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Some Basic Concepts of Chemistry

Chemistry :- The branch of science that deals with the studies of the preparation, properties, structure and reaction of material substances is called chemistry.

Some Facts that tells about chemistry in Ancient India :-

- In India, chemistry was called Rasayan Shastra, Rastantra, Ras kriya or Rasvidya.
- Rigueda, describes tanning of leather and dying of cotton.
- Kautilya's Arthashastra :- describes the production of Salt From sea.
- Sushruta Samhita :- explains importance of Alkalies.

- Chakra samhita :- maintioned Indians who knew how to prepare sulphuric acid, nitric acid, etc. It is the oldest Ayurvedic epic of India

- Rasopanishada describes the preparation of gunpowder mixture.

- Chakrapani discovered mercury sulphide.

- Varahmihir's Brihat Samhita informs preparation of glutinous material to be applied on walls walls of house & temples.

Chemistry is devided into following branches :-

1.1. Physical Chemistry

The branch of chemistry which deals with macroscopic as well as physical phenomenon in a universe. It is generally the impact of physical property on the chemical property as well as the structure of substance.

1.2 Inorganic Chemistry

The branch of chemistry that studies compounds that do not contain carbon and hydrogen atoms is called 'inorganic chemistry.' simply put, it is the opposite of organic chemistry. substances that do not have carbon - hydrogen bonds include metals, salts and chemicals.

1.3 Organic chemistry

The discipline which deals with the study of the structure, com...... and the chemical properties of organic componds is known as organic chemistry. It envolves the study of carbon

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and its compounds.

1.4 Biochemistry

Biochemistry is that branch of chemistry that explores the chemical process in organisms and associated with them. it's a laboratory-based science that connects biology and chemistry. By using chemical knowledge and technology, biochemists can under-stand and solve biological problems.

1.5 Analytical chemistry

It is the branch of chemistry which uses instruments and analytical techniques to determine structure, functionality and properties of a substance.

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Nature of Matter

Matter :- Anything which has mass and occupies space.

States of matter

(i) Solids have definite shape and definite volume.
(ii) Liquids have definite volume but do not have definite shape.
(iii) Gases have neither fixed volume nor fixed shape.

- Matter
  - mixture
    - Homogenous
    - Heterogenous
  - Pure substances
    - Elements
    - Compounds

Pure substance :- All constituent particles of a substance are same in chemical nature.

mixture :- Contains particles of two or more pure substances.

Homogenous :- Components mix uniformly throughout.

Heterogenous :- Composition is not uniform.

Elements :- Consists of only one type.

Compounds :- When two or more atoms of different elements combine together in a definite ratio, the molecules of compound is obtained.

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★ Physical and chemical properties of matter

- Physical Properties can be measured or observed without changing the identity or the composition of substance.
  e.g- colour, odour, melting point, boiling point, density, etc.

- Chemical Properties can be observed only through performing a chemical reaction.
  e.g. Composition, combustibility, reactivity, acidity or bisicity, etc.

★ Measurement of Physical Propertees

Earlier, two different system of measurement, i.e English system and the Matric system were used.

⇓

Developed in france, was based on decimal System.

The International System of Units [SI]

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- It was established by the 11th General Conference on weights and measures (CGPM from Conference General des Poids et Measures).

- The CGPM is an inter-governmental treaty organisation created by diplomatic treaty known as metre convection, which was signed in Paris in 1875.

- India has a National metrology Institute (NMI), which maintains standards of measurements
- This responsibility has been given to the National physical Laboratory (NPL), New delhi.

The SI System has seven base units.

	Base physical quantity	Symbol for quantity	Name of SI unit	Symbol for SI Unit
1	Length	l	metre	m
2	mass	m	Kilogram	kg
3	Time	t	Second	S
4	Electric current	I	Ampere	A
5	Thermo dynamics temperature	T	kelvin	K
6	Amout of Substance	n	mole	mol
7	Luminous Intensity	I_v	candela	cd

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Mass and weight

- Mass of a substance is the amount of matter present on it.
- mass is always constant.
- mass of a substance is determined by analytical balance.
- SI unit is kilogram (kg).

- weight is the force exerted by gravity on an object.

Volume

Volume is the amount of space occupied by a substance.
It has unit of [length]^3 or (m)^3

$1L = 1000 mL$

$1000 cm^3 = 1 dm^3$

Density

- Density of a substance i its amount of mass per unit volume.

- Density tells how closely the particles of a substance are packed.

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

Temperature

- SI unit of temperature is kelvin(k)
- There are three common sales to measure the temperature, °C (degree celcius). °F (degree Fahrenheit) k (kelvin):

From 0° to 100° b/w 32° to 212°

- Negative values are possible in celcius scale but not in kelvin

Relationship b/w Fahrenheit and celcius

$$\boxed{°F = \frac{9}{5}(°C) + 32}$$

Relationship b/w kelvin and celcius

$$\boxed{k = °C + 273.15}$$

Scientific Notation

Scientific notation is a method for expressing very large or very small number in a compact Form which is especially useful in chemistry when dealing with atomic masses or the number of molecules.

Any number can be written in the Form $N \times 10^n$, where:

- N is the digit term, a number that is between 1.000... and 9.999...
- n is the exponent which can be a positive or negative number. (integer)

Examples

- The number 232.508 can be written as 2.32508×10^2. Here the decimal was moved two places to the left, so the exponent is 2.

- The number 0.00016 can be written as 1.6×10^{-4}. Here the decimal was moved Four places to the Right, so the exponent is -4.

Calculations using Scientific Notation

Multiplication and Division

These operations follow the standard rules for exponents

- Multiplication: multiply the digit terms and add the exponents.
  Example: $(5.6 \times 10^5) \times (6.9 \times 10^8) = (5.6 \times 6.9) \times 10^{5+8} = 38.64 \times 10^{13}$, which should be written as 3.864×10^{14}.

- Division: Divide the digit terms and subtracts the exponents
  Example: $\frac{2.7 \times 10^{-3}}{5.5 \times 10^4} = (2.7 \div 5.5) \times 10^{-3-4} = 0.4909 \times 10^{-7}$, which should be written as 4.909×10^{-8}.

Addition and subtraction

For these operation, the exponent of the numbers must be same before you can add or subtract the digit terms.

- **Addition Example :-** To add 6.65×10^4 and 8.95×10^3 you First adjust one number to match the exponent of the other $(6.65 \times 10^4) + (0.895 \times 10^4)$. Then, Add the digit terms:

$(6.65 + 0.895) \times 10^4 = 7.545 \times 10^4$

- **Subtraction Example :-** To subtract 4.8×10^{-3} from 2.5×10^{-2} adjust the exponent: $(2.5 \times 10^{-2}) - (0.48 \times 10^{-2})$. Then subtracts the digit terms.

$(2.5 - 0.48) \times 10^{-2} = 2.02 \times 10^{-2}$.

Significant Figures

Significant Figures are the meaning-ful digits in measured or calculated value. They include all the digits known with certainity plus one Final digit that is estimated or uncertain. This is important because every experimental measurement has some level of uncertainity.

Rules for Identifying significant figures

1. Non-zero digits are always significant.

- Example:- 285 cm has 3 significant Figures; 0.25 mL has 2.

2. Zeros between non-zero digits are significant

- Example:- 2.005 has 4 significant figures.

3. zero to the left of the first non-zero digits are not significant.

- Example! 0.03 has 1 significant figure; 0.0052 has 2.

4. Zeros at the end of a number are significant if the number contains a decimal point.

- Example :- 0.200g has 3 significant figures.
- For a number like 100, it has only 1 significant figures. However, 100. has 3 and 100.0 has 4 using scientific notation.

(1×10^2, 1.0×10^2, or 1.00×10^2) is the best way to way to remove ambiguity.

5. Exact numbers such as those from counting objects (e.g. 2 balls) or from definations, have an infinite number of significant figures

6. For numbers in scientific notation all digits in the digit term (N) are significant

- Example

4.01×10^2 has 3 significant figures.

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Calculations with significant figures

Addition and subtraction

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

Example:- when adding 12.11, 18.0 and 1.012, the sum is 31.122 since 18.0 has only one digit after the decimal, the result must be rounded to 31.1

multiplication and devision

The result should be reported with the same number of significant figures as the measurement with the fewest significant figures.

Example:- 2.5 × 1.25 = 3.125 Because 2.5 has only two significant figures, the result is rounded to

3.1.

Rounding off Rules

1. If the digit to be removed is >5 the preceding number is increased by one. (e.g., 1.386 rounds to 1.39)

2. If the digit to be removed is <5, the preceding number is not changed (e.g., 4.334 rounds to 4.33)

3. If the digit to be removed is =5, the preceding number is increased by one if it is odd, and not changed if it is even.

- Example :- 6.35 is rounded to 6.4.
- Example :- 6.25 is rounded to 6.2

Accuracy Vs Precision

These two terms are often confused but are distinct concepts in measurement.

- **Accuracy** :- Refers to how close a measurement is to the true value.
- **Precision** :- Refers to how several measurement of the same quantity are to one another.

Example scenario (True value = 2.00g)

Student A measures 1.95g and 1.93g. These values are precise but not accurate.

Student B measures 1.94 g and 2.05 g. These are neither precise nor accurate.

Student C measures 2.01g and 1.99g. These are both accurate and precise.

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Dimensional Analysis

while calculating, there is a need to convert units from one System to the other. The method used to accomplish this is called Factor label method or unit Factor method or dimensional analysis.

Laws of Chemical Combinations

1 Law of conservation of Mass

- Given by Antoine Lavoisier in 1789
- It states that "matter can neither be created nor destroyed.
- mass of Reactant = Mass of product

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Ques 1.5g of C_2H_6 on complete combustion give 4.4g of CO_2 and 2.7g of H_2O. Show that it follows the law of conservation of mass.

Sol- molar mass of C_2H_6 =

$$2 \times 12 + 1 \times 6 = 24 + 6 = 30g.$$

$$n = \frac{1.5}{30} = \frac{1}{20} \text{ or } 0.05 \text{mole}$$

2 mole of C_2H_6 require 7 mole of O_2

0.05 mole require $\frac{7}{2} \times 0.05$ mole of O_2

$$\Rightarrow \frac{7}{2} \times \frac{1}{20} = \frac{7}{40} \text{ mole of } O_2$$

$$n = \frac{\text{mass}}{\text{molar mass}}$$

$$\frac{7}{40} = \frac{\text{mass}}{32} \Rightarrow \boxed{\text{mass} = 5.6\,g}$$

According to law of conservation of mass :- $1.5 + 5.6 = 4.4 + 2.7$

$7.1 = 7.1$ Hence proved.

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Law of Definite Proportion

- Given by Joseph Proust.
- He stated that "a given compound always contains exactly the same proportion of elements by weight whatever the source from where it is obtained."

e.g cupric carbonate

	% of Cu	% of C	% of O
Natural	51.35	9.74	38.91
Synthetic	51.35	9.74	38.91

Ques 0.243g of Mg on burning with oxygen yield 0.4030g of MgO. In another experiment, 0.182g of Mg burn with O_2 and give 0.302g MgO. Proven that this law follow law of definite composition.

Ans In experiment, $mg = 0.243\,g$

$mgO = 0.4030g$

$O_2 = 0.4030g - 0.243$

$= 0.160g$

$M : O \Rightarrow \frac{0.243}{0.160} = 1.51$ (approx)

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In exp. II, $Mg = 0.182\ g$

$MgO = 0.302\ g$

$O_2 = 0.302 - 0.182 = 0.120\ g.$

$$M:O \Rightarrow \frac{0.182}{0.120} = 1.51 \text{ (approx)}$$

Law of Multiple Proportion

- Given by John Dalton in 1803
- "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."
- Example:

$\Rightarrow H_2 + O_2 \rightarrow H_2O$
(2g) (16g) [18g]

$\Rightarrow H_2 + O_2 \rightarrow H_2O_2$
(2g) + (32g) (34g) (Hydrogen Peroxide)

Diff. masses of O (16g & 32g) which commbine with a fixed mass of hydrogen (2g) bear a simple ratio i.e 16:32 or 1:2

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Ques N_2O, NO, N_2O_3, NO_2, N_2O_5. Show that these compounds of same element Follow Law of multiple proportion.

Ans

N	O
28	16
14	16
28	48
14	32
28	80

we have to make no. of one element Fix

N	O
28	16
28	32
28	48
28	64
28	80

Diff mass of O (16, 32, 48, 64, 80) which combines with N (28g) bears a simple ratio.
1 : 2 : 3 : 4 : 5.

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4. Gay Lussac's Law of Gaseous Volumes

- Given by Gay lussac in 1808.
- When gases react they do so in a simple ratio by volume, provided all gases are at same temperature and pressure.

Example

$2H_2 + 1O_2 \rightarrow 2H_2O_2$
(2 Volume) (1 Volume) (2 Volume)
100 mL 50 mL 100 mL

$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$
(1 volume) (1 volume) (2 volume)
20 mL 20 mL 40 mL.

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Que When 100 mL of CH_4 burn it produces CO_2 and H_2O. Find volume of O_2 required and CO_2 produced.

Ans

$$CH_4(g) + 2O_2 \longrightarrow CO_2 + 2H_2O(l) + Heat$$

(1 volume) (2 volume) (1 vol) (2 vol)

100 mL 200 mL 100 ml

Ques 60 mL of $N_2(g)$ combine with 60 mL of $H_2(g)$. Find the composition of resultant mixture.

Ans

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

(1 volume) (3 volume) (2 volume)

20 mL 60 mL 40 mL

3 volume = 60 mL

1 volume = 20 mL

Ammonia (NH_3) = 40 mL.

Nitrogen - 20 mL

Unreactive Nitroge = 40 mL

5. Avogadro's Law

- Given by Amedo Avagadro in 1811.
- "Equal volumes of all gases at the same temperature and pressure should contain equal number of molecules.
- $\boxed{\text{Vol. of gas} \propto \text{No. of molecules}}$

Example :-

100 ml of O_2 $\rightarrow$ Y molecules

100 mL of CH_4 $\rightarrow$ Y molecules

100 mL of CO_2 $\rightarrow$ Y molecules

10 mL of H_2 $\rightarrow$ $\frac{Y}{10}$ molecules

50 mL of CO_2 $\rightarrow$ $\frac{Y}{2}$ molecules

Dalton's Atomic Theory

- In 1808, Dalton published 'A New system of chemical philosphy' in which he proposed :-

1. Matter consits of indivisible atom.
2. All atoms of a given element have identical properties including identical mass. Atoms of different element have different mass.
3. Compounds are formed when atoms of different elements combine in fixed ratio.
4. chemical reactions involve reorganisation of atoms. These are neither created or nor destroyed in a chemical reaction.

Atomic Mass

- Atomic mass is the mass of a atom.
- The present system of atomic mass is based on C-12 as standard since 1961

$$\boxed{1 \text{ atomic mass units} = \frac{1}{12} \text{th mass of C-12 atom}}$$

- At present 'amu' has replaced by 'u' which is known as Unified mass.

Average Atomic Mass

The Average atomic mass of an element is weighted average of the atomic masses of its isotopes.

$$\text{Average atomic mass} = \sum_{n=1}^{n} \frac{(\%\text{ abundance})_i \times (\text{mass}_i)}{100}.$$

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For Example

Abundance of carbon - 12 isotope → 99%

Abundance of carbon - 13 isotope → 1%

Average atomic mass of carbon = $\sum_{n=1}^{n}$

$$\frac{[(99)\times(12)]+[(1)\times(13)]}{100} = 12.01 \text{ amu}$$

Molecular Mass

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

Example:- molecular mass of methane

$(CH_4) = 12 \times 1 + 1 \times 4$
$= 12 + 4$
$= 16u$

Molecular mass of glucose

$(C_6H_{12}O_6) = 6 \times 12 + 12 \times 1 + 6 \times 16$
$= 72 + 12 + 96$
$= 180u.$

Formula Mass

- Formula mass of a substance is a sum of the atomic masses of all atoms in a formula unit of the compound.

- Same as molecular mass.

- The only difference is that we use word formula unit for those substances whose constituent particles are time.

e.g- NaCl
Formula unit mass =
$= 1 \times 23 + 1 \times 35.5$
$= 58.5$ u.

Mole Concepts of molar mass

- Mole is a number which is used to count microscopic particles
- mole is the SI unit of the substances (7th base quantity)
- 1 mole contains exactly $6.02214076 \times 10^{23}$ elementary entities. (Avagadro Constant) (N_a)
- molar mass in grams is numerically equal to atomic/ molecular/ Formala mass in u.

e.g. Molecular mass of water = 18 u
molar mass of water = 18 g.

$$\boxed{n = \frac{\text{Given mass}}{\text{Molar mass}}}$$

$$\boxed{n = \frac{\text{molecules / atoms Given}}{6.022 \times 10^{23}}}$$

$$\boxed{n = \frac{\text{volume Given (At STP)}}{22.4 \quad (22.4\ dm^3)}}$$

Que Find no. of moles present in

1) 11.2 of N_2 gas STP → $n = \frac{volume}{22.4} = \frac{11.2}{22.2} = 0.5$ mol

ii) 16g of CH_4 gas → $n = \frac{Mass}{molar\ mass} = \frac{16}{16} = 1$ mol

iii) 12×10^{24} molecules of O_2 gas

→ $n = \frac{no\ of\ particles}{6.22 \times 10^{23}} = \frac{12 \times 10^{24}}{6 \times 10^{23}} = 20$ moles

Percentage Composition

$\% \text{ composition} = \frac{\text{mass of that element in compound}}{\text{molar mass of ~~m~~ element}} \times 100$

Example

mass of hydrogen in water $= \frac{2}{\cancel{18}_9} \times 100 = 11.1\%$

Empirical formula for Molecular Formula

- An empirical formula represents the simplest whole number ratio of various atoms present in a compound.

- Molecular formula shows exact no. of atoms in a compound.

- Example

Eethane (C_2H_6) $\longrightarrow$ (C_1H_3)
(molecular Formula) (Empirical Formula)

$\boxed{\text{molecular Formula} = n \times \text{empirical Formula}}$

$\boxed{\text{mol. weight} = n \times \text{empirical weight}}$

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Ques A compound has molecular weight 90 and its empirical formula is CHO_2. Find molecular formula.

Sol^n

$$\text{mol. wt} = n \times \text{Emp. wt}$$

$$n = \frac{\text{Mol. wt}}{\text{Emp. wt}}$$

$$n = \frac{90}{45} = 2$$

Mol. Formula = $2 \times$ Emp. Formula
$= 2 \times CHO_2$
$= C_2H_2O_4$

Ques A Compound has vapor density 45 and emp. Formula CHO_2. Find molecular Formula.

Sol^n $\boxed{\text{Mol. wt} = 2 \times \text{Vapour Density}}$

$$\text{Mol. wt} = 2 \times 45 = 90$$

$$n = \frac{\text{mol. wt}}{\text{emp. wt}} = \frac{90}{45} = 2$$

Mol. Formula = $C_2H_2O_4$

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Ques Find the mol. Formula of a compound in which $C = 40\%$, $H = 6.7\%$, $O = 53.3\%$. and its molar mass is 60.

Soln

Element	%	At. wt	mole (At. ratio)	Simplest ratio
C	40	12	$\frac{40}{12} = 3.33$	1
H	6.7	1	6.7	2
O	53.3	16	$\frac{53}{16} = 3.33$	1

Empirical Form = $C_1 H_2 H_1$

$$n = \frac{\text{mol. wt}}{\text{emp. wt}} = \frac{60}{30} = 2$$

$$\text{mol. Formula} = 2 \times CH_2O$$
$$= C_2 H_4 O_2$$

Stoichiometry

Stoichiometry :- Deals with calculation of masses of reactant and products.

$① CH_4(g) + ② O_2 g \rightarrow ① CO_2(g) + ② H_2O(g)$

Stoichiometric coefficient
(represent moles)

Ques If 16g of Methane is burnt. Find ① mass of CO_2 produced?

Solⁿ $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

$n(\text{methane}) = \frac{\text{mass}}{\text{molar mass}} = \frac{16}{16}$ - Moles

$\Rightarrow$ 1 mole of CH_3 produced 1 mole of CO_2

$n(\text{carbondioxide}) = \frac{\text{mass}}{\text{molar mass}}$

$1 = \frac{x}{44}$

$\boxed{x = 44g}$

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Ques Find the volume of CO_2 produced when 50g $CaCO_3$ is heated?

$1\,CaCO_3 \longrightarrow 1\,CaO + 1\,CO_2$

1 mole $CaCO_3$ → 1 mole CO_2

Solve $n(CaCO_3) = \frac{\cancel{50}}{\cancel{100}_{2}}$ = 0.5 mole $CaCO_3$ → $\underline{0.5\text{ Mole } CO_2}$

$= \underline{0.5\text{ Moles}}$

$\boxed{n = \frac{\text{Volume}}{22.4}}$

$0.5 \times 22.4 = \text{vol}$

vol. = 11.2 L

Que $2C_4H_{10} + 13O_2 \longrightarrow 8CO_2 + 10H_2O$

Find mass of CO_2 produced, when 5.8 g Butan is burnt.

Soln $n(\text{Butane}) = \frac{\cancel{5.8}}{\cancel{580}\ 10} = 0.1$ moles

2 moles of C_4H_{10} → 8 mole of CO_2

0.1 mole of C_4H_{10} → $\frac{\cancel{8}^{4}}{\cancel{2}} \times \frac{0.1}{\cancel{10}_{5}}$ mole of CO_2

⇒ 0.4 mole of CO_2

$n(CO_2) = \frac{\text{Given Mass}}{\text{Molar Mass}}$

$0.4 \times 44 = \text{mass}$

$\boxed{\text{mass} = 20.6\,g}$

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Ques $CuCO_3 \xrightarrow{\Delta} CuO + CO_2$

Find mass of $CuCO_3$ required to produce 11.2 L of CO_2 STP.

Solve $n = \frac{11.2}{22.4} = 0.5$ moles

1 mole of $CuCO_3 \rightarrow$ 1 mol. CO_2

0.5 mole " $\rightarrow$ 0.5 mol. CO_2

$n = (CO_2) = \frac{mass}{molar\ mass}$

$= 0.5 = \frac{mass}{124}$

$\boxed{mass = 62\ g}$

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Ques $2Ca(NO_3)_2 \xrightarrow{\Delta} 2CaO + 4NO_2 + O_2$

Find mass of CaO produced when 5.6 L of NO_2 is produced (STP)?

$$n(NO_2) = \frac{\text{Volume}}{22.4L\,(STP)} = \frac{5.6}{22.4} = 0.25 \text{ moles}$$

4 mol. of $NO_2 \longrightarrow$ 2 Mole of CaO

0.25 Mol. $NO_2 \rightarrow \frac{\cancel{2}}{\cancel{4}_2} \times 0.25$ Mole of CaO

$$= \frac{1}{8} \text{ Moles}$$

$$n(CaO) = \frac{\text{mass}}{\text{molar mass}}$$

$$\frac{1}{\cancel{8}} \times \cancel{56}^{7} = \text{mass}$$

$$\Rightarrow \boxed{\text{mass} = 7g}$$

Reactions in Solutions

The concentration of solution can be expressed of the following ways:—

1. weight by weight percentage (w/w%)
2. volume by volume percentage (v/v%)
3. weight by volume percentage (w/v%)
4. PPM (Parts Per million)
5. PPB (Part per Billion)
6. Mole Fraction
7. Molarity
8. Molality
9. Normality

1. weight by weight percentage [w/w%]

e.g. A solution of NaCl in water is 5% (w/w%)

it means, 5g NaCl is present in 100g of water

$$\boxed{(w/w\%) = \frac{\text{mass of solute}}{\text{mass of sol}^n} \times 100}$$

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Ques 10g sucrose is dissolved in 100g of water. Find weight by weight %.

Solⁿ Solute = 10g sucrose

solⁿ = 100+10 = 10 = 110g solⁿ

$$[w/w\%] = \frac{\text{Mass of Solute}}{\text{Mass of Sol}^n} \times 100$$

$$= \frac{1\cancel{0}}{11\cancel{0}} \times 100 = \frac{100}{11} = 9.09\%$$

Que 10g of NaCl is present 100mL of Solution. Find (w/w%). If density of solⁿ is 1.2 g/m.

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

$1.2 \times 100 = \text{mass}$

$\Rightarrow \text{mass} = 120\,g$

$$(w/w\%) = \frac{\text{mass of solute}}{\text{mass of sol}^n} \times 100$$

$$= \frac{1\cancel{0}}{10\cancel{0}} \times 100$$

$$= \frac{100}{12} = 8.33\%$$

Ques. A solution prepared by adding 2g of a substance A to 18g of water. calculate mass percent of solute.

Solⁿ mass of solute (A) = 2g
mass of solution = 2g + 18g = 20g

$$\text{mass percent of A} = \frac{\text{mass of A}}{\text{mass of Sol}^n} \times 100$$

$$\Rightarrow \frac{2}{\cancel{20}_{10}} \times 100 = 10\%$$

2. Volume by Volume Percentage (V/V%)

e.g A HNO_3 (eq) Solⁿ is 7% (V/V%)
7 mL of HNO_3 is present in 100 mL of Solⁿ

$$\boxed{(V/V\%) = \frac{\text{volume of solute}}{\text{volume of sol}^n} \times 100}$$

3. weight by volume Percentage (W/V%)

A sugar solution is 3% (W/V%)

⇒ 3g of sugar is present in 100 mL of solⁿ

$$\boxed{[W/V\%] = \frac{\text{mass of solute}}{\text{volu. of sol}^n} \times 100}$$

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Ques A sugar solution is 10% (w/v%) Find (w/w) if density of solution is 1.25 g mL.

Soln 10g sugar is present in 100 mL soln

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

$$1.25 \times 100 = Mass$$

$$\Rightarrow \boxed{Mass = 125g}$$

$(w/w\%) = $

$$\frac{\text{mass of solute}}{\text{mass of soln}} \times 100$$

$$= \frac{\cancel{10}^{2}}{\cancel{125}} \times 100$$

$$\underline{8\%}$$

Ques A sugar solution is 10% (w/w) find (w/v%) if density is 1.2 g/mL

Soln 10g sugar is present in 100g soln

$$Density = \frac{mass}{volume}$$

$$1.2 \times 10 = \frac{}{Volume}$$

$$\cancel{1.2}$$

$$Volume = 8.33\% \text{ or } \frac{100}{12}$$

$(w/v\%) = $

$$\frac{\text{mass of solute}}{\text{vol. of soln}} \times 100$$

$$= \frac{10}{\frac{100}{12}} \times 100$$

$$= 10 \times 12$$

$$= 120\%$$

Ques A solution of HNO_3 is 5% (w/w). Find the mass of HNO_3 present in 100 mL of solution (density of $soln = 1.4 g/mL$)

Soln $5g\ HNO_3 \longrightarrow 100g$ solution

Density = $\frac{mass}{volume}$

$1.4 \times 100 = mass$

$\Rightarrow mass = 140$.

$(w/w) = \frac{\text{mass of solute}}{\text{mass of soln}} \times 100$

$5 = \frac{x}{140} \times 100$

$\boxed{x = 7g}$

4. **PPM (Part Per Million)** = 10^6 Part

$$\boxed{\text{no. of PPM} = \frac{\text{Mass of solute}}{\text{Mass of soln}} \times 10^6}$$

Ques O_2 is dissolved in water as 8×10^{-4} g of O_2 is 100g of water. Find concentration of O_2 is PPM.

Soln no. of PPM = $\frac{\text{Mass of solute}}{\text{Mass of soln}} \times 10^6$

$= \frac{8 \times 10^{-4}}{100} \times 10^6$

$= \frac{8 \times \cancel{10^2}}{\cancel{100}}$

$= 8$ PPM

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Ques Sea water contains 53 PPM of Na_2CO_3. Find mass of Na_2CO_3 present in a glass of water (250 mL)

Sol no. of PPM = $\frac{\text{Mass of solute}}{\text{Mass of sol}^n} \times 10^6$

$53 = \frac{x}{250} \times 10^6$

$\frac{53 \times 25}{10^5} = x$

$\Rightarrow x = \frac{53 \times \cancel{25}}{1000 \times \cancel{100}_{4}} = \frac{53}{4 \times 10^3} = 13.25 \times 10^{-3}$

5. P.PB (Part Per Billion) = 10^9 part.

NaCl → 5 PPB

⇒ 5g NaCl is present in 10^9g of Sol^n

$$\boxed{\text{no. of PPB} = \frac{\text{Mass of solute}}{\text{Mass of sol}^n} \times 10^9}$$

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6. Mole Fraction

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

$A \rightarrow 1$ mole $\Rightarrow$ Mole fraction $\Rightarrow X_A$ (zeta) $\quad \eta_A$ (eta)

$A \rightarrow n_A$ mol

$B \rightarrow n_B$ mol

$C \rightarrow n_C$ mol

$$\boxed{X_C = \frac{n_C}{n_A + n_B + n_C}} \qquad \boxed{X_A + X_B + X_C = 1}$$

Ques It is given that NaOH is 40g and H_2O is 54g. Find X_{NaOH} and X_{H_2O}.

Sol

$H_2O = 2 \times 1 + 16 = 18$ g

$n_{H_2O} = \frac{54}{18} = 3$ Moles

$NaOH = 23 + 16 + 1 = 40$ g

$n_{NaOH} = \frac{\text{mass}}{\text{molar mass}} = \frac{40}{40} = 1$ mole.

$$\boxed{X_{NaOH} = \frac{1}{4}} \qquad \boxed{X_{H_2O} = \frac{3}{4}}$$

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Ques A solution has 46% (w/w%) Ethyle alcohol (C_2H_5OH) in water. Find mole fraction of ethyle alcohol.

$C_2H_5OH = 46\,g$

Sol

Ethyle Alcohol = 46 g

water = 100 − 46 = 54 g.

$$n_{C_2H_5OH} = \frac{46}{46} = 1 \text{ mole}$$

$$n_{water} = \frac{54}{18} = 3 \text{ moles}$$

$$X_{C_2H_5OH} = \frac{1}{4}$$

7. Molarity

molarity (M) – Number of Moles of Solute present in 1 litre (L) of Solution.

1 m of NaOH – 1 Mol of NaOH is present in 1 litre of Solⁿ

$$\boxed{\text{Molarity (M)} = \frac{\text{No. of moles of Solute}}{\text{Vol. of Solution in litres}}}$$

marality depend on temp.

molar Solution – M = 1 (1M)

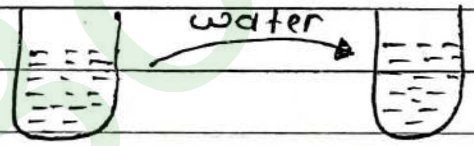
Bi Molar – M = 2

Semi Molar – M = 1/2

Deci Molar – M = 10/10

Centi Molar – M = 1/100

Molarity of Dilution



M_1, V_1 $\qquad$ M_2, V_2

$m_1 = \frac{n}{V_1}$ $\qquad$ $M_2 = \frac{n}{V_2}$

$n = M_1 \times V_1$ —(I) $\qquad$ $n = m_2 \times V_2$ —(II)

From (I) and (II) we get

$$\boxed{M_1 \times V_1 = M_2 V_2}$$

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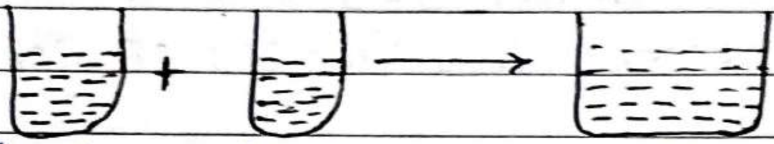
Molarity of mixture



M_1, V_1 M_2, V_2 $M, (V_1+V_2)$

$$\boxed{M = \frac{M_1V_1 + M_2V_2}{V_1+V_2}}$$

$$M_1V_1 + M_2V_2 = M(V_1+V_2)$$

Ques Find molarity of :-

(i) 40g NaOH dissolved in 250 mL solution.
$\to 40$

Solⁿ $n = \frac{\text{Given mass}}{\text{Molar mass}}$

$$n = \frac{\cancel{40}\,g}{\cancel{40}\,g}$$

$$\boxed{n = 1\text{ mole}}$$

$$m = \frac{n\text{ (in solute)}}{V\text{ (in L)}}$$

$$\Rightarrow \frac{1}{\frac{\cancel{250}}{\cancel{1000}_4}} = \underline{\underline{4\,M}}$$

Ques 4.9 of H_2SO_4 present in 500 cm^2 solution.

$1L = 1000 mL$
$1L = 1000 cm^3$

Soln $n = \frac{\text{Given mass}}{\text{Molar mass}}$

$H_2SO_4 \rightarrow 2\times1 + 1\times23 + 16\times4$
$\rightarrow 2 + 32 + 64$
$\Rightarrow 98g$

$n = \frac{\cancel{4.9}}{\cancel{980}_{20}}$

$\boxed{n = 1/20 \text{ moles}}$

$M = \frac{\frac{1}{\cancel{20}\ 10}}{\frac{\cancel{500}}{\cancel{1000}\ 2}} = \frac{1}{10} = 0.1 M$

(iii) mass of Na_2CO_3 present in 100 mL of 3m Solution. (Na = 23, C = 12, O = 16)

Soln

$m = \frac{n}{V(\text{in} L)}$

$3 = \frac{n}{\frac{100}{1000}}$

$n = 0.3 \text{ moles}$

$n = \frac{\text{Given mass}}{\text{molar mass}}$

$0.3 = \frac{x}{106}$

$\boxed{x = 31.8g}$

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(iv) 10% (w/w%) (aq) Solution of H_2SO_4. If density of Sol^n is 1.1 g/mL.

Solve: 10g of H_2SO_4 is present in 100g of solution

$(w/w\%) = 4$

$$M = \frac{n}{V(\text{in L})} = \frac{\frac{\text{Given mass}}{\text{molar mass}}}{\text{Vol (in L)}}$$

$$M = \frac{\frac{10}{98}}{\frac{100}{1.1\times1000}} (L) = \frac{10\times11}{98} = \frac{55}{49}$$

$$\frac{\text{Mass}}{\text{Density}} = \text{Vol.}$$

$$\frac{100}{1.1} = \text{volume}$$

$$\boxed{\text{Volume} = \frac{100}{1.1}}$$

(V) If 120g Urea (NH_2CONH_2) is dissolved in 1000g water Find molarity of the solution. If density of Solution is 1.12 g/mL (N = 14, H = 1, C = 12, O = 16)

Solar: Molar mass of NH_2CONH_2 = 60g.

$$n = \frac{\text{Given mass}}{\text{molar mass}}$$

$$\frac{\cancel{120}^{2} g}{\cancel{60} g}$$

$$\boxed{n = 2 \text{ moles}}$$

$$M = \frac{\text{no. of mass in solute}}{\text{Vol. of } sol^n (\text{in L})}$$

$$= \frac{20}{1}$$

$$\boxed{M = 20M}$$

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

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

$$= \frac{\cancel{1120}^{10} \times 100}{\cancel{1.12}}$$

$$\boxed{\text{Vol.} = 1000 \text{ mL} = 1L}$$

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(VI) Find molarity of water in density in 1g/mL.

Soln Let us consider 100 mL of water is present

$$\text{mass of water} = \frac{1000}{100\,g} \quad 1 \times 100\ mL.$$

$$M = \frac{\frac{\text{Given mass}}{\text{molar mass}}}{\text{vol (mL)}} = \frac{\frac{\cancel{100}}{18}}{\frac{\cancel{100}}{1000}} = \frac{1000}{18} = 55.555M$$

Ques Find the mass of KOH present in 150 cm^3 of semi molar aq. solution, (K = 19)

Soln

$$m = \frac{\text{no. of moles in solute}}{\text{Vol. of Solution in Litre}}$$

$$150\,cm^3 = \frac{150}{1000} = 3/20\,L$$

$$\frac{1}{2} = \frac{n}{3/20}$$

$$\Rightarrow \boxed{n = \frac{3}{40}}$$

$$n = \frac{\text{Mass}}{\text{molar mass}}$$

$$\frac{3}{40} = \frac{\text{Mass}}{36}$$

$$\text{Mass} = \frac{3}{40} \times 36$$

$$\boxed{\text{mass} = 2.7g}$$

Molarity of Ions

Que If 53g Na_2CO_3 is dissolved in 500mL Solution Find malarity of (i) Na_2CO_3 (ii) Na^+ (iii) CO_3^{2-}
(Na=23)

Solve $500\,mL = \frac{1}{2}$ Litre

$n = \frac{Mass}{Molar\,mass}$

$n = \frac{53}{106}$

(i) $M = \frac{\frac{53}{106}}{\frac{1}{2}} = \frac{\frac{1}{2}}{\frac{1}{2}} = 1M$

(ii) $Na_2CO_3 \rightleftharpoons Na_2^+ + CO_3^2$
(1 mole) (2 mole) (1 mole)

$Na^+ = 2M$

(iii) $CO_3^{2-} = 1M$

molarity of Devision.

Ques 0.2 M, 100 ML of HNO_3 Solution is diluted with 100 mL water. Find new Molarity.

Soln $M_1 = 0.2$, $V_1 = 100\,mL$
$M_2 = x$, $V_2 = 200\,mL$

$M_1 \times V_1 = M_2 \times V_2$

$\frac{0.2 \times 100}{200} = M_2$

$\boxed{M_2 = 0.1M}$

Ques If 0.2 M, 100 ML of HCl is Mixed with 0.1 m, 300 ML HCl solution, Find molarity of mixture.

Sol

$M_1 = 0.2$, $V_1 = 100$ mL

$M_2 = 0.1$ V_2 300 mL

$$M = \frac{M_1 \times V_1 + M_2 \times V_2}{V_1 + V_2}$$

$$M = \frac{0.2 \times 100 + 0.1 \times 300}{100 + 300}$$

$$= \frac{20 + 30}{400}$$

$$\frac{\cancel{50}}{\cancel{400}_{8}}$$

$$\boxed{m = 1/8}$$

8. Molality

- No of mole of Solute present in 1kg of Solvent

Denoted by 'm'

Does not change with temperature

molality of HCl is 3m ⇒ 3 moles of HCl is present in 1 kg water (solvent)

$$\boxed{m = \frac{\text{no. of moles of solute}}{\text{mass of solvent in 1 kg}}}$$

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Que Find molality of 20g of NaOH dissolved in 100g solution
(NaOH → 40g)

Soln $m = \dfrac{n}{\text{mass of solvent (in kg)}} = \dfrac{\frac{20}{40}}{\frac{80}{1000}} = \dfrac{25}{4} = 6.25\,m$

Equivalent weight and Gram

The reactants have always same number of Gram Equivalent
- No of moles of Reactant can be different but no. of Gram equivalent will always same.

$$\boxed{\text{Eq. mass} = \frac{\text{Mol. wt}}{x\ (\text{valency Factor})}}$$

* For acids, x = Basicity (No of H^+ ions released)
For Bases, x = Acidity (No. of OH^- ions released)
For salts, x = Total positive charge of cation.

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Que Find eq. wt of (1) H_2SO_4 (2) NaOH and (3) Na_2CO_3

Soln 1. $Eq \cdot wt = \frac{Mol \cdot wt}{x} = \frac{98}{2} = 49g$

2. $Eq \cdot wt = \frac{Mol \cdot wt}{x} = \frac{40}{1} = 40g$

3. $Eq \cdot wt = \frac{Mol \cdot wt}{x} = \frac{106}{2} = 53g$

⇒ Equivalent weight depends on reaction

$H_2SO_4 + NaOH \longrightarrow NaHSO_4 + H_2O$
$x = 1$

$$\boxed{\text{No. of gm equivalent} = \frac{\text{Mass}}{\text{Eq} \cdot \text{wt}}}$$

$$\boxed{\text{No. of gm eq.} = \text{no. of moles} \times x}$$

Ques Find no. of gm eq. present in

(i) 0.4 g of NaOH

$$\text{No. of gm eq} = \frac{\text{Mass}}{\text{Eq.wt}}$$

$$= \frac{\text{Mass}}{\frac{\text{Mol.wt}}{x}}$$

$$= \frac{0.4}{\frac{40}{1}}$$

$$= 0.01g$$

(ii) 0.98 g of H_2SO_4

$$\text{Eq.wt} = \frac{\text{Mol.wt}}{x}$$

$$\Rightarrow \frac{98}{2} = 49g$$

No. of gm eq =

$$\frac{\text{mass}}{\text{Eq.wt}} = \frac{0.9}{49}$$

$$= 0.02 g$$

Ques Find no of gm eq present in 3 moles of H_2SO_4

Soln

$$\text{no. of gm eq} = \text{no of moles} \times X$$

$$= 3 \times 2$$

$$= 6g\ eq.$$

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9. Normality (N)

No. of gm. eq. of solute which are present in 1 L solⁿ

$$\boxed{N = \frac{\text{no. of gm. eq. of solute}}{\text{vol. of sol}^n \text{ (in L)}}}$$

Relationship b/w Normality & molarity

$$N = \frac{\text{No. of gm. eq. solute}}{\text{Vol. of sol}^n \text{ (in L)}} \quad \text{(I)} \qquad M = \frac{\text{No. of moles of solute}}{\text{Vol. of sol}^n \text{ (in L)}} \quad \text{(II)}$$

Divide (I) and (II), we get

$$\frac{N}{M} = \frac{\text{no. of gm. eq solute}}{\text{no of moles of solute}} = \frac{\frac{\text{Mass}}{\text{Eq. wt}}}{\frac{\text{Mass}}{\text{Mol. wt}}} = \frac{\text{Mol. wt}}{\text{Eq. wt}}$$

$$\frac{N}{M} = \frac{\text{mol. wt}}{\frac{\text{mol. wt}}{x}} = x$$

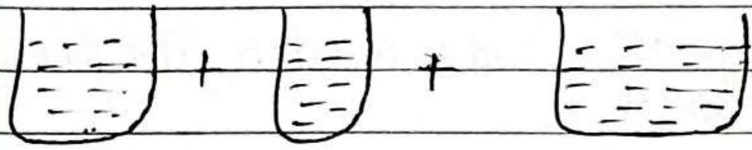
$$\Rightarrow \frac{N}{M} = x$$

$$\boxed{N = M \times x}$$

Normality of Mixtures
(Acid + Acid / Base + Base)

$$\boxed{N = \frac{N_1V_1 + N_2V_2}{V_1 + V_2}}$$

Acid + Base



N_a, V_a $\quad$ N_B, V_b $\quad$ N

$V_a + V_b$

IF $N_aV_a > N_bV_b \longrightarrow$ Acidic solⁿ

$N_aV_b < N_aV_a \longrightarrow$ Basic solⁿ

$N_aV_a = N_aV_b \longrightarrow$ Neutral solⁿ

- Acid solⁿ $\boxed{N = \dfrac{N_aV_a - N_bV_b}{V_a + b_b}}$

- Basic solⁿ $\boxed{N = \dfrac{N_bV_b - N_aV_b}{V_a + V_b}}$

Neutral solⁿ $\boxed{N_aV_a = N_bV_b}$