1803. It was called Atomic Weight.
- Atomic weight was originally defined relative to that of lightest element, hydrogen, taken as 1.
- Both chemists & physicists started using oxygen as basis. Atomic no- 16 (Whole number)
- It was later found that natural Oxygen contains other isotopes to oxygen too.
- Chemists picked naturally occurring Oxygen, which is mixture of O-16, O-17 & O-18. (Atomic number – 16.008).
- Physicists picked up Carbon-12 as base based on mass spectrometry & mass of one carbon-12 atom is a whole number.
- This led to 2 different tables of atomic mass.
- Present system of atomic masses is based on carbon-12 as the standard and has been agreed upon in 1961 as only the mass of one carbom-12 atom is a whole number.
- In this system, 12C is assigned a mass exactly 12 atomic mass unit(amu) and masses of all other atoms are given relative to this standard.

Average Atomic mass

Many naturally occurring elements exist as more than one isotope.

When we take into account the existence of these isotopes and their relative abundance (per cent occurrence), the average atomic mass of that element can be computed.

Isotopes	Relative abundance	AMU
C 12	98.892	12
C 13	1.108	13.0035
C 14	$2 * 10^{-10}$	14.0037

Molecular Mass

Molecular mass is the sum of atomic masses of the elements present in a molecule.

E.g. Molecular mass of methane, (CH_4)

$= (12.011 \text{ u}) + 4\ (1.008 \text{ u})$

$= 16.043 \text{ u}$

Formula Mass

- Some substances such as sodium chloride do not contain discrete molecules as their constituent units.
- In such compounds, positive (sodium) and negative (chloride) entities are arranged in a three-dimensional structure.
- The formula such as NaCl is used to calculate the **formula mass** instead of molecular mass as in the solid state sodium chloride does not exist as a single entity.

Formula mass of sodium chloride

= atomic mass of sodium + atomic mass of chlorine

= 23.0 u + 35.5 u = 58.5 u

Mole concept and molar masses

- Atoms and molecules are extremely small in size and their numbers in even a small amount of any substance is really very large.
- To handle such large numbers, a unit of similar magnitude is required.
- Just as we denote one dozen for 12 items, score for 20 items, gross for 144 items, we use the idea of mole to count entities at the microscopic level (i.e. atoms/molecules/particles, electrons, ions, etc).

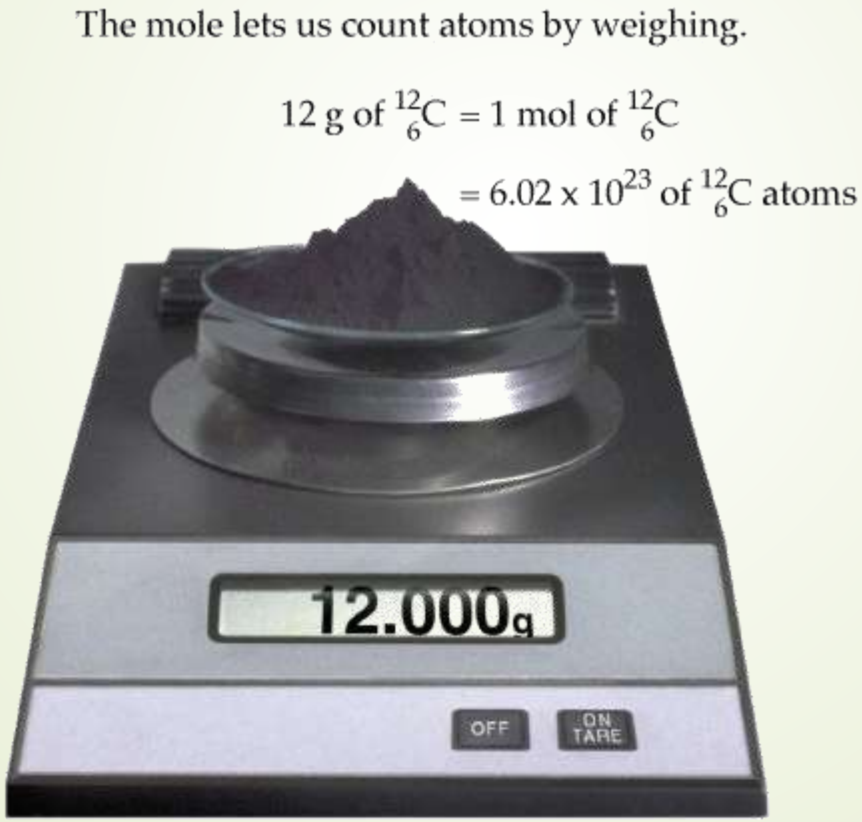
One mole is the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 g (or 0.012 kg) of the 12C isotope.

- It may be emphasized that the mole of a substance always contain the same number of entities, no matter what the substance may be.

1 Mole C-12

The mass of a carbon–12 atom was determined by a mass spectrometer and found to be equal to 1.992648 x 10^{-23} g.

Knowing that one mole of carbon weighs 12 g, the number of atoms in it is equal to :



1 mole of :


Sucrose
Copper (II) sulfate pentahydrate
Copper
Mercury (II) oxide
Sulfur
Sodium chloride

Molar mass

The mass of one mole of a substance in grams is called its Molar mass.

- The molar mass in grams is numerically equal to atomic/molecular/formula mass in u.
- Molar mass of water = 18.02 g mol^{-1}
- Molar mass of sodium chloride = 58.5 g mol^{-1}

Percentage composition

Suppose an unknown or new compound is given to you, the first question you would ask is:

What is its formula or what are its constituents and in what ratio are they present in the given compound?

For known compounds also, such information provides a check whether the given sample contains the same percentage of elements as is present in a pure sample. One can check the purity of a given sample by analysing this data.

Let us understand it by taking the example of water (H_2O).

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

Empirical Formula & Molecular Formula

Empirical formula is the simplest whole number ratio of various atoms present in a compound.

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

If the mass per cent of various elements present in a compound is known, its empirical formula can be determined.

Molecular formula can further be obtained if the molar mass is known.

Stoichiometry

The word ‘stoichiometry’ is derived from two Greek words - *stoicheion* (meaning *element*) and *metron* (meaning *measure*).

Stoichiometry deals with the calculation of masses (sometimes volumes also) of the reactants and the products involved in a chemical reaction.

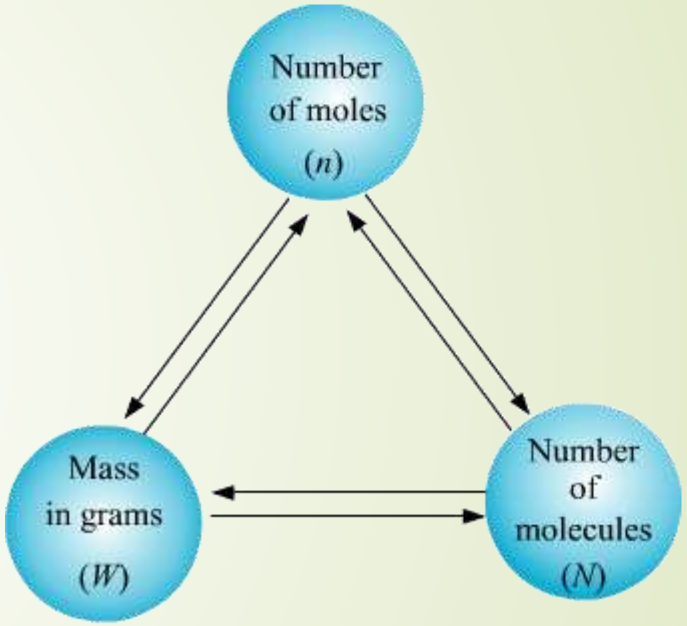
$$Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$$

$$Mg(s) + O_2(g) \rightarrow MgO(s)$$

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

- **One mole** of $CH_4(g)$ reacts with **two moles** of $O_2(g)$ to give **one mole** of $CO_2(g)$ and **two moles** of $H_2O(g)$.
- One molecule of $CH_4(g)$ reacts with 2 molecules of $O_2(g)$ to give one molecule of $CO_2(g)$ and 2 molecules of $H_2O(g)$.
- 16 g of CH_4 (g) reacts with 232 g of O_2 (g) to give 44 g of CO_2 (g) and 218 g of H_2O (g).
- 22.7 L of $CH_4(g)$ reacts with 45.4 L of O_2 (g) to give 22.7 L of CO_2 (g) and 45.4 L of $H_2O(g)$.

E.g.



Limiting Reagent

Many a time, the reactions are carried out when the reactants are not present in the amounts as required by a balanced chemical reaction.

In such situations, one reactant is in excess over the other.

The reactant which is present in the lesser amount gets consumed after sometime and after that no further reaction takes place whatever be the amount of the other reactant present.

The reactant which gets consumed, limits the amount of product formed and is, therefore, called the **Limiting reagent**.

Reactions in Solutions

A majority of reactions in the laboratories are carried out in solutions.

Therefore, it is important to understand as how the amount of substance is expressed when it is present in the form of a solution.

The concentration of a solution can be expressed by the following ways.

i. Mass per cent or weight per cent (w/w %)

ii. Mole fraction

iii. Molarity

iv. Molality

i. Mass percentage (w/w)

The mass percentage of a component of a solution is defined as:

$$\text{Mass \% of a component} = \frac{\text{Mass of the component in the solution}}{\text{Total mass of the solution}} \times 100$$

For e.g. - If a solution is described by 10% glucose in water by mass, it means that 10 g of glucose is dissolved in 90 g of water resulting in a 100 g solution.

Commercial bleaching solution contains 3.62 mass percentage of sodium hypochlorite in water.

ii. Mole fraction

Commonly used symbol for mole fraction is 'x' and subscript used on the right hand side of x denotes the component. It is defined as

$$\text{Mole fraction of a component} = \frac{\text{Number of moles of the component}}{\text{Total number of moles of all the components}}$$

For e.g. in a binary mixture, if the number of moles of A and B are n_A and n_B respectively, the mole fraction of A will be

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

Note: $x_1 + x_2 + \ldots\ldots\ldots\ldots\ldots\ldots + x_i = 1$

iii. Molarity

Molarity (*M*) is defined as number of moles of solute dissolved in one litre (or one cubic decimeter) of solution

$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Volume of solution in litre}}$$

For e.g. 0.25 mol L^{-1} (or 0.25 M) solution of NaOH means that 0.25 mol of NaOH has been dissolved in one litre (or one cubic decimetre)

iv. Molality

Molality (m) is defined as the number of moles of the solute per kilogram (kg) of the solvent and is expressed as:

$$\text{Molality (m)} = \frac{\text{Moles of solute}}{\text{Mass of solvent in kg}}$$

For e.g. 1.00 mol kg^{-1} (or 1.00 m) solution of KCl means that 1 mol (74.5 g) of KCl is dissolved in 1 kg of water.

End ☺