Welcome to

Aakash + BYJU'S LIVE

Some Basic Concepts of Chemistry

Classification of matter

Anything that **has mass** and **occupies space.**

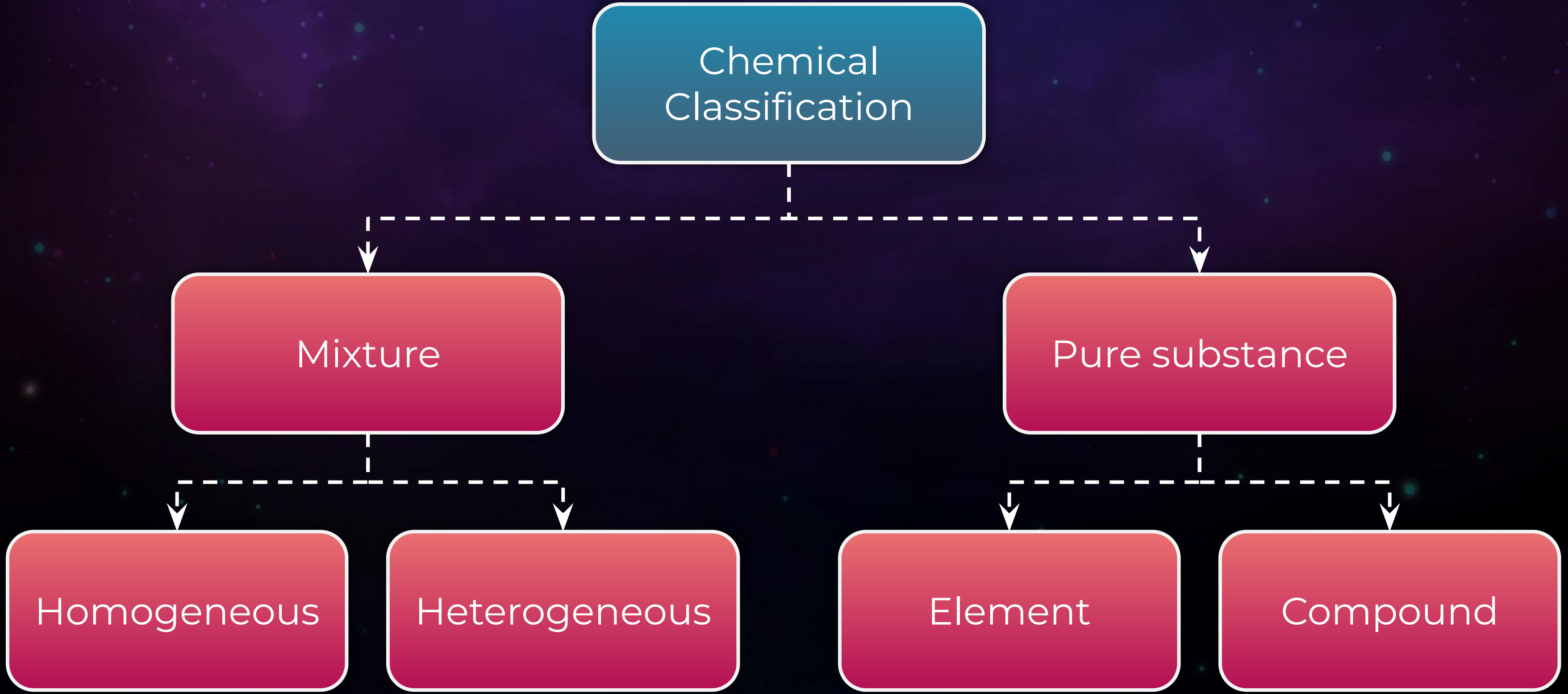
Matter

- Physical classification
  - **Solid:** Definite shape & volume
  - **Liquid:** No specific shape but definite volume
  - **Gas:** No fixed shape & volume
- Chemical classification

Classification of matter

- Matter
  - Physical classification
  - Chemical classification
    - **Mixture:** Variable composition
    - **Pure substance:** Fixed composition

Classification of matter



Properties of matter

Properties of Matter

Physical properties

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

Colour, melting point, boiling point, density etc.

Chemical properties

Properties which can be measured only by a **chemical reaction**.

Acidity or basicity, combustibility, etc.

Measurement of Physical Properties

Any **quantitative observation** or measurement such as length, area, volume, etc. is represented by a **number followed by units** in which it is measured.

Number

↓

Length of a room = **6 m**

↑

Unit in which length is measured

Unit

A unit is the **standard of reference** chosen to measure or express any physical quantity.

S.I. system has **seven base units** pertaining to the seven fundamental scientific quantities.

The International System of Units (S.I.)

Base Physical Quantity	Symbol for Quantity	Name of S.I. unit	Symbol for S.I. unit
Length	l	Meter	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

The International System of Units (S.I.)

Derived unit

The physical quantities unit which are derived using the S.I. base units in combinations.

Example

Unit of area (m^2),
unit of density ($kg\ m^{-3}$)

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

Multiple	Prefix	Symbol
10^{-9}	nano	n
10^{-6}	micro	μ
10^{-3}	milli	m
10^{-2}	centi	c
10^{-1}	deci	d

Prefixes Used in the SI System

Multiple	Prefix	Symbol
10	deca	da
10^2	hecto	h
10^3	kilo	k
10^6	mega	M
10^9	giga	G

Multiple	Prefix	Symbol
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, while weight is the **force exerted by gravity** on an object.

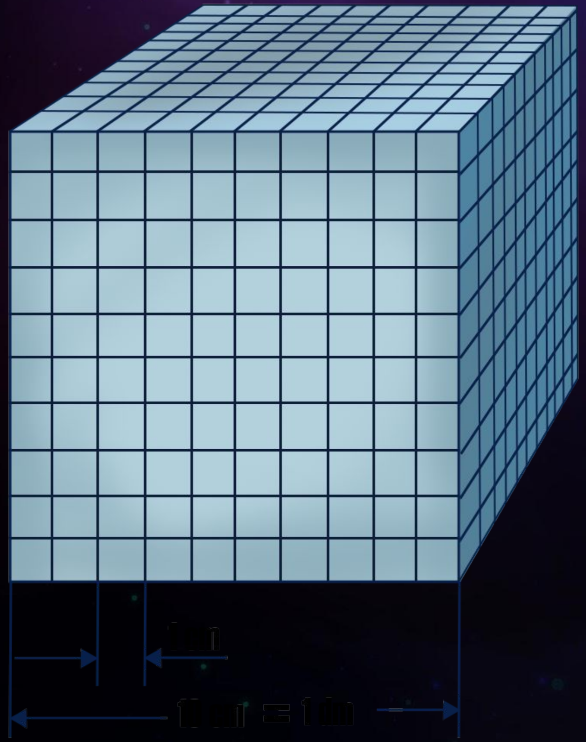
The S.I. unit of mass is kilogram

Volume

Volume is the **amount of space occupied** by a substance. It has the unit of $(\text{length})^3$

In S.I. system, volume has unit of m^3

Volume



Volume of one unit = $1\ cm^3$ = 1 mL

Total volume = $1000\ cm^3$ = 1000 mL

= $1\ dm^3$ = 1 L

Density

Density is the amount of **mass present per unit volume.**

It's S.I. unit is **kg m^{-3}**.

Temperature

There are three common scales to measure temperature

01
°C
(degree celsius)

02
°F
(degree fahrenheit)

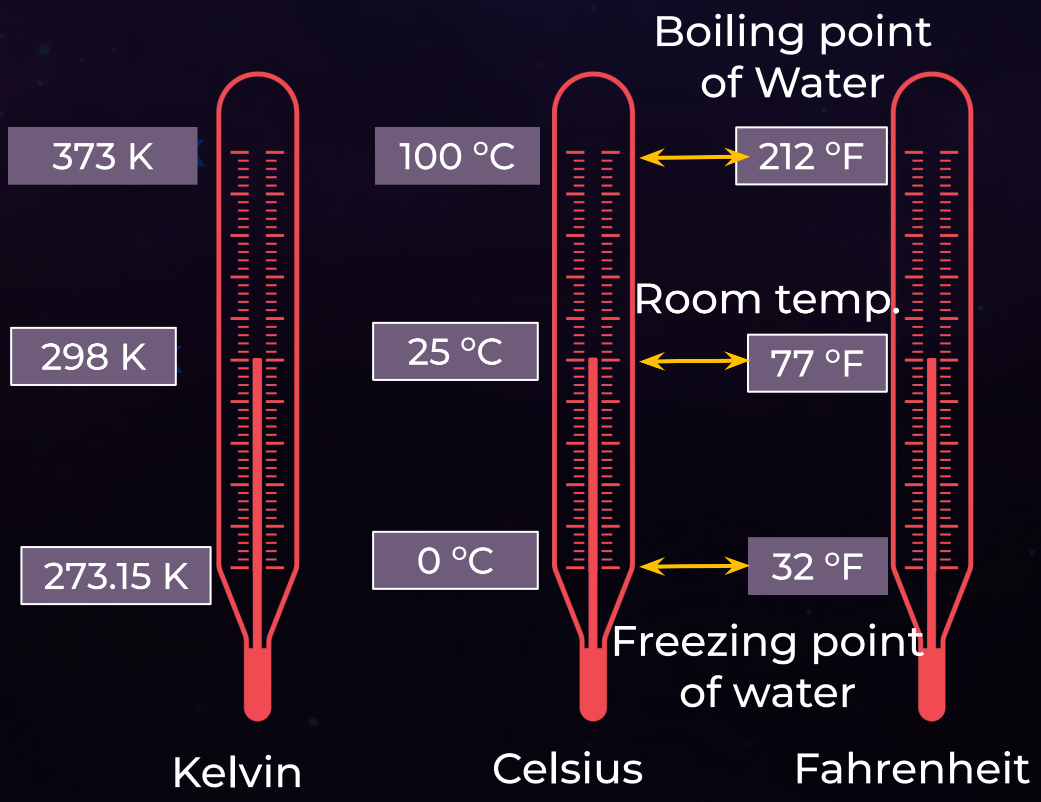
03
K
(kelvin)

S.I. unit is kelvin.

$$°F = \frac{9}{5}(°C) + 32$$

$$K = °C + 273.15$$

Temperature



Pressure

SI Unit = N/m^2 or Pa

Generally, **atm (atmospheric pressure)** is used in chemistry and sometimes other units are also used.

Pressure (Unit Conversions)

1 atm	=	1.01325 bar (approximated as 1 bar)
1 atm	=	760 mm Hg
1 atm	=	760 torr (Hence, 1 mm Hg = 1 torr)
1 atm	=	1.01325×10^5 N/m^2 or Pa; which can be taken approx. equal to 10^5 N/m^2

Scientific Notation

Any number can be represented in the form: $N \times 10^n$.

Where,
n - Exponent having any positive or negative values
N - Number which varies between 1.000... and 9.999....

For example

342.505 can be written as 3.42505×10^2 in scientific notation.

Addition and Subtraction in Scientific Notation

Example

$$3.425 \times 10^4 + 3.425 \times 10^3 = 3.425 \times 10^4 + 0.3425 \times 10^4$$

$$= 3.7675 \times 10^4$$

Multiplication and Division in Scientific Notation

In both the operations, exponents are **added and subtracted** as per their positive and negative values.

Example: **Multiplication**

$$10^x \times 10^y = 10^{x+y}$$

where x and y are integers

$$(3.425 \times 10^5) \times (3.425 \times 10^8) = 1.173 \times 10^{14}$$

Example: **Division**

$$\frac{10^x}{10^y} = 10^{x-y}$$

where x and y are integers

$$\frac{3.425 \times 10^5}{3.425 \times 10^8} = 1 \times 10^{5-8} = 1 \times 10^{-3}$$

Precision and Accuracy

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.

Precision and Accuracy

Rules for Determining the Number of Significant Figures

01 All **non-zero digits** are significant.

Example: 175 cm has **three** significant figures.

03 Zeros **between two non-zero digits** are significant.

Example: 2.005 has **four** significant figures.

02 Zeros **preceding to first non-zero** digit are not significant.

Example: 0.03 has **one** significant figure.

Rules for Determining the Number of Significant Figures

04

Zeroes at the **end or right of a number** are significant, provided they are on the right side of the decimal point.

Example: 6.00 cm has **three** significant figures.

05

Exact numbers have an **infinite number** of significant figures.

Example: In 4 pens or 40 copies, there are **infinite** significant figures.

Significant Figures

Addition and Subtraction of Significant Figures

In case of addition and subtraction, the final result should be reported to same number of decimal places as the number carrying **minimum number of decimal places** to the right.

Multiplication and Division of Significant Figures

In case of multiplication and division, the final result should be reported as having the same number of significant digits as the number with **least number of significant digits**.

Rounding Off

01

If the rightmost digit to be removed is more than 5, the **preceding number** is increased by one.

Example

1.386 is rounded off to 1.39.

Rounding Off

02

If the rightmost digit to be **removed is less than 5**, the preceding number is not changed.

Example

4.334 is rounded off to 4.33.

Rounding Off

03

If the rightmost digit to be removed is 5, then the preceding number is not changed if it is an **even number,** but it is increased by one if it is an **odd number**.

Example

6.35 is round off to 6.4. If 6.25 is to be rounded off it is rounded off to 6.2.

Dimensional Analysis

The unit factor by which multiplication is to be done is that unit factor which gives the **desired units.**

The **numerator** should have that part which is required in the **desired result**.

Units can be handled just like **other numerical part**.
It can be cancelled, divided, multiplied, squared, etc.

Dozen Analogy

No. of Dozens = No. of Items / 12

½ Dozen Classes	=	**6** Classes
3.5 Dozen Students	=	**42** Students
51 Chalks	=	**(51/12)** Dozen Chalks
	=	**4.25** Dozen Chalks

Definition of Mole

1 mole of a substance is defined as the number of entities same as the number of atoms present in **12 g of C^{12} isotope**.

This is equal to **Avogadro's number.**

Avogadro's number = 6.023×10^{23}

Mole-Particle Conversion

No. of Items = (No. of Mole) x N_A

½ Mole of Classes	=	**3.0115×10^{23}** Classes
	=	**(N_A/2)** Classes
3.5 Mole of Students	=	**21.0805×10^{23}** Students
	=	**3.5 N_A** Students
1 millimole of Chairs	=	**$6.023 \times 10^{23} \times 10^{-3}$** Chairs
	=	**(N_A/1000)** Chairs

Mole - Mass conversion

'Molar mass' g of the substance contains $\longrightarrow$ 1 mole of the substance.
Therefore,

'W' g of the substance contains $\longrightarrow$ $\dfrac{\textbf{W}}{\textbf{Molar mass}}$ **mole** of the substance

Hence,

$$\text{Moles} = \frac{\text{Mass}}{\text{Molar Mass}}$$

Atomic Mass Unit

The quantity $\frac{1}{12}$ x **(mass of an atom of C–12)** is known as **atomic mass unit**.

1 *amu* = 1 Dalton *(Da)* = 1 *u* where, **u** stands for unified mass

The actual mass of one atom of C-12 = **1.9924×10^{-26} kg**

$$\textbf{1 amu} = \frac{1.9924 \times 10^{-26}}{12}\ \text{kg}$$

$$= 1.66 \times 10^{-27}\ \text{kg}$$

$$= 1.66 \times 10^{-24}\ \text{g}$$

$$= 1/N_A\ \text{g}$$

Atomic Mass

Mass of an atom

Equal to summation of mass of **subatomic particles**

Neutron — n^0

Electron — e^-

Proton — p^+

Mass of Subatomic Particles

Neutron (n^o)	:	1.68×10^{-27} kg	or	1.68×10^{-24} g
Proton (p^+)	:	1.67×10^{-27} kg	or	1.67×10^{-24} g
Electron (e^-)	:	9.1×10^{-31} kg	or	9.1×10^{-28} g

Mass of Subatomic Particles

Mass (n^o)	:	mass (p^+)	:	mass (e^-)
1.68×10^{-24} g	:	1.67×10^{-24} g	:	9.1×10^{-28} g
1	:	0.994	:	$\frac{1}{1837}$
1.0087 amu		1.00728 amu		0.0005 amu

or

Mass (n^o) $\approx$ mass (p^+) >> mass (e^-)

Hence,
mass of electron is negligible w.r.t. proton or neutron

Atomic Number and Mass Number

$_{Z}E$

Element

"A" is Mass Number

"Z" is Atomic Number

Atomic number (Z) : Total no. of **protons**

Mass Number (A) : Total no. of (**protons + neutrons**)

Atomic Mass Unit

Atomic Mass = (**Mass Number**) x (**1.66 x 10^{-24} g**)

= (A) X (1 **amu**)

= (A) x (1 **u**)

= (A) x (1 **Da**)

Molecular Mass

Molecular mass numerically indicates the **mass of a molecule.**

Summation of mass of all the atoms that are contained in a molecule.

Example: Ammonia (NH_3)

1 N atom

Atomic mass = 14.00 u

3 H atoms

Atomic mass = 1.008 u

Molecular mass = (1 x 14.00 + 3 × 1.008) u

= 17.024 u

Molar Mass

Mass of 1 banana x 12 bananas = mass of 1 dozen banana

Similarly,

Mass of 1 atom x N_A atoms = molar mass (of atoms)

Mass of 1 molecule x N_A molecules = molar mass (of molecules)

Mass of 1 ion x N_A ions = molar mass (of ions)

Unit of Molar mass = g/mol

Molar Mass

Example:

$$\text{Mass of 1 Carbon atom} \times N_A = \text{Molar Mass of Carbon}$$

$$= 12\ u \times 6.022 \times 10^{23}$$

But, $\boxed{1\ u = \dfrac{1}{N_A}\ g}$

Therefore,

$$12\ u \times 6.022 \times 10^{23} = \frac{12 \times 6.022 \times 10^{23}}{6.022 \times 10^{23}}\ g$$

$$= 12\ g$$

$$= \text{Molar Mass of carbon}$$

g-Atomic Mass, g-Molecular Mass and g-Ionic Mass

g-atomic mass : It is the mass of 1 mole of atoms of a type in grams.

g-molecular mass : It is the mass of 1 mole of molecules of a type in grams.

g-ionic mass : It is the mass of 1 mole of ions of a type in grams.

Gram Atoms, Gram Molecules and Gram Ions

1 Gram Atom is 1 mole of atoms.

1 Gram Molecule is 1 mole of molecules.

1 Gram Ion is one mole of ions.

Example of Concept of Averages

क्रमांक
S.No. SSCE/ 2008/ 303622

केन्द्रीय माध्यमिक शिक्षा बोर्ड
CENTRAL BOARD OF SECONDARY EDUCATION
अंक विवरणिका MARKS STATEMENT
सीनियर स्कूल सर्टिफिकेट परीक्षा, 2008
ALL INDIA SENIOR SCHOOL CERTIFICATE EXAMINATION, 2008

नाम Name ABHINAV BISWAS
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माता का नाम Mother's Name APARNA BISWAS
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विषय कोड SUB. CODE	विषय SUBJECT	प्राप्तांक MARKS OBTAINED				स्थितीय ग्रेड POSITIONAL GRADE
		लि. TH	प्र. PR	योग TOTAL	योग शब्दो में TOTAL IN WORDS	
301	ENGLISH CORE	087	XXX	087	EIGHTY SEVEN	A1
041	MATHEMATICS	095	XXX	095	NINETY FIVE	A1
042	PHYSICS	058	030	088	EIGHTY EIGHT	A1
043	CHEMISTRY	066	029	095	NINETY FIVE	A1
083	COMPUTER SCIENCE	062	030	092	NINETY TWO	A1
048	PHYSICAL EDUCATION	065	023	088	EIGHTY EIGHT	A1
500	WORK EXPERIENCE					B1
502	PHY & HEALTH EDUCA					A1
503	GENERAL STUDIES					A2

संक्षिप्तियों का अर्थ : Abbreviations
AB : विषय में अनुपस्थित Absent in the Subject
EX : छूट - प्राप्त Exempted
FP : प्रयोगात्मक में असफल Fail in Practical
FT : लिखित में असफल Fail in Theory

परिणाम Result PASS

दिल्ली Delhi
दिनांक Dated 23-05-2008

MSharma
परीक्षा नियंत्रक
Controller of Examinations

Example of Concept of Averages

Solution:

$$\text{The average marks scored by the student} = \frac{\text{Sum of marks scored by the student in each subject}}{\text{Total number of subjects}}$$

$$= \frac{87 + 95 + 88 + 96 + 92 + 88}{6} = \mathbf{91}$$

Similarly, in case of elements, a sample consist of more than 1 kind of atoms called isotopes.

Therefore,

Mass of a sample of atoms is also represented as weighted average mass and is called as **average atomic mass.**

Isotopes

Isotopes are those particles which have same number of protons but different number of neutrons.

$_{1}H^{1}$ Hydrogen	$_{1}H^{2}$ Hydrogen	$_{1}H^{3}$ Hydrogen	$_{8}O^{16}$ Oxygen	$_{8}O^{18}$ Oxygen
Protium	**Deuterium**	**Tritium**	**^{16}O Isotope**	**^{18}O Isotope**

Percentage Abundance

Percentage abundance is defined as the **percentage value** of the **quantity of isotopes** available in nature for a **given element.**

Average Atomic mass

Average atomic mass = **$(\%\ \text{abundance})_1\ \text{Mass}_1 + (\%\ \text{abundance})_2\ \text{Mass}_2 + \ldots$**

Example: Carbon-12 → 99%, Carbon-13 → 1%

Average atomic mass of carbon= $(12 \times 0.99) + (13 \times 0.01)$ g = 12.01 g

Hence, 12.01 is the average atomic mass of carbon, whereas, atomic mass of carbon-12 is 12 u.

Average Molecular Mass

$$= \frac{(\text{no. of molecules})_1 \times (\text{Molecular Mass})_1 + (\text{no. of molecules})_2 \times (\text{Molecular Mass})_2 + \ldots}{\text{Total no. of particles}}$$

Specific Activity

Specific Activity(A) is defined as the activity of 1 g of substance (pure or impure) and is given by $A = \lambda \times N$ where N is the number of radioactive atoms.

Unit = yr^{-1}

Dalton’s Atomic Theory

- Matter consists of **tiny** particles called **atoms**.
- All the atoms of a given element have **identical** properties including identical **mass**. Atoms of different elements differ in mass.
- Compounds are formed when **atoms** of **different** elements combine in a **fixed** ratio.
- Chemical reactions involve **rearrangement** of atoms.

Drawbacks

- According to the theory, atom is indivisible but atom can be **divided** into **electrons**, **neutrons** and **protons**.

- Atoms of same element can have **different masses** as in case of **isotopes** and **isobars**.

- Reactions that **does not** react with **simple** whole number **ratio** of reactants are known as **non-stoichiometric reactions.**

Law of Conservation of Mass

Statement: It states that **matter can neither be created nor destroyed** in ordinary chemical and physical changes.

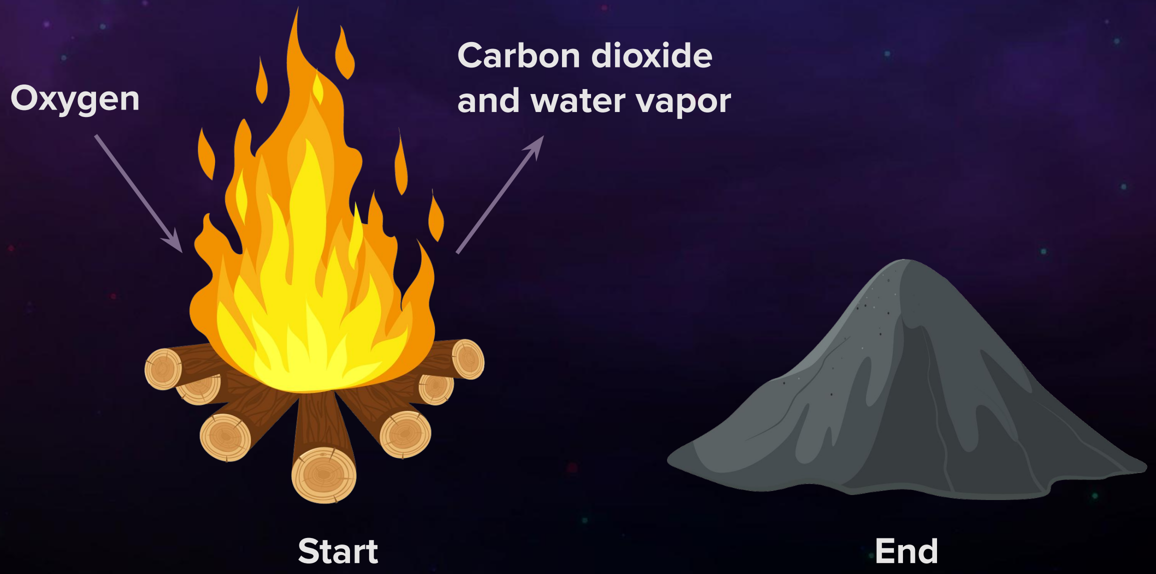
Explanation: In a chemical reaction, the **total mass of reactants** is always equal to the **total mass of products** formed.

Law of Conservation of Mass

Reactants:
4 H + 2 O
atoms

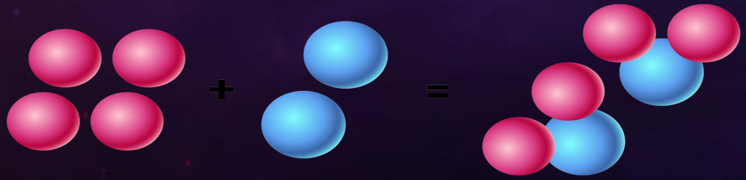
Products:
4 H + 2 O
atoms

Let's Understand



Law of Definite Proportions

Statement: A given compound always contains **exactly the same proportion** of elements by weight **irrespective** of the **source** or **method of preparation**.



2 : 1

4 parts hydrogen

2 parts oxygen

2 gaseous water

2 H atoms : 2 amu

+

1 O atoms : 16 amu

Molecular weight of water = 18 amu

% Hydrogen by weight : 11.11%

% Oxygen by weight : 88.89%

Law of Multiple Proportions

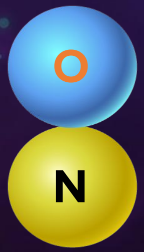
Statement:

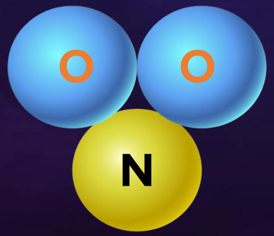
If **two elements can combine** to form more than one compound, the **masses of one element** that combine with a **fixed mass of the other element**, are in the **ratio of small whole numbers**.

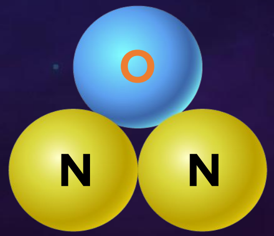
Alternate way to phrase:

If **two elements can combine** to form more than one compound, for a **fixed mass of the any one element** in both the compounds, the **ratio of masses of the other element** in the two compounds comes out to be in **small whole numbers**.

Law of Multiple Proportions







NO	NO_2	N_2O
14 : 16	14 : 32	28 : 16
1 : 1	1 : 2	2 : 1

Law of Multiple Proportions

Compound	Mass ratio of N : O	Taking fixed mass of N	Divide by lowest amount of oxygen reacting	Ratio of oxygen
N_2O	28 : 16	28 : 16	16/16	1
NO	14 : 16	28 : 32	32/16	2
N_2O_3	28 : 48	28 : 48	48/16	3
NO_2	14 : 32	28 : 64	64/16	4
N_2O_5	28 : 80	28 : 80	80/16	5

Law of Multiple Proportions

	CO	CO_2	C_3O_2
Mass ratio of C : O	12 : 16	12 : 32	36 : 32
Taking fixed mass of C as 36 g/mol	36 : 48	36 : 96	36 : 32

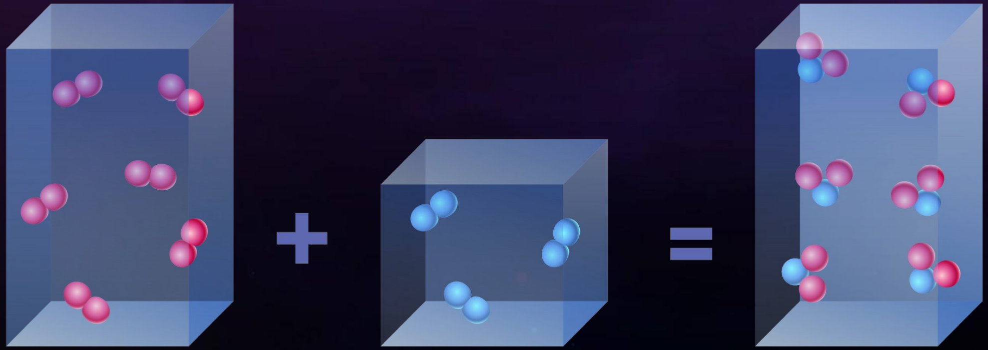
Law of Multiple Proportions

Taking mass of O and dividing by lowest mass, which is 32 we get:

	CO	CO_2	C_3O_2
	48 : 32	96 : 32	32 : 32
Ratio of O	$\overset{3}{\cancel{48}} : \overset{2}{\cancel{32}}$	$\overset{3}{\cancel{96}} : \overset{1}{\cancel{32}}$	$\overset{1}{\cancel{32}} : \overset{1}{\cancel{32}}$
	1.5	3	1
	3	6	2

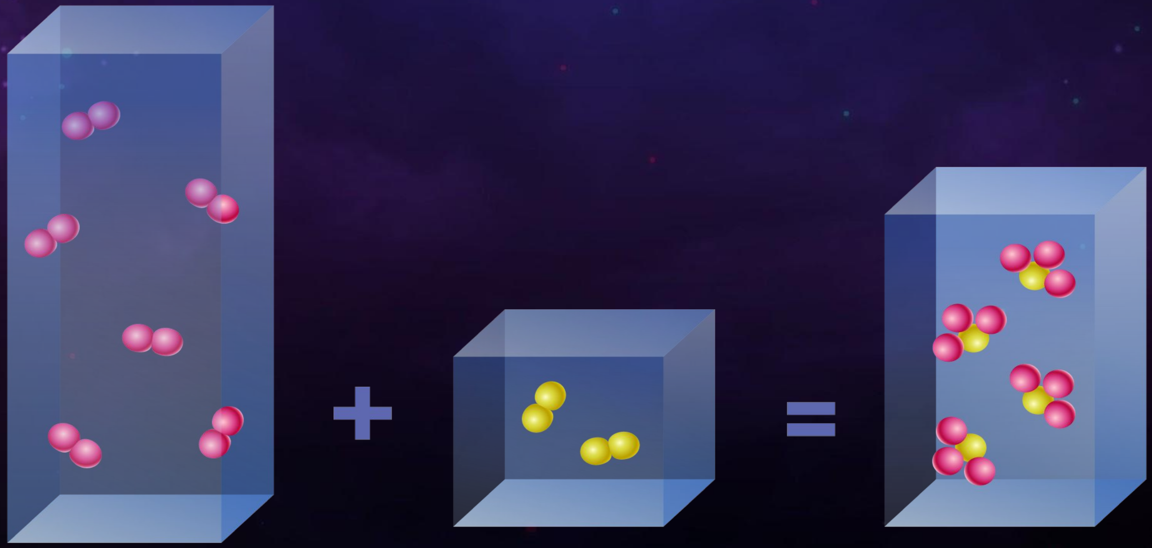
Gay Lussac's Law

Statement: 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**.



$$2H_2\,(g) + O_2\,(g) \longrightarrow 2H_2O\,(g)$$

Gay Lussac's Law



$$3H_2\,(g) \quad + \quad N_2\,(g) \longrightarrow 2NH_3\,(g)$$

Avogadro Law

Example: Comparison of some gases at same temperature and pressure

He N_2 CH_4

Avogadro Law

Statement: **Equal volumes** of gases at the **same temperature and pressure** should contain **equal** number of **molecules/ particles.**

$$V \propto n$$

$$\Rightarrow \frac{V}{n} = k$$

$$\Rightarrow \frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Relative density

It is the density of substance **relative** to the density of another substance, at the same T and P.

$$R.D. = \frac{\rho_{substance}}{\rho_{reference}}$$

It is a **dimensionless** quantity.

Specific gravity

Specific gravity for **liquids** is measured with respect to **water** at 4°C.

$$\text{S.G. (for liquid)} = \frac{\rho_{\text{substance}}}{\rho_{\text{water at } 4^\circ\text{C}}}$$

For **gases** it is measured with respect to **air** at STP.

$$\text{S.G. (for gas)} = \frac{\rho_{\text{gas at STP}}}{\rho_{\text{air at STP}}}$$

It is a **dimensionless** quantity.

Vapour density

Mathematical form - Density of the gas with respect to **hydrogen** gas at the **same** temperature and pressure

Significance - It indicates **how heavy a gas is** with respect to the **lightest gas**.

It is a **dimensionless** quantity.

Vapour density

$$\text{Vapour Density} = \frac{\textbf{density} \text{ of gas A at some T and P}}{\textbf{density} \text{ of } H_2 \text{ gas at same T and P}}$$

$$= \frac{\textbf{mass} \text{ of gas A } \textbf{in 1 mL} \text{ at some T and P}}{\textbf{mass} \text{ of } H_2 \text{ gas } \textbf{in 1 mL} \text{ at same T and P}}$$

$$= \frac{\textbf{N particles} \times (\textbf{mass of one particle} \text{ of gas A})}{\textbf{N particles} \times (\textbf{mass of one particle} \text{ of } H_2 \text{ gas})}$$

$$= \frac{\text{M amu}}{\text{2 amu}}$$

$$= \frac{M}{2}$$

Molar Volume at STP

S.T.P. (Standard Temperature and Pressure)

At STP condition: Temperature = **0°C** or **273 K**

Pressure = **1 Bar**

Molar volume at STP = **22.7 L**

Molar volume = volume of 1 mole of gas

1 atm = 1.01325 Bar

Molar Volume at STP

$\rho(\mathbf{H_2}$ at $\mathbf{STP}) = 0.089$ g/L

Now,

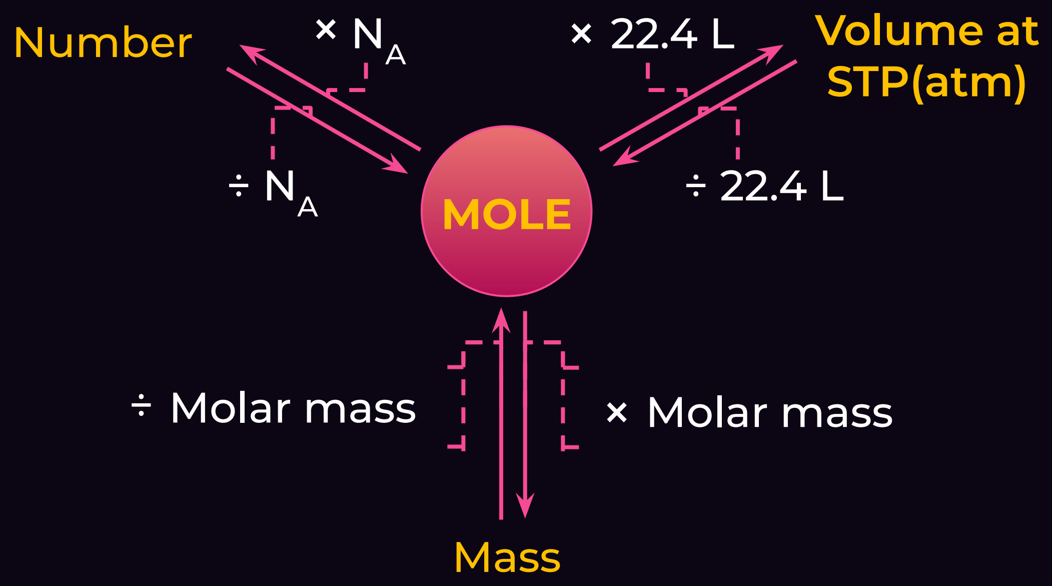
$$\text{Density} = \frac{M}{V}$$

$$0.089 \text{ g/L} = \frac{2 \text{ g/mol}}{V}$$

$$\text{Hence, molar volume } V_m(\text{L/mol}) = \frac{2 \text{ g/mol}}{0.089 \text{ g/L}}$$

$\mathbf{V_m}$ = **22.7 L** at **STP**

Interconversion of mole-volume, mass and number of particles



Percentage Composition

Defined as mass of an element present in 100 g compound.

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

Molar mass of CO_2 = 12 + 2(16) = 44 g

Example: CO_2

$$\text{Mass \% of Carbon} = \frac{12}{44} \times 100 = 27.27\ \%$$

$$\text{Mass \% of Oxygen} = \frac{(2 \times 16)}{44} \times 100 = 72.72\ \%$$

Empirical and Molecular Formula

Empirical Formula: It represents the simplest **whole number** ratio of various atoms present in a compound.

Determined by the **mass percent** of various elements present in the compound.

$C_6H_{12}O_6$

CH_2O

Empirical and Molecular Formula

Molecular Formula: It shows the **exact** number of **different** types of **atoms** present in a molecule of a compound.

Obtained by using the **molar mass** of each element.

$C_6H_{12}O_6$

CH_2O

Empirical and Molecular Formula

Molecular formula = Empirical formula × n

$$n = \frac{\text{Molecular formula mass}}{\text{Empirical formula mass}}$$

$C_6H_{12}O_6$

CH_2O

Empirical and Molecular Formula

Empirical formula mass for $\mathbf{CH_2O}$ = 12 + 2 + 16 = 30

$$n = \frac{\text{Molecular formula mass}}{\text{Empirical formula mass}}$$

$$\boxed{n = \frac{180}{30} = 6}$$

Molecular formula = $\mathbf{C_nH_{2n}O_n}$

Thus, **Molecular formula** is $\mathbf{C_6H_{12}O_6}$

Principle of Atom Conservation

Total number of atoms of a particular element is **same** in **reactants and products**.

Balancing Equations

$C + O_2 \longrightarrow CO$ Not Balanced

$2C + O_2 \longrightarrow 2CO$ Balanced

$Zn + HCl \longrightarrow ZnCl_2 + H_2$ Not Balanced

$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$ Balanced

Balancing Equations

$NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O$ Not Balanced

$2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O$ Balanced

Balancing Equations

$$Ca_3(PO_4)_2 + SiO_2 + C \rightarrow CaSiO_3 + CO + P_4$$

Not Balanced

$$2Ca_3(PO_4)_2 + 6SiO_2 + 10C \rightarrow 6CaSiO_3 + 10CO + P_4$$

Balanced

Stoichiometry

Calculations based on the **quantitative** relationship (**mole /mass/volume/ number of molecules**) between the **reactants** and the **products.**

$$1\,N_2(g) + 3\,H_2(g) \rightarrow 2\,NH_3(g)$$

For a balanced chemical reaction these are called the **stoichiometric coefficients.**

Stoichiometry

$$1\,N_2(g) + 3\,H_2(g) \rightarrow 2\,NH_3(g)$$

	$1\,N_2(g)$	$3\,H_2(g)$	$2\,NH_3(g)$
Initial moles	10	30	0
	-2 moles	**-6 moles**	**+4 moles**
Intermediate moles	8	24	4
	-8 moles	**-24 moles**	**+16 moles**
Final moles	0	0	20

Stoichiometry

The **reactant** reacts and the **products** are formed in the **molar ratio**, which is **same** as the **ratio of their stoichiometric coefficients**.

The stoichiometric coefficients **a**, **b**, **c** or **1**, **3**, **2** are **not the given number of molecules or the number of moles**.

$$a\,N_2(g) + b\,H_2(g) \longrightarrow c\,NH_3(g)$$

$$1\,N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g)$$

Stoichiometric coefficients sum i.e. $1 + 3 \neq 2$

i.e., total **sum of stoichiometric coefficients of reactants** need **not be equal** to the total **sum of stoichiometric coefficients of products.**

Stoichiometry

$$2\,H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O\,(g)$$

or

$$H_2\,(g) + \frac{1}{2}\,O_2\,(g) \longrightarrow H_2O\,(g)$$

This **doesn't** mean that $\frac{1}{2}$ O_2 molecule is reacting.

Stoichiometry

It means that when $\mathbf{H_2}$, $\mathbf{O_2}$ combine and $\mathbf{H_2O}$ is produced, it happens in a **1 : 1/2 : 1** ratio or **2 : 1 : 2**

$$1\,N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g)$$

$$\frac{\mathbf{n}\,(N_2,\text{consumed})}{\mathbf{n}\,(H_2,\text{consumed})} = \frac{1}{3}$$

$$\frac{\mathbf{n}\,(N_2,\text{consumed})}{\mathbf{n}\,(NH_3,\text{produced})} = \frac{1}{2}$$

$$\frac{\mathbf{n}\,(H_2,\text{consumed})}{\mathbf{n}\,(NH_3,\text{produced})} = \frac{3}{2}$$

Stoichiometry

$$1\,N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g)$$

$$\frac{\mathbf{n}\,(N_2,\ \text{consumed})}{1} = \frac{\mathbf{n}\,(H_2,\ \text{consumed})}{3}$$

$$\frac{\mathbf{n}\,(N_2,\ \text{consumed})}{1} = \frac{\mathbf{n}\,(NH_3,\ \text{produced})}{2}$$

$$\frac{\mathbf{n}\,(H_2,\ \text{consumed})}{3} = \frac{\mathbf{n}\,(NH_3,\ \text{produced})}{2}$$

$$\frac{\mathbf{n}\,(N_2,\ \text{consumed})}{1} = \frac{\mathbf{n}\,(H_2,\ \text{consumed})}{3} = \frac{\mathbf{n}\,(NH_3,\ \text{produced})}{2}$$

Stoichiometry

$$\frac{\mathbf{n}\ (N_2,\ \text{initial})}{1} \neq \frac{\mathbf{n}\ (H_2,\ \text{initial})}{3} \neq \frac{\mathbf{n}\ (NH_3,\ \text{final})}{2}$$

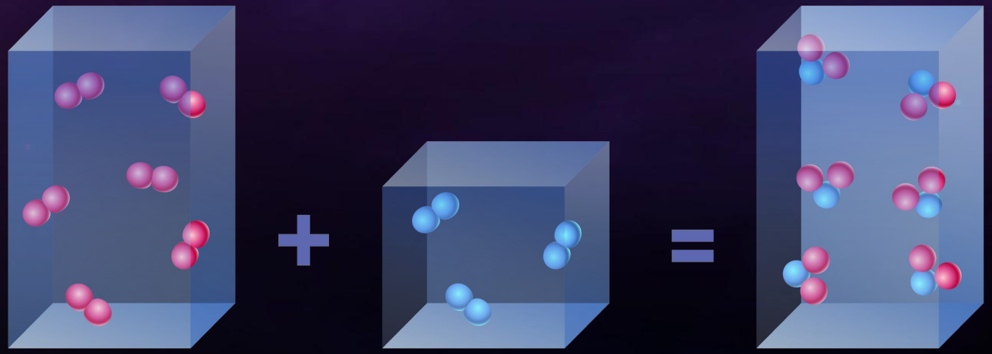
Necessarily Necessarily

$$\frac{\mathbf{n}\ (N_2,\ \text{cons.})}{1} = \frac{\mathbf{n}\ (H_2,\ \text{cons.})}{3} = \frac{\mathbf{n}\ (NH_3,\ \text{prod.})}{2}$$

Definitely Definitely

Conclusion

Stoichiometry can be applied to **volume of gaseous reactants** and **products** at same **temperature and pressure.**



$$2H_2\,(g) \quad + \quad O_2\,(g) \longrightarrow H_2O\,(g)$$

Limiting Reagent

The **limiting reagent (or limiting reactant)** is the one which is consumed first in a chemical reaction and determines the amount of product formed.
We will take an example to understand this concept.

Example:

$$2A + 3B \longrightarrow 4C + 5D$$

Starting with **8** moles of **A** and **6** moles of **B**, find the moles of **C** formed.

Here, 8 moles of A need 12 moles of B (because stoichiometric ratio is 2:3) but we have only 6 moles of B.

So, B is the limiting reagent and A is in excess. So, only 4 moles of A is consumed with 6 moles of B and 8 moles of C and 10 moles of D are formed.

% Yield

The reactions not always yield 100 % but sometimes its production is lesser than the theoretical amount(calculated amount).

In such case, % yield can be given as:

$$\% \text{ yield} = \frac{\text{Actual amount of product}}{\text{Theoretical amount of product}} \times 100$$

% Yield

Let's see an example:

Example:

$$2A + 3B \longrightarrow 4C + 5D \qquad\qquad \text{yield} = 70\%$$

Starting with 8 moles of A and 6 moles of B, find the moles of C formed.
Here, B is the limiting reagent because for its 6 moles to react, only 4 moles of A is required.

We know that the amount of product will be determined only by limiting reagent.

So, after consumption of 6 moles of B, (4/3) x 6 moles = 8 moles of C formed (if yield is 100%).

But it is 70%, so the amount of moles of C = 8 x (70/100) = 5.6 moles.

Percentage Purity

Percentage of a specified compound or element in an **impure sample**.

$$\% \text{ Purity} = \frac{\text{Actual amount of desired species in the sample}}{\text{Total amount of the sample}} \times 100$$

What is a solution?

Homogeneous mixture of two or more substances

Concentration Terms

Molarity

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

Mass of solute in **grams** present in **per 100 mL of solution**

SI unit

mol L^{-1}

mol dm^{-3}

M

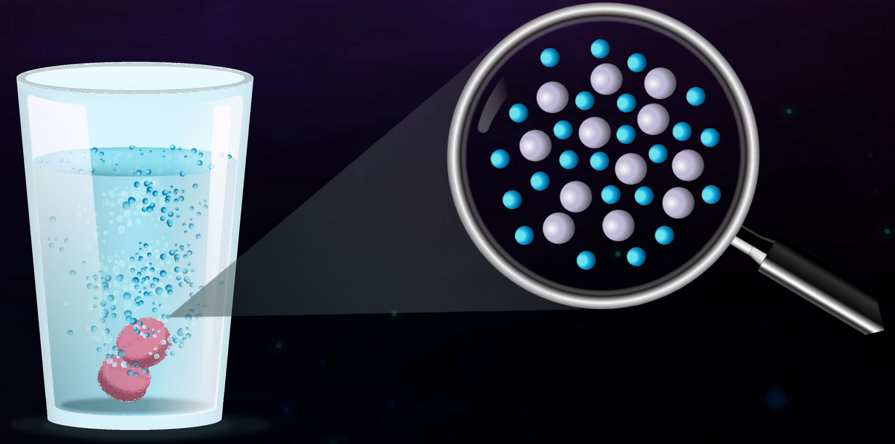
Molar

Concentration Terms

Percentage concentration (%w/v)

$$\% \text{ w/v} = \frac{\text{mass of solute in g}}{\text{volume of solution in mL}} \times 100$$

Number of moles of the **solute** present in **1 litre** of the **solution**



Let's understand

0.9% w/v means **0.9 g** of **solute** present in **100 mL** of **solution**

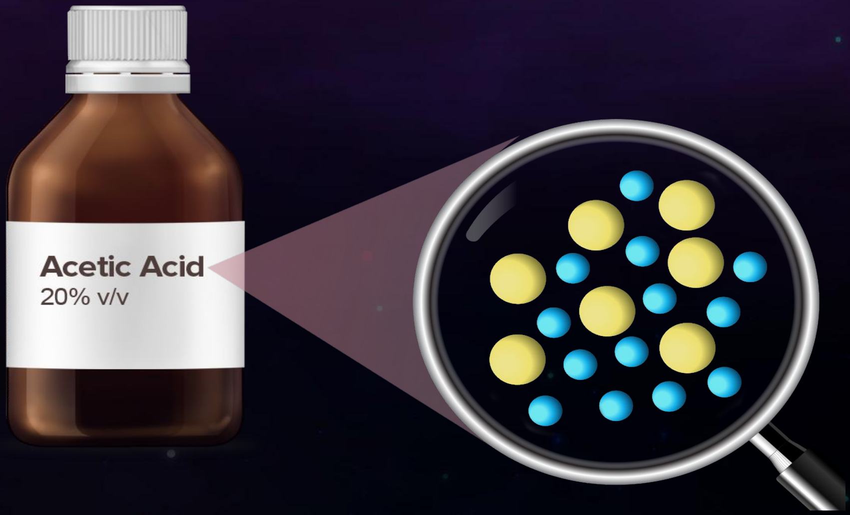
Percentage concentration (%v/v)

Volume of solute in **mL** present in **per 100 mL** of **solution**

$$\% \ v/v = \frac{\text{volume of solute in mL}}{\text{volume of solution in mL}} \times 100$$

Let’s understand

20% v/v or by volume means **20 mL** of **solute** present in **100 mL** of **solution**



Strength of solution

Unit = g/L

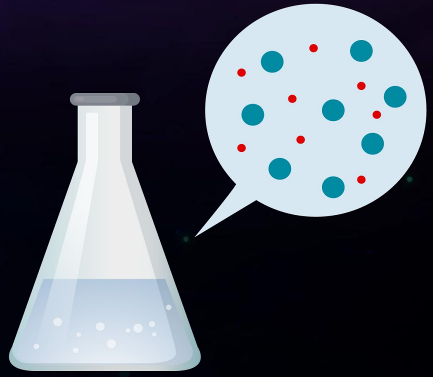
Concentration of solution expressed in **gram/litre**

Molality

Number of moles of solute present in 1 kg of solvent

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

SI unit is **mol kg^{-1}** or **m(molal)**



1.0 Molal NaOH

Mole fraction

If substance '**A**' is dissolved in substance '**B**' and $\mathbf{n_A}$ and $\mathbf{n_B}$ are their respective moles, then

Ratio of the number of moles of a particular **component to the total number of moles** of the solution

$$\textbf{Mole fraction of A} = \frac{\textbf{No. of moles of A}}{\textbf{No. of moles of solution}}$$

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

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

Mole fraction

For a solution containing **i** number of components, we have

Mole fraction is a **pure number** and has **no units.**

$$X_i = \frac{n_1}{n_1 + n_2 + ... + n_i} = \frac{n_1}{\Sigma n_i}$$

where, $X_1 + X_2 + ... + X_i = 1$

Mole Percentage

If substance ‘**A**’ is dissolved in substance ‘**B**’, n_A and n_B are their respective moles, then

Ratio of the **number of moles** of a particular **component** to the **total number of moles** of the Solution multiplied by 100.

$$\text{Mole percentage of A} = \frac{\text{No. of moles of A}}{\text{No. of moles of solution}} \times 100$$

Mole Percentage

Mass of solute present in **per 100 g** of **solution**

$$\% \, w/w = \frac{\text{mass of solute in g}}{\text{mass of solution in g}} \times 100$$

Let’s understand

72% w/w or by weight means **72 g** of **solute** present in **100 g** of **solution**

TRIMANOC 72% w/w

Parts per million (ppm, 10^{-6})

Number of parts of the **solute** present in **every 1 million parts** of the **solution**

$$\text{ppm (w/w)} = \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \times 10^6$$

$$\text{ppm (w/v)} = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 10^6$$

$$\text{ppm (v/v)} = \frac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}} \times 10^6$$

PPM Significance

CO_2 [ppm]	Air Quality
2100	**BAD** Heavily contaminated Indoor air Ventilation required
2000	
1900	
1800	
1700	
1600	
1500	**MEDIOCRE** Contaminated indoor air Ventilation recommended
1400	
1300	
1200	
1100	
1000	**FAIR**
900	
800	**GOOD**
700	
600	**EXCELLENT**
500	
400	

Nitrite (NO_2^-)

- 0 ppm
- 0.25 ppm
- 0.50 ppm
- 1.0 ppm
- 2.0 ppm
- 5.0 ppm

Nitrate (NO_3^-)

- 0 ppm
- 5.0 ppm
- 10 ppm
- 20 ppm
- 40 ppm
- 80 ppm
- 160 ppm

Parts Per Billion (ppm, 10^{-9})

Number of parts of the **solute** present in **every 1 billion parts** of the **solution**

$$\text{ppb} = \frac{\text{Mass of solute (g)}}{\text{Mass of solution (g)}} \times 10^9$$

Interconversiom of Concentration Terms

Relation between Molarity and Strength

$$S = M \times M_o$$

Where,

S = Strength of solution

M = Molarity of solution

M_o = Molar mass

Relation between % w/v and % w/w

$$\% \text{ w/v} = \% \text{ w/w} \times \rho$$

Where,

ρ = density of the solution

Interconversiom of Concentration Terms

Relation between Molarity and %/v

$$M = \frac{10}{M_{solute}} \times \% \ w/v$$

M_{solute} = Molar mass of the solute

Relation between Molarity and Molality

$$m = \frac{1000 \times M}{1000 \times \rho - M \times M_{solute}}$$

Interconversiom of Concentration Terms

Relation between Molality and Mole fraction

$$m = \frac{X_{solute}}{X_{solvent}} \times \frac{1000}{M_{solvent}}$$

M = Molality

χ = Mole fraction

$M_{solvent}$ = Molecular mass of solvent

Interconversiom of Concentration Terms

Relation between Molarity and Mole fraction

$$M = \frac{X_B \times 1000 \times \rho}{X_A M_A \times X_B M_B}$$

Where,

ρ = Density of the solution

X_B , X_B = Mole fraction

M_A, M_B = Molar mass

M = Molarity

Dilution

Consider a solution having volume $\mathbf{V_1}$ and molarity $\mathbf{M_1}$

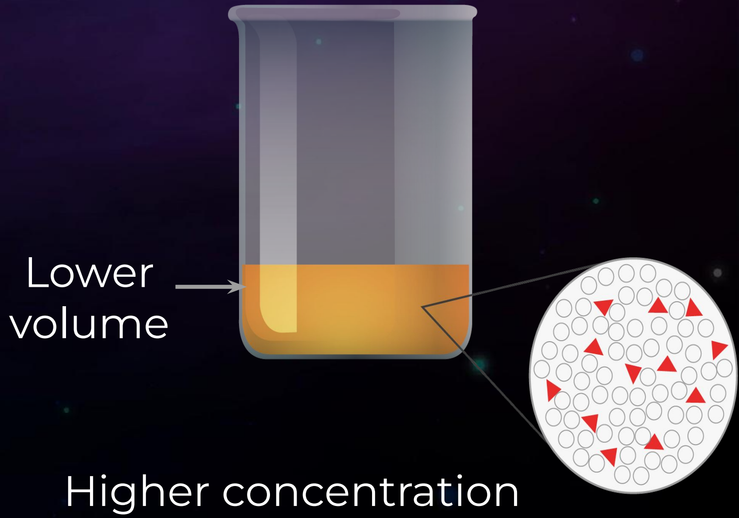
Dilute it upto volume $\mathbf{V_2}$ mL

Concentration decreases to $\mathbf{M_2}$

M_1

V_1

Before adding solvent

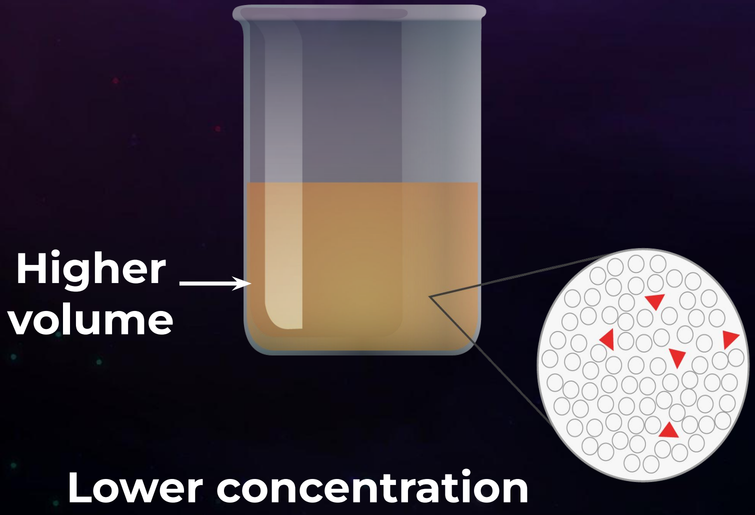


Dilution

After adding solvent

M_2

V_2



$$M_1V_1 = M_2V_2$$

M_1 = Initial molarity of the solution
M_2 = Final molarity of the solution
V_1 = Initial volume of the solution
V_2 = Final volume of the solution

Dilution

150 mL of water (further added)

150 mL of 2 Molar
Aq. solution of NaCl

300 mL of 1 Molar
Aq. solution of NaCl

Addition or mixing

Solution 1 having volume $\mathbf{V_1}$ and molarity $\mathbf{M_1}$

Mixed with another solution of **same solute**

Solution 2 having volume $\mathbf{V_2}$ & molarity $\mathbf{M_2}$

$$M_1V_1 + M_2V_2 = M_RV_{Final}$$

M_R = Resultant molarity of the solution

V_{final} is not necessarily equal to $(V_1 + V_2)$

Addition or mixing

Resultant Molarity (M_R) $= \dfrac{M_1 V_1 + M_2 V_2}{V_{Final}}$

150 mL of 2 M NaCl (aq.) + 150 mL of 2 M NaCl (aq.)

Addition or mixing

M_1	M_2	M_R
V_1	V_2	V_{Final}
n_1	n_2	$n_1 + n_2$

Addition or mixing

$$M_{mix} = \frac{\text{Total Moles}}{\text{Total Volume}} = \frac{n_1 + n_2}{V_{Total}}$$

Neutralisation - Mixing of acid and base

Acid + **Base**

ACID

Solution 1 having volume $\mathbf{V_1}$ and molarity of $\mathbf{H^+}$ ions $\mathbf{M_1}$

Mixed with another solution of **different solute**

BASE

Solution 2 having volume $\mathbf{V_2}$ & molarity of $\mathbf{OH^-}$ ions $\mathbf{M_2}$

Neutralisation - Mixing of acid and base

$$M_R = \frac{M_1V_1 - M_2V_2}{V_1 + V_2}$$

M_R = Resultant molarity of the solution

Gravimetric Analysis

Gravimetric analysis by weight is the process of isolating and weighing an element or a definite compound of an element in as pure form as possible.

Gravimetric analysis is an analytical technique based on the measurement of mass of solid substances and/or volume of gaseous species.

Gravimetric Analysis

Consider this reaction :

$CaCO_3(s)$	$\longrightarrow$	$CaO(s)$	+	$CO_2(g)$
1 mol		1 mol		1 mol
100 g		56 g		44 g or 22.4 L at S.T.P

Equivalent Mass

Number of parts by mass of an element which reacts or displaces from a compound **1.008** parts by mass of **hydrogen**, **8** parts by mass of **oxygen** or **35.5** parts by mass of **chlorine.**

$$2Mg + O_2 \longrightarrow 2MgO$$

32 g O_2 reacts with 48 g of Mg

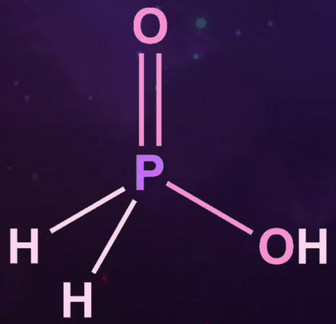
$$8 \text{ g of } O_2 = \frac{48 \times 8}{32} = 12 \text{ g}$$

Equivalent weight of Mg = 12 g

Equivalent Mass

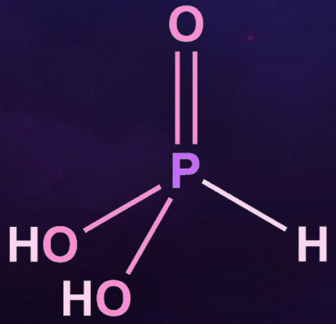
For **Acids**	Valency factor	=	Number of replaceable H^+ ions per molecule
For **Bases**	Valency factor	=	Number of replaceable OH^- ions per molecule
For **Salts**	Valency factor	=	Total number of +ve or -ve charge present in the compound

Some Special Cases of Acids



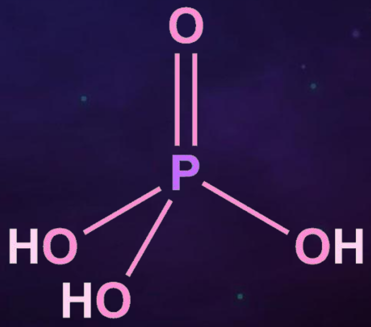
H_3PO_2

$n_f = 1$



H_3PO_3

$n_f = 2$



H_3PO_4

$n_f = 3$

Equivalents

An equivalent is an amount of a substance that reacts with a definite amount of another substance in a given chemical reaction.

Significance of n-Factor

$$\text{Equivalents} = n \times n_f = M \times V \times n_f$$

Where,

n = Number of moles

M = Molarity of the solution

n_f = n-factor

V = Volume of the solution

$$\text{Equivalent mass} = \frac{\text{Molar mass}}{n_f}$$

Equivalent Mass of Different Species

For Elements:

Amount of an element that reacts with or displace 1 g of hydrogen, 8 g of oxygen or 35.5 g of chlorine.

$$\text{Equivalent Mass} = \frac{\text{Atomic Mass}}{\text{Valency of the Element}}$$

For Ions:

$$\text{Equivalent Mass} = \frac{\text{Formula Mass of Ion}}{\text{Charge on Ion}}$$

Equivalent Mass of Different Species

For Ionic compounds:

$$\text{Equivalent Mass} = \frac{\text{Molecular Mass}}{\text{Charge on Constituent Cation or Anion}}$$

For Acid:

$$\text{Equivalent Mass} = \frac{\text{Molecular Mass}}{\text{Basicity of Acid}}$$

For Base:

$$\text{Equivalent Mass} = \frac{\text{Molecular Mass}}{\text{Acidity of Base}}$$

Normality

Number of gram
equivalents of solute
dissolved
per litre of solution

Normality

$$\text{Normality} = \frac{\text{Number of equivalents of solute}}{\text{Volume of solution (L)}}$$

$$\text{Number of equivalents} = \frac{\text{Mass of the species}}{\text{Equivalent mass}}$$

$$= \frac{\text{Mass of the species}}{\dfrac{\text{Molar mass}}{\text{n - factor}}}$$