Stoichiometry

→ Understand & appreciate the role of chemistry in different spheres of life.

→ Explain various laws of chemical combination

→ Appreciate significance of atomic mass, avg atomic mass.

→ molecular mass & formula mass;

→ Describe the terms - mole & molar mass.

→ Calculate the mass % of different elements constituting a compound.

→ Determine the Empirical formula & molecular formula for a compound from Experimental data.

→ perform the stoichiometric calculations.

\# Chemistry is nothing but the science of atoms, molecules and their transformations

\# chemistry is the branch of science that studies the composition, props & interaction of matter

Why study chemistry?

we see diamond, coal → temp, props, heat of comb etc.

plastic cover formation

To know about fibres, soaps we use everything the knowledge of chemistry is must.

can we count the number of atoms & molecules in a given mass of matter? say in 10gm of gold.
can we see atoms & molecules?
These answers can be given by the knowledge of chemistry

Advantages of studying chemistry:-

with the better understanding of chemical principles it has now become possible to design and synthesize new materials having specific magnetic, electric & optical props.

→ Eg:- Safer alternatives to CFCs have been successfully synthesised.

Some problems still exist:-

Management of the green house gases like methane, CO_2 etc,
understanding of bio-chemical processes, use of enzymes for large scale production of chemicals & synthesis of new exotic materials are some of the intellectual challenges for the future generation of chemists.

properties of matter:-

physical props are those props which can be measured or observed without changing the identity
Eg:- colour, odour, M.P, B.P, density etc,

the measurement or observation of chemical props require a chemical change to occur. chemical props
Eg:- include acidity, basicity, combustibility etc,

prefixes used in the SI system:

10^{-6}	micro	10^{-9}	nano
10^{-3}	milli	10^{-12}	pico
10^{-2}	centi	10^{-15}	femto
10^{-1}	deci		
10	deca		
10^{2}	hecto	10^{15}	peta
10^{3}	kilo	10^{18}	exa
10^{6}	mega	10^{23}	zeta
10^{9}	giga	10^{24}	yotta
10^{12}	tera		

Eg:- 4500 can be written as 4.5×10^{3}.

Convert the following into basic units.

(1) 28.7 pm ↳ pico metre

$1 pm = 10^{-12} m$.

$28.7 pm = 28.7 \times 10^{-12} m$.

$= 2.87 \times 10^{-11} m$

(2) $25365 mg = 25365 \times 10^{-6} kg$.

$= 2.5365 \times 10^{-2} kg$

If the speed of light is $3 \times 10^{8} ms^{-1}$. calculate the distance covered by light in 2 ns.

$D = S \times t = 3 \times 10^{8} \frac{m}{\cancel{sec}} \times 2 \times 10^{-9} \cancel{sec}$

$= 6 \times 10^{-1} m$

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Law of chemical Combination

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

(1) law of conservation of mass

(2) law of definite proportions

(3) law of multiple proportions

(4) Gay Lussac's law of gaseous volumes

(5) Avagadro law.

(1) law of conservation of mass :- It states that matter can neither be created nor destroyed.

This law was put forth by Antoine lavoisier in 1789. he discovered oxygen ?

He performed careful experimental studies for combustion laws, noting reactants & products for reaching to the above conclusion.

(2) law of multiple proportions :- According to this law, if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers.

This law was proposed by Dalton in 1803

this is const

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O \quad (2g + 16g = 18g)$ ∴ 16 : 32

$H_2 + O_2 \rightarrow H_2O_2 \quad (2g + 32g = 34g)$ 1 : 2

Law of definite proportions:-

It stated that a given compound always contains exactly the same proportion of elements by weight.

It was given by a French chemist, Joseph Proust in 1806.

He observed 2 samples of cupric carbonate

- one of which was of natural origin
- other was synthetic one

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

Cupric ~~chlate~~ carbonate	% of copper	% of carbon	% of O_2
natural	51.35	9.74	38.91
synthetic	51.35	9.74	38.91

This law is also called as law of definite composition.

Gay Lussac's law of gaseous volumes:-

This law was given by Gay Lussac in 1808.

He observed that when gases combine or are produced in a chemical rxn they do so in a simple ratio by volume provided all gases are at same temp & pressure.

Hydrogen + oxygen → water

100 ml   50 ml   100 ml

100 : 50 = 2 : 1 → simple ratios

500Rs Day 10cm 5Rs 1Rs

Avogadro law

In 1811, Avogadro proposed that Equal volumes of gases at the same temp., & pressure should contain Equal number of molecules.

Avogadro made a distinction b/n atoms & molecules.

1 mole of Ideal gas at STP $\rightarrow$ 22.4 ltr.

$\downarrow$

6.023×10^{23}

Question time

The following data are obtained when dinitrogen and dioxygen react together to form different compound. Which law of chemical combination is obeyed by the above experimental data?

Mass of dinitrogen	Mass of Hydrogen
14×2	$16 \times 2 = 32$
14×2	$32 \times 2 = 64$
28	32
28	80

$32 : 64 : 32 : 80 = 2 : 4 : 4 : 5$

This is a simple whole number ratio, thus, it obeys law of multiple proportions.

The mass of H_2 required to form different kind of compounds are in the whole number ratio.

Dalton's atomic theory

* Matter consists of indivisible atoms but not true
X as atoms are divided into protons, neutrons etc.
* All the atoms of a given Element have identical props including identical mass. Atoms of different Elements differ in mass.
* compounds are formed when atoms of different Elements combine in a fixed ratio.
* chemical ~~Elements~~ rxns involve reorganisation of atom
✓ These are neither created nor destroyed in a chemical rxn.
but in radioactive Elements, atoms get broken
so the above statement is false.

Atomic mass

Mass of an atom.
How is atomic mass of atom determined today?
Sophisticated techniques eg: mass spectrometry is used
How was atomic mass of atom determined in 19th Century?
They used to find using stoichiometry.
took base Element as Hydrogen with atomic no =1

$$H_2 + \frac{1}{2}O_2 \rightarrow H_2O$$

By doing lot of chemical ~~calc~~ calculate the atomic mass of O_2 (stoichiometry)

Atomic mass: History

→ The first scientists to determine atomic mass was John Dalton in 1803. It was called atomic weight that time.

→ Atomic weight was originally defined relative to that of lightest element hydrogen taken as 1.

But since hydrogen is unreactive. It can't react with most of the elements.

→ Thus, they took O_2 as base instead Hydrogen. Both chemist & physics started using O_2 as base. Atomic mass 16 (whole number).

→ It was later found that natural oxygen contains other isotopes to O_2 too.

→ chemists picked naturally occurring O_2, which is a mixture of isotopes of O_2-16, O_2-17 & O_2-18 (Atomic mass 16.008).

→ physicists picked up carbon 12 as base based on mass spectrometry & mass of one carbon-12 atom is a whole number i.e., 12

Since it reduces calculations physicists still use carbon as base i.e., $12 \times 5 = 60$
& $16.008 \times 5 =$ (little tough)

→ present system of atomic masses is based on carbon-12 as the standard and has been agreed upon in 1961 as only the mass of one carbon-12 atom is a whole number.

→ In this system, ^{12}C is assigned a mass of exactly 12 atomic mass unit (amu) and masses of all other atoms are given relative to this standard.

Avg atomic mass :-

Many naturally occurring elements exist as more than one isotope. When we take into account the existence of these isotopes and their relative abundance (percent occurrence).

Eg:-

Isotopes	Relative abundance	AMU
C12	98.892	12
C13	1.108	13.0035
C14	2×10^{-10}	14.0037

$$\left[\frac{98.892}{100}\times 12+\frac{1.108}{100}\times 13.00354+\frac{2\times10^{-10}}{100}\times 14.0037\right]$$

is the Avg. atomic mass of carbon.

11^{ly} O16, O17, O18 $\Rightarrow$ Avg. atomic mass = 16.008

Calculate the atomic mass (avg) of chlorine using the following data:

Isotopes	Relative abundance	AMU
Cl 35	75.77	34.9689
Cl 37	24.23	36.9659

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$$RAM\ Cl = \frac{75.77}{100} \times 34.9689 + \frac{24.23}{100} \times 36.9659$$

$$= 26.4952 + 8.9568$$

$$= 35.452$$

Q) Use data given in following table to calculate molar mass of naturally occuring argon isotopes

Isotope	molar mass	Abundance
36 Ar	35.96755 g	0.337 %
38 Ar	37.96272 g	0.063 %
40 Ar	39.9624 g	99.6 %

$$\frac{0.337}{100} \times 35.96755 + \frac{0.063}{100} \times 37.96272 + \frac{99.6}{100} \times 39.9624$$

$= 39.947 \rightarrow$ Relative atomic mass of argon

Molecular mass :- It is the sum of atomic masses of all the elements present in a molecule

molecular mass of methane (CH_4)

$$= (12.011\ u) + 4(1.008\ u) = 16.043\ u$$

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Q] Calculate the molecular mass of the following
Q, (1) H_2O (2) CO_2 (3) CH_4.

→ H_2O → $2\times H + 1\times O = 2\times 1 + 1\times(16) = 18$ amu

CO_2 → $1\times C + 2\times O = 12 + 2\times 16 = 44$ amu

CH_4 → $1\times C + 4\times 4$ → $12 + 4\times 1 = 16$ amu.

Formula mass :- Some substances such as sodium chloride do not contain discrete molecules as their constituent units Eg: NaCl.
↳ $Na^+ + Cl^-$

formula mass of NaCl = atomic mass of sodium + atomic mass of chlorine

= 23.04 + 35.54 = 58.54

Mole concepts

Guess the no. of molecules in 1 gm of salt?
It will be huge number.
To handle such large numbers, a unit of similar magnitude is required

In SI system, mole was introduced as seventh base quantity for the amount of a substance.

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

1gm of H_2 → x

12gm of C → x

16gm of O_2 → x atoms.

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this no. of entities in 1mol is so important as it is given a separate name and symbol, known as Avagadro constant.

1mol of H_2 atoms = 6.022×10^{23} atoms.

1mol of H_2O molecule = 6.022×10^{23} molecule

1mol of NaCl = 6.022×10^{23} units of NaCl.

for 6.022×10^{23} atoms - 1gm of H_2

where as for

6.022×10^{23} atoms - 16gm of O_2.

as wt of $O_2 > H_2$.

Lets suppose wt of H_2 molecule as a.

i.e., a → ~~1gm~~ $a\times6.022\times10^{23}$ = 1gm H_2

$12a\times6.022\times10^{23}$ = 12gm C.

$16a\times6.022\times10^{23}$ = 16gm O_2

Molar mass:-

→ The mass of one mole of a substance in grams is called molar mass. 6.023×10^{23} atoms of O_2 = 16gm

→ The molar mass in grams is numerically equal to atomic/molecular/formula mass in u,

Eg:- molar mass of O_2 = 16.02gm.

i.e., mass of 6.023×10^{23} atoms of O_2 = 16.02gm

* Percentage composition

Ratio in which its constituents present in a given compound.

H_2O — $H \rightarrow 2\times1$ moles — 2gm

$O \rightarrow 16$ — 16 gm.

mass % of $H_2 = \frac{2gm}{2+16} = \frac{2}{18g} \times 100 \rightarrow 11.18\%$

mass % of $O_2 = \frac{16}{18} \times 100 = 88.79\%$

Info from % composition : Empirical formula.

An Empirical formula represents the simplest whole number ratio of various atoms present in a compound. Eg:- $C_6H_6 \rightarrow$ E.F C_1H_1

Molecular formula shows the exact no. of different types of atoms present in a molecule of a compound.

1. A compound contains 4.07% of hydrogen, 24.27% of carbon & 71.65% of chlorine. Its molar mass is 98.96g. What are its E.F & M.F's?

100 gm of compound =
- 4.07 g of H_2 $\rightarrow \frac{4.07}{1.001}$ mol $\simeq 4$ mol
- 24.27 g of C $\rightarrow \frac{24.27}{12.01}$ mol $\simeq 2$ mol
- 71.65 g of Cl $\rightarrow \frac{71.65}{35.5}$ mol. $\simeq 2$ mol.

It means ratio of $H_2 : C : Cl$

$= 4 : 2 : 2$

$= 2 : 1 : 1$

So, the E.F is CH_2Cl.

Molar mass $[CH_2Cl] = 12 + 2 \cdot 1 + 35.5 = 49.5$

$$\frac{98.96}{48.5} \simeq 2$$

$\Rightarrow$ molecular formula: [Empirical formula]$_n$

$= [CH_2Cl]_2$

$= C_2H_4Cl_2$ //

Q] Calculate the mass % of different elements in Sodium sulphate, (Na_2SO_4).

Sol]: molar mass of Na_2SO_4 =

$2 \times 23 + 34 + 4 \times 16 = 142$ gm.

In 142 gm of Na_2SO_4 — 46 gm of Na, 34 gm of S, 64 gm of O_2.

% Na $\rightarrow \frac{46}{142} \times 100 = 32.4\%$

% S $\rightarrow \frac{34}{142} \times 100 = 22.6\%$

% $O_2 \rightarrow \frac{64}{142} \times 100 = 45\%$

Find the E.F of an oxide of iron which has 69.9% iron & 30.1% dioxygen by mass.

In 100 gm of Fe oxide — 69.9 gm of Fe $\Rightarrow \frac{69.9}{55.85} = 1.25$ mol

30.1 gm of $O_2 \Rightarrow \frac{30.1}{32} = 1.88$ mol

Ratio $\Rightarrow \frac{1.25}{1.88} = 1:1.5$

$\simeq 2:3$ (as we should convert it into a whole numb].

[Fe_2O_3] — Empirical formula,

<u>Stoichiometry</u>:— It deals with the calculations of masses (sometimes volumes also) of the reactants & the products involved in a chemical rxn.

This is done using balanced chemical rxn.

<u>What is chemical rxn?</u>

whenever a chemical change occurs, we can say that a chemical rxn has taken place.

Magnesium + oxygen → magnesium oxide.

<u>Balanced chemical Eqn</u>:— The total mass of the elements in the products of a chemical rxn has to be equal to the total mass of the elements present in the reactants.

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

→ 1 mol of CH_4 reacts with 2 moles of O_2 to give mole of CO_2 & 2 moles of H_2O.

→ 1 molecule of $CH_4(g)$ reacts with 2 molecules of $O_2(g)$ to give 1 molecule of $CO_2(g)$ and 2 molecules of $H_2O(g)$.

→ 16g of $CH_4(g)$ reacts with 2×32g of $O_2(g)$ to give 44 g of $CO_2(g)$ and 2×18g of $H_2O(g)$.

Q] Calculate the amount of water (g) produced by the combustion of 16g of CH_4.

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

16gm 32gm 44gm 36gm

Limiting Reagent :- In a chemical rxn, reactant which is present in the lesser amount gets consumed after sometime and after that no further rxn takes place whatever be the amount of the other reactant present. Hence, the reactant which gets consumed, limits the amount of product formed and is, therefore, called the limiting reagent

$$A + B \rightarrow P$$

limiting reagent

20gm of R_1 + 5gm of R_2 $\rightarrow$ O.

10gm of R_1 + 5gm of R_2 $\rightarrow$.

Limiting reagent

this has to be 20 but it is less and is equal to 10

Thus, it is L.R.

Eq] 50 kg of N_2 and 10kg of H_2 (g) are mixed to produce NH_3(g). Calculate the NH_3(g) formed. Identify the limiting reagent in the production of NH_3 in this situation.

$$N_2 + 3H_2 \rightarrow 2NH_3$$

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$N_2 + 3H_2 \rightarrow 2NH_3$

I have 50 Kg. 10 Kg

Lets see how much they are required?

From the rxn we have,

1 mole of N_2 requires 3 moles of H_2

28 gm of N_2 " = 3×2 gm of H_2.

1 gm of N_2 requires = ? $\frac{1 \times 6}{28}$ of H_2

1 gm of H_2 requires = $\frac{28}{6}$ gm of N_2.

10 Kg of H_2 requires = $\frac{28}{6} \times 10 \times 10^3$ gm of N_2

= 46.67 Kg //.

So, we can see N_2 (50 Kg given) is more than that req. (46.67 Kg).

Thus, limiting reagent is H_2 as it get consumed early.

From the rxn

3 mol of H_2 → 2 mol of NH_3.

or 6 gm of H_2 → 2×17 gm of NH_3.

relying only on H_2 as the limiting reagent, 10 Kg of H_2 → ? $\frac{10 \times 10^3 \times 34}{6}$ gm of NH_3.

= 56.66 Kg //

Q] calculate the amount of CO_2 that could be produced when.

(1) 1 mole of carbon is burnt in air

$C + O_2 \rightarrow CO_2$

1 mol of C → 1 mole of CO_2

(2) 1 mole of carbon is burnt in 16g of dioxygen

1 mole of C → 1 mole of O_2

12 gm of C requires $\underline{32}$ gm of O_2 is the required amount

? — 16 gms ~~..~~ 16gm < 32gm

O_2 is the limiting reagent

$\frac{\cancel{16}^{6} \times \cancel{12}}{\cancel{32}_{2}}$ gm of C.

So, 6 gm of C is required.

12gm → 44 gm of CO_2

6 gm → ?

$\frac{\cancel{6} \times \cancel{44}^{22}}{\cancel{12}_{2}}$ gm of CO_2 is produced

$\underline{\text{or}}$ 1 mol of carbon + 1 mol of O_2 → 1 mol of CO_2

12gm of C + 32gm of O_2 → 44 gm of CO_2.

32gm of O_2 → 44 gm of CO_2.

16gm of O_2 → $\frac{16 \times 44}{32}$ gm of CO_2

= 22gm of CO_2

(5) 2 moles of carbon are burnt in 16g of dioxygen
Since O_2 is the limiting reagent, 16 gm of O_2 will always form 22gm of CO_2.

Q) In a rxn $A + B_2 \rightarrow AB_2$. Identify limiting reagent, if any, in following rxn mixtures.

* 300 atoms of A + 200 molecules of B.
* 2 mol A + 3 mol B.

Sol:-

$$A + B_2 \rightarrow AB_2$$

1 mol of A + 1 mol of B → 1 mol of AB_2.

1 molecule of A + 1 mol of B → 200 mol of AB_2

200 molecule of A + 200 mol of B

But 300 mol of A is given. (300 > 200)

hence B is the limiting reagent.

2 mol A + 3 mol B.

A is the limiting reagent

Q) Dinitrogen & dihydrogen react with each other produce ammonia according to the following chemi Eqn: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

(*) calculate the mass of ammonia produced 2×10^3 g of N_2 reacts with 10^3 g of H_2.

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28 gm N_2 + 6 gm H_2 → 34 gm of NH_3.

2×10^3 g of N_2 –

(2) will any of the two reactants remain unreacted,

(3) If yes, which one & what would be its mass,

Sol. 1 mol of N_2 + 3 mol of H_2 → 2 mol of NH_3.

28 gm of N_2 + 6 gm of H_2 → 34 gm of NH_3.

from the eqn we can see that

for 28 gm of N_2 we require 6 gm of H_2.

Hence, H_2 should be the L.R. ~~X~~

but still lets try to exactly find it

28 gm of N_2 → 6 gm of H_2

So, 1 gm of N_2 → $\frac{6}{28}$ gm of H_2.

for 2 Kg of N_2 → $\frac{6}{28} \times 2 \times 10^3$

$= 428.6$ gm of H_2 is requi

How many gms of H_2 we have 1000 gm.

Hence, N_2 is the L.R.

H_2 unreacted $= 1000 - 428.6 = 571.4$ gm will be left.

we also have to find the mass of ammonia produced. Since N_2 is the L.R we get

28 g of N_2 → 34 gm of NH_3.

2×10^3 g of N_2 → $\frac{2 \times 10^3 \times 34}{28} = 2428.57$ gm

Conc. in solns

A majority of rxn in the laboratories are carried out in solns.

The conc. of a soln is the amount of substance present in its given volume can be expressed in any of the following ways.

- Mass % or wt. % (w/w %)
- Mole fraction
- Molarity
- Molality

(1) Mass % or wt. %. It is one way of representing the conc. of an element in a compound or a component in a mixture. Mass % is calculated as the mass of a component divided by the total mass of the mix, multiplied by 100 %.

H_2O

mol of H 2 (2 at H) → 2×1 gm of H_2

1 ~~mol~~ ~~mole~~ of O_2 → 16 gm of O_2

$$\text{mass \% of } H_2 = \frac{2}{16+2} \times 100 = \%$$

$$\text{Mass \% of } O_2 = \frac{16}{18} \times 100 = \%$$

Q] A soln is prepared by adding 2g of a substance A to 18g of water. Calculate the mass % of the th.

2gm of solute

18 gm of water

Total 20gm of mixt

$$\% \text{ substance} = \frac{2}{20} \times 100 = 10\%$$

* Mole fraction :- It is the ratio of no. of moles of a particular component to the total no. of moles of the soln.

3 moles of x

7 moles of y

$$\text{mole fraction of } x = \frac{3}{10} = 0.3$$

$$y = 0.7$$

** Molarity (M)

It is defined as the no. of moles of solute in 1 lt of soln.

Eg:- A 4gm of sugar cube (sucrose: $C_{12}H_{22}O_{11}$) is dissolved in a 350ml teacup filled with hot water. What is the molarity of the sugar soln?

$$\text{molar mass } (C_{12}H_{22}O_{11}) = 12\times12 + 22\times1 + 11\times16$$

$$= 342 \text{ gm/mol}$$

342gm ~~food~~ = 1 mol of sucrose.

4gm — $\frac{4}{342}$ mol of sucrose = 0.117 M,

$$M = \frac{\text{mol. of sucrose}}{\text{Vol. of soln (L)}} = \frac{0.117}{0.350} = 0.332\ \frac{\text{mol}}{\text{L}}$$

molality (m) :- No. of moles of solute present in 1 Kg of solvent. It is denoted by m.

$$m = \frac{\text{moles of solute}}{\text{Kg of solvent}} = \frac{\text{mole}}{\text{Kg}},$$

Q] above question, what is the molality of the sugar soln? Given, density of water @ 20° = 0.995 g/ml.

Mol. wt of sucrose = 342 gm/mole.

$$\text{mol. of sucrose} = \frac{4}{342} = 0.117\ \text{mol}.$$

$$\text{density} = \frac{m}{V}.$$

mass of water = density × vol.

$$= 0.975\ \frac{\text{g}}{\text{ml}} \times 350.$$

$$= 0.341\ \cancel{\times 10^{3}}\ \text{Kg}$$

∴ mass of water 0.341 Kg.

$$m = \frac{0.117}{0.341} = 0.634\ \frac{\text{mol}}{\text{Kg}}$$

Q) Calculate the mass of sodium acetate (CH_3COONa) required to make 500ml of 0.375 molar aqueous soln. Molar mass of sodium acetate is 82.0245 $g\,mol^{-1}$:

Sol:- 0.375 molar soln means

0.375 moles of CH_3COONa in 1000 ml of H_2O

or $\frac{0.375}{2}$ moles of CH_3COONa in 500 ml of H_2O.

1.875 ←

given:- 1 mole of CH_3COONa = 82.0245 gm.

1.875 mol of " = ?

1.875 × 82.0245 gm

= 15.38 gm,

Q) A sample of drinking water was found to be severely contaminated with chloroform, $CHCl_3$. The level of contamination was 15 ppm (by mass).

(1) Express this in percent by mass of mass%?

$$15ppm = \frac{15 \text{ parts of } CHCl_3}{\text{million parts of } H_2O}$$

$$\text{mass \%} = \frac{15}{10^6} \times 100 = 1.5 \times 10^{-3} \%$$

(2) Determine the molality of chloroform in the water sample.

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$$\text{molality}(m) = \frac{\text{mol of } CHCl_3}{\text{kg of solvent}}$$

Lets take 1kg of solvent = 1000 gm = 1.5×10^{3}

10^{6} gm = 15

2gm = 15×10^{-6} gm.

10^{-3} gm = ?

$$\frac{10^{-3} \times 15 \times 10^{-6}}{1} \text{ gm}$$

$= 15 \times 10^{-6}$ kg.

$°F = \frac{9}{5} \, 1.8 °C + 32$