UNIT 01 | CHAPTER 1.1 – 2.8

General & Physical Chemistry

* Section "A" *

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* Language of Chemistry *

Chemical Arithmetic:-

Dalton's Atomic Theory & Laws of Stoichiometry:-

★ postulates of Dalton's atomic theory are:-

1) All matter consists of extremely small particles called atom.
2) Atom of an element are identical in all respect.
3) Atoms of different element are entirely different and have different properties.
4) Atom can neither be created nor be destroyed by any chemical reaction.
5) Atoms combine with eachother in whole number ratio in a compound.

★ Law of Stoichiometry:-

↳ There are five basic laws of Chemical Combination, they are:-

1) Law of Conservation of mass.
2) Law of definite proportions. [law of Constant Composition]
3) Law of multiple proportions
4) law of reciprocal proportions [The law of Equivalent proportions]
5) Gay. Lussac's law of Gaseous volume.

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[I]. Law of Conservation of mass:-

↳ It States, "In any chemical reaction the total mass of reactant is equal to the total mass of the products." Matter is neither Created nor destroyed. This law Can be illustrated by taking reaction between Sodium Chloride (NaCl) and Silver Nitrate ($AgNO_3$).

Let w_1 gram of NaCL solution react with w_2 gram of $AgNO_3$ to give w_3 gram of white precipiate (ppt.) of AgCL (Silver chloride) and w_4 gram of $NaNO_3$ (Sodium Nitrate).

$$\underset{w_1}{NaCl} + \underset{w_2}{AgNO_3} \longrightarrow \underset{w_3}{AgCl} + \underset{w_4}{NaNO_3}$$

According to law of Conservation of mass,
Total mass of reactants = Total mass of products
i.e. $\boxed{w_1 + w_2 = w_3 + w_4}$

This law is in accordance with dalton's atomic theory. In that, atom can neither be Created nor be destroyed. Therefore, this law is also called law of indestructibility because total mass and energy remains conserved.

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[2]. Law of definite proportion [Law of Constant Composition].

↳ This law was given by Louis Proust in 1799. It states "A pure chemical compound is always composed of similar elements combined together in a fixed ratio by weight irrespective of their sources and method of preparation."

This law can be illustrated by taking water as example,

Water is obtained from various sources like, rain, river, tap, sea etc. Whatever may be the source of water, it always contains hydrogen & oxygen elements combined in fixed ratio of 1:8 by weight.

If $\frac{W_1}{W_2}$, $\frac{W_3}{W_4}$ & $\frac{W_5}{W_6}$ are the weight ratio of elements in a compound obtained by three different sources or methods then according to this law,

$$\boxed{\frac{W_1}{W_2} = \frac{W_3}{W_4} = \frac{W_5}{W_6}}$$

★ Example,

↳ In an experiment 5.4 gm of Cupric oxide [CuO] was obtained by treating 4.32 gm of Cupper with concentrated (conc.) Nitric acid [Conc. HNO_3] and subsequent ignition while in another experiment 2.3 gm of CuO on reduction gave 1.84 gm of Cupper. Show that these data illustrate the law of Constant Composition.

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Soln:- In 1st experiment:-

$$Cu + HNO_3 \longrightarrow Cu(NO_3)_2 \xrightarrow{\Delta} \underset{5.4\,gm}{CuO}$$

Wt. of CuO = 5.4 gm
Wt. of Cu = 4.32 gm
Wt. of O = (5.4 - 4.32) = 1.08 gm

$\therefore$ Ratio of weight of Cu:O in CuO = $\frac{4.32}{1.08}$ = 4:1

In another experiment;-

$$\underset{2.3\,gm}{CuO} + H_2 \longrightarrow \underset{1.84\,gm}{Cu} + H_2O$$

Wt. of CuO = 2.3 gm
Wt. of Cu = 1.84 gm
Wt. of O = (2.3 - 1.84) = 0.46 gm

$\therefore$ Ratio of wt. of Cu:O in CuO = $\frac{1.84}{0.46}$ = 4:1

Here, in both experiment the ratio weight of Cu & O in CuO compound is same i.e. 4:1. Which illustrate the law of definite proportion/constant proportion.

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[3] Law of Multiple proportions:-

↳ This law was proposed by John Dalton in 1804. It states, "When two element combine to give two or more compounds, the weight of one element which combined with fixed weight of another element bears a simple whole number ratio.

$$H + O \rightarrow \begin{cases} H_2O & 2, 16 \\ H_2O_2 & 2, 32 \end{cases}$$

Illustration:-

Let Hydrogen combine with oxygen element to give hydrogen oxide (H_2O) & Hydrogen peroxide (H_2O_2).

In H_2O two parts of hydrogen and 16 part of oxygen by weight combine to form it.

In H_2O_2 two parts of hydrogen and 32 parts of oxygen by weight combine to form it.
The ratio of weight of oxygen in H_2O & H_2O_2 which combine with 1 part of hydrogen is 8:16 i.e. 1:2 which is simple whole number ratio. Hence this illustrate the law of multiple proportion.

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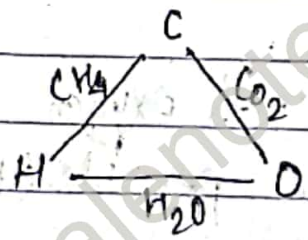
[4] **Law of reciprocal proportion [Law of equivalent proportion]:-**

↳ It was given by Richter in 1972. It states, "The weight ratio of two different elements which combine with the fixed weight of third element in two different compounds is either same or simple multiple of the ratio in which they combine with each other."

This law can be illustrated by taking Hydrogen, Oxygen and Carbon element which combine with eachother to produce H_2O, CO_2 and CH_4 respectively.



In CH_4, 12 gram of Carbon combines with 4 gram of hydrogen. In CO_2, 12 gram of Carbon combines with 32 gram of oxygen.

The ratio of weight of hydrogen & oxygen in methane and CO_2 which combine with 12 gram of carbon is 1:8 —(i)

Again,
Hydrogen combines with oxygen to give water in H_2O two gram of Hydrogen combines with 16 gram of oxygen.
The ration of weight of H_2 & O in H_2O is 1:8 —— (ii)

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The ratio (i) and ratio (ii) are same which illustrate the law of reciprocal proportion.

Since the ratio are equal. this law is also called law of equivalent proportion.

[5] law of gaseous volume [Gay-Lussac's law]:-

↳ This law was given by Gay-Lussac in 1805. This law gives the relationship between volume of reactant and product.

It states, "When two gases combine together to give gaseous product. The volume of reactant and product bear a simple whole number ratio."

Example: $H_2 + 3N_2 \longrightarrow 2NH_3$

1 vol^m 3 vol^m $\longrightarrow$ 2 vol^m

Ratio of volume is 1:3:2.

Atomic Mass & Molecular mass:-

↳ The atomic mass of an element is the number of parts by weight of that element which shows how many times the element is heavier than $\frac{1}{2}$th weight of 1 atom of C^{12} isotopes. It is denoted by A and its unit is amu (atomic mass unit).

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i.e, $A = \dfrac{\text{Wt. of an element}}{\left(\frac{1}{12}\right)^{th} \text{ part of } C^{12} \text{ isotopes}}$

or, $A = \dfrac{\text{Wt. of an element}}{1 \text{ amu}}$

OR,

$A = \dfrac{12 \times \text{Wt. of an element}}{\text{Wt. of } C^{12} \text{ isotopes}}$.

★ Atomic mass Unit:-

↳ $\frac{1}{12}$, wt. of one atom of C^{12} isotopes is called atomic mass unit.

$1 \text{ amu} = \left(\frac{1}{12}\right)^{th}$ mass of one atom of C^{12} isotopes.

Calculate the value of 1 amu in gram?

↳ We know,

~~gm~~ ~~of~~ gram-atom weight = gram atomic weight

or, 1 gm atom of Carbon = 12 gm of Carbon

or, Mass of 6.023×10^{23} atom of Carbon = 12 gm of Carbon

or, Mass of 1 atom of Carbon = $\dfrac{12}{6.023 \times 10^{23}}$ gm of Carbon,

or, $\frac{1}{12}$ mass of 1 atom of Carbon = $\dfrac{12}{6.023 \times 10^{23}} \times \dfrac{1}{12}$ gm of Carbon

$\Rightarrow$ $\boxed{1 \text{ amu} = 1.66 \times 10^{-24} \text{ gm}}$

Atomic weight of Some element:-

SN	Element	Atomic weight.
1.	H	1.008 amu
2.	He	4.003 amu
3	Li	6.94 amu
4	Be	9.012 amu
5	B	10.81 amu
6	C	12.01 amu
7.	N	14.006 amu
8	O	15.99 amu
9	F	18.99 amu
10	Ne	20.17 amu
11	Na	22.98 amu
12	Mg	24.3 amu
13	AL	26.98 amu
14	Si	28.08 amu
15	P	30.97 amu
16	S	32.06 amu
17	Cl	35.45 amu
18	Ar	39.94 amu
19	K	39.098 amu
20	Ca	40.08 amu

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Fractional Atomic Weight of an element:-

↳ The same element has different isotopes. The percentage abundance of each isotopes in nature is different; Atomic weight is not the weight of single isotope. It is calculated by taking average weight of different isotopes of an element according to their percentage abundance in nature. Therefore atomic mass of an element is found in fractional number.

For example,

Chlorine exists as two isotopic forms Cl^{35} (75%) and Cl^{37} (25%)

$$\text{Atomic weight of Chlorine} = \frac{35\times75 + 37\times25}{75+25}$$
$$= 35.5 \text{ grams.}$$

Mass number is sum of number of protons and neutrons present in isotopes of element. Therefore, mass number is always in whole number.

For example,

Cl^{35} isotopes contains 17 protons & 18 neutrons

$$\text{Mass number of } Cl^{35} \text{ isotopes} = 17 + 18$$
$$= 35$$

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Molecular weight of Substance:-

↳ Number which shows how many times 1 molecules of a substance is heavier than $(\frac{1}{12})^{th}$ of mass of one atom of C^{12} isotopes is known as molecular weight of substance.

i.e.

$$\text{Molecular mass} = \frac{\text{Wt. of 1 molecules of given Compound}}{1 \text{ amu.}}$$

Calculation of molecules weight:-

Molecular wt. of H_2 = no. of atoms × atomic wt.

= 2 × 1.008

= 2.016 amu.

Molecular wt. of $H_2O = (2 \times 1) + (1 \times 16) = 18$ amu.

Molecular wt. of $CO_2 = (1 \times 12) + (2 \times 16) = 44$ amu

Molecular wt. of $NaOH = (1 \times 23) + (1 \times 16) + (1 \times 1) = 40$ amu

Molecular wt. of $CaCO_3 = (1 \times 40) + (1 \times 12) + (3 \times 16) = 100$ amu

Molecular wt. of $H_2SO_4 = (2 \times 1) + (32) + (16 \times 4) = 98$ amu

Molecular wt. of $Na_2CO_3 = (2 \times 23) + (12) + (3 \times 16) = 106$ amu

Mole Concept:-

↳ Amount of substance which contain fixed number of particles equal to Avogadro's number (i.e, 6.023×10^{23}) is called one mole of that substance. Mole can be expressed in terms of mass & volume.

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Expression of mole in terms of mass

↳ Mole is also defined as the expression of atomic or molecular mass in gram.

Atomic mass in term of gram (gram-atom)

↳ Atomic weight expressed in gram is called gram-atom or gram atomic weight.

$\boxed{\text{1 mole of atom = gram-atom = gram atomic weight = Avogadro's no.}}$

e.g.,

1 mole of hydrogen = 1 gram-atom = 1.008 gm = 6.023×10^{23} atoms

Molecular wt. in terms of gram (gram-molecules)

↳ Molecular wt. expressed in gram is called gram-molecules or gram molecular weight.

$\boxed{\text{1 mole of molecule = 1 gram-molecule = gram molecular wt. = } N_A}$

e.g.:

1 mole of CO_2 = 1 gm-molecule = 44 gram = 6.023×10^{23} CO_2 molecules

Expression of mole in terms of Volume

↳ Mole is defined as the volume occupied by 22.4 litres of gas at NTP. It is called molar volume.

For e.g.:

1 mole of O_2 = 22.4 Litre at NTP.

We can write:-

1 mole of NH_3 = 6.023×10^{23} NH_3 molecules = 22.4 litres of NH_3 at NTP.

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* Calculate the number of mole of 4 gm of Carbon.

↳ Soln: We know that,

12 gm of C = 1 mole

1 gm of C = $\frac{1}{12}$ mole

4 gm of C = $\frac{4}{12}$ mole

= 0.333 mole. #

* Calculate the number of mole of 2 atom of sulphur.

↳ We have,

6.023×10^{23} atom = 1 mole

1 atom = $\frac{1}{6.023 \times 10^{23}}$ mole.

2 atom = $\frac{2}{6.023 \times 10^{23}}$ mole

= 3.32×10^{-24} mole. #

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For molecules:-

no. of moles of molecules = $\frac{\text{Wt. in gram}}{\text{molecular wt.}}$

= $\frac{\text{Given no. of molecule}}{N_A}$

* How many moles are present in 630 gm of HNO_3?

↳ We have, molecular wt. of HNO_3 = $1 + 14 + (3 \times 16)$ = 63

Now,

No. of moles of HNO_3 = $\frac{\text{Wt. in gm}}{\text{Molecular wt.}}$ = $\frac{630}{63}$ = 10 moles. #

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* How many moles are present in 3.6 gm of water.

↳ Soln:-

$$\text{No. of mole of water}(H_2O) = \frac{\text{Wt. in gm}}{\text{Molecular wt.}} = \frac{3.6}{18} = 0.2 \text{ mole}$$ #

* How many molecules are present in 0.9 gm of Water?

↳ Soln:-

18 gm of water $= 6.023 \times 10^{23}$ molecules

1 gm of Water $= \frac{6.023 \times 10^{23}}{18}$ molecules

0.9 gm of Water $= 3.34 \times 10^{23} \times 0.9$ molecules

$= 3.0115 \times 10^{22}$ molecules. #

For gases at NTP:-

$$\text{no. of mole of gas} = \frac{\text{Given vol}^m \text{ of gas in litres}}{22.4 \text{ litres}}$$

$$= \frac{\text{vol}^m \text{ of given mass in ml}}{22400 \text{ ml}}$$

Calculation is expressed in following terms:-

a) Mass - Mass relation
b) Mass - Volume relation
c) Volume - Volume relation

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g) Mass-mass Relation:-

↳ Reactant and product related with each other in terms of mass on basis of involved in the balanced chemical equation.

If mass of any one reactant or product is given, the mass of another reactant or product can be calculated easily from balanced chemical equation.

Example:-

* Calculate the amout of CaO produced on heating 5 gm of $CaCO_3$?

↳ Soln:-

$$CaCO_3 \xrightarrow{\Delta} CaO + CO_2$$

100 gm of $CaCO_3$ on heating ⟶ 56 gm of CaO

1 gm of $CaCO_3$ on heating ⟶ $\frac{56}{100}$ gm of CaO

5 gm of $CaCO_3$ on heating ⟶ $\frac{56 \times 5}{100}$ gm of CaO

= 2.8 gm of CaO #

* Calculate the amount of CO_2 produced of 10 gm of Na_2CO_3 is heated with dil. HCl.

↳ Soln:-

$$Na_2CO_3 + 2HCl \longrightarrow 2NaCl + CO_2 + H_2O$$

Here,

106 gm of Na_2CO_3 ⟶ 44 gm of CO_2

10 gm of Na_2CO_3 ⟶ $\frac{44}{106} \times 10$ gm of CO_2

= 4.15 gm of CO_2 #

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* What weight of 60% pure dilute H_2SO_4 is required to decompose 25 gm of Chalk ($CaCO_3$).

↳ Soln:-

The balanced chemical eqn is.

$$CaCO_3 + H_2SO_4 \longrightarrow CaSO_4 + H_2O + CO_2$$

100 gm of $CaCO_3$ ⟶ 98 gm of H_2SO_4

1 gm of $CaCO_3$ ⟶ $\frac{98}{100}$ gm of H_2SO_4

25 gm of $CaCO_3$ ⟶ $\frac{98 \times 25}{100} = 24.5$ gm of H_2SO_4

Now,

$x \times 60\% = 24.5$

or, $x \times \frac{60}{100} = 24.5$

$\Rightarrow \boxed{x = 40.83 \text{ gm}}$

OR,

60 gm of H_2SO_4 is present in 100 gm of H_2SO_4

24.5 gm of H_2SO_4 is present in $\frac{100}{60} \times 24.5 = 40.83$ gm

b) Mass-Volume Relation:-

* What volume of CO_2 is required to react with 20 gm of NaOH?

↳ Soln

$$2NaOH + CO_2 \longrightarrow Na_2CO_3 + H_2O$$

80 gm of NaOH react with 22.4 ltr of CO_2

20 gm of NaOH react with $\frac{22.4}{80} \times 20$ ltr. of CO_2

$= 5.55$ ltr. of CO_2

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c) Volume-Volume Relation:-

* What volume of O_2 is necessary for complete combustion of 5 ltr. of propane gas at NTP?

↳ Sol:-

$$C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O + Q$$

Now;

22.4 ltr. of C_3H_8 needs $(22.4 \times 5) = 112$ ltr. of O_2 to react

1 ltr. of C_3H_8 needs $\frac{112}{22.4}$ ltr. of O_2 to react

5 ltr. of C_3H_8 needs $\frac{112}{22.4} \times 5$ ltr. of O_2 to react

$= 25$ ltr. #

Empirical formula:-

↳ It is defined as the simplest chemical formula of compound in which there is a actual whole number atom are present.

Example: Empirical formula of glucose $\Longrightarrow$ (CH_2O)

$$\% \text{ Composition} = \frac{\text{Wt. of individual atom}}{\text{Molecular Wt. of molecule}} \times 100\%$$

For e.g.:- $Na_2S_2O_3$

$$\% \text{ Composition of Na} = \frac{46}{150} \times 100\% = 29.1\%$$

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% Composition of S = $\frac{64}{150} \times 100\% = 40.5\%$

% Composition of O = $\frac{48}{150} \times 100\% = 30.38\%$

* An organic Compound Contains 32% Carbon and 4% hydrogen determine empirical formula of Compound.

Elements	% Composition	Atomic Wt.	No. of Mole	Simple Ratio	Simple Whole no. Ratio
C	32%	12	32/12 = 2.67	$\frac{2.67}{2.67} = 1$	$1 \times 2 = 2$
H	4%	1	4/1 = 4	$\frac{4}{2.67} = 1.5$	$1.5 \times 2 = 3$
O	64%	16	64/16 = 4	$\frac{4}{2.67} = 1.5$	$1.5 \times 2 = 3$

∴ Empirical formula of organic Compound is $C_2H_3O_3$.

Molecular Formula:-

1) Calculate empirical formula of Compound.
2) Calculate molecular weight = 2 x v.d.
3) Calculate value of n = $\frac{\text{Molecular formula wt.}}{\text{Empirical formula wt.}}$
4) Calculate molecular formula = (Emperical formula)$_n$

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* Find the molecular formula of an organic Compound which give the following % Composition.

C = 26.6% ; the v.d. of the compound is 45 amu.
H = 2.22%

Element	% Composition	Atomic wt.	no. of mole	Simple ratio	Simplest whole no. ratio
C	26.6	12	$\frac{26.6}{12} = 2.21$	$\frac{2.21}{2.21} = 1$	1
H	2.22	1	$\frac{2.22}{1} = 2.22$	$\frac{2.22}{2.21} = 1$	1
O	71.18	16	$\frac{71.18}{16} = 4.44$	$\frac{4.44}{2.21} = 2$	2

∴ Empirical formula of Compound = CHO_2 #

Molecular weight = 2 × 45 = 90

Empirical weight = 12 + 1 + 32 = 45

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

Molecular Formula = $2(CHO_2)$
$= C_2H_2O_4$ #

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Limiting reagent (Limiting reactant):-

↳ The reactant which is in deficit amount and finished first in reaction mixture is known as Limiting reagent. It is called so as it limits the formation of product in chemical reaction. The remaining reactant left over unreacted in the reaction mixture is called excess reactant. Limiting reactant is essential in stoichiometric calculation to determine the amount of product formed in chemical reaction.

* 10.6 gm of pure Na_2CO_3 if treated with 7.9 gm of HCl to produce NaCl, H_2O & CO_2.

Q) Find the limiting reagent & Calculate mole of unreacted reagent left over.

Soln:-

The balanced chemical eqn is;

$$\underset{1\,mol}{\overset{106\,gm}{Na_2CO_3}} + \underset{2\,mol}{\overset{73\,gm}{2HCl}} \longrightarrow \underset{2\,mol}{\overset{117\,gm}{2NaCl}} + \underset{1\,mol}{\overset{18\,gm}{H_2O}} + \underset{1\,mol}{\overset{22.4\,ltr}{CO_2}}$$

Here,

106 gm of Na_2CO_3 react with 73 gm of HCl

1 gm of Na_2CO_3 react with $\frac{73}{106}$ gm of HCl

10.6 gm of Na_2CO_3 react with $\frac{73 \times 10.6}{106} = 7.3$ gm of HCl

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$$10.6 \text{ gm of } Na_2CO_3 \begin{cases} \xrightarrow{\text{Consume}} 7.3 \text{ gm of } HCl \\ \text{given } 7.9 \text{ gm of } HCl. \end{cases}$$

Amount of HCl unreacted $= 7.9 - 7.3 = 0.6$ gm #

no. of mole of unreacted HCl $= \frac{0.6}{36.5} = 0.016$ mole #

b) What volume of CO_2 gas is produced at NTP.

* Soln:-

106 gm of Na_2CO_3 gives 22.4 ltr. of CO_2 at NTP

1 gm of Na_2CO_3 gives $\frac{22.4}{106}$ ltr. of CO_2 at NTP.

10.6 gm of Na_2CO_3 gives $\frac{22.4}{106} \times 10.6$

$= 2.24$ ltr of CO_2 at NTP #

c) Calculate the mass of NaCl formed.

* Soln:

106 gm of Na_2CO_3 gives 117 gm of NaCl

10.6 gm of Na_2CO_3 gives $\frac{117}{106} \times 10.6$

$= 11.7$ gm of NaCl #

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Avogadro's Hypothesis:-

↳ It state that, "Under similar condition of temperature and pressure, equal volume of all gases contain the equal numbers of molecules or numbes of moles."

Mathematically,

Volume of gas $\propto$ numbes of moles

i.e., $V \propto n$ [$\because$ P & T are constant]

Explanation:-

Hydrogen + Oxygen ⟶ Water

$2H_2 + O_2 \longrightarrow 2H_2O$

2 Vol^m	1 Vol^m	2 Vol^m
2n mol.	n mol.	2n mol.
2 mol	1 mol	2 mol
1 molecules	$\frac{1}{2}$ molecules	1 molecules

This means that 1 molecules of H_2O formed by combination of 1 molecule of H_2 & half molecule of O_2. So, combination bet^n fraction of molecule is possible.

Application of Avogadro's hypothesis:-

[1] Determination of molecular wt. & Vapour density:-

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Vapour density may be defined as the ratio of wt. of any volume of gas at NTP. to the wt. of same volume of Hydrogen gas at NTP.

i.e., $VD = \dfrac{\text{Wt. of any Vol}^m \text{ of gas at NTP.}}{\text{Wt. of same vol}^m \text{ of } H_2 \text{ gas at NTP}}$

On applying Avogadro's Law (i.e. $V \propto n$)

$$VD = \frac{\text{Wt. of n molecules of any gas}}{\text{Wt. of n molecules of } H_2 \text{ gas}}$$

$$= \frac{n(\text{Wt. of 1 molecules of any gas})}{n(\text{Wt. of 1 molecule of } H_2 \text{ gas})}$$

$$= \frac{1}{2}\left[\frac{\text{Wt. of 1 molecule of gas}}{\text{Wt. of 1 atom of } H_2 \text{ gas}}\right]$$

$$= \frac{1}{2} \text{ molecular wt.}$$

$\therefore \boxed{\text{Molecular Wt.} = 2 \times VD}$

[2] Determination of Atomicity of elementary gas:-

Let,

H_2 gas combine with chlorine gas to form HCl.

Experimentally,

Hydrogen gas + Chlorine gas → Hydrogen Chloride gas

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$$H_2 + Cl_2 \longrightarrow 2HCl$$

H_2	Cl_2	$2HCl$
↓	↓	↓
1 vol^m	1 vol^m	2 vol^m

Applying Avogadro's Law (i.e. $V \propto n$)

n molecule	n molecule	2n molecule
1 molecule	1 molecule	2 molecule
$\frac{1}{2}$ molecule	$\frac{1}{2}$ molecule	1 molecule

It is found that 1 molecule of HCl contains 1 atom of H_2 which comes from $\frac{1}{2}$ molecules of H_2

$\frac{1}{2}$ molecule of H_2 = 1 atom of H_2

1 molecule of H_2 = 2 atom of H_2

Thus,

Hydrogen is diatomic.

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[3] Determination of Molar volume of gas:-

Soln:-

We know:

Molecular wt. = 2 x VD

but,

$$VD = \frac{\text{Wt. of 1 ltr. of gas at NTP}}{\text{Wt. of 1 ltr. of } H_2 \text{ gas at NTP.}}$$

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So,

$$\text{Molecular wt.} = \frac{2 \times \text{Wt. of 1 ltr. of gas at NTP}}{\text{Wt. of 1 ltr. of } H_2 \text{ at NTP}}$$

We know;

Wt. of 1 ltr. of H_2 gas at NTP = 0.089 gm

Therefore,

$$\text{molecular wt.} = \frac{2 \times \text{wt. of 1 ltr. of gas}}{0.089}$$

$$\text{or, Molecular wt.} = 22.4 \times \text{Wt. of 1 ltr. of gas}$$

$$= \text{Wt. of 22.4 ltr. of gas}$$

Hence, the gram molecular volume of all gases at NTP occupies 22.4 ltr.

[4] Determination of molecular ~~wt.~~ Formula from volumetric Composition:-

↳

By experiment;

* It is found that Nitrous Oxide Contains its own volume of Nitrogen and its vapour density is 22. Find its molecular formula.

* Here,

1 volm of nitrous oxide Contains 1 volm of Nitrogen

Let,

1 volm of Nitros oxide Contains n molecules

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Therefore,

1 vol^m of Nitrogen also contains n molecules

n molecules of Nitrous Oxide contains n molecules of N_2
1 molecules of Nitrous oxide contains 1 molecule of N_2
1 molecule of Nitrous Oxide contains 2 atom of N_2

$\therefore$ Nitrogen is diatomic.

Hence, formula of Nitrous oxide should be

$$N_2O_x$$

i.e. $(2 \times 14) + (x \times 16) = 44$

$\Rightarrow \boxed{x = 44}$ [$\because$ Molecular wt. $= 2 \times VD = 2 \times 22 =$

Hence, Actual molecular formula of Nitrous Oxide is N_2O. #

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