Mole Concept

Introduction :

There are a large number of objects around us which we can see and feel.
Anything that occupies space and has mass is called matter.

Ancient Indian and Greek Philosopher's beleived that the wide variety of object around us are made from combination of five basic elements : Earth, Fire, Water, Air and Sky.
The Indian Philosopher kanad (600 BC) was of the view that matter was composed of very small, indivisible particle called *"parmanus"*.
Ancient Greek Philosophers also believed that all matter was composed of tiny building blocks which were hard and indivisible.
The Greek philosopher Democritus named these building blocks as atoms, meaning indivisible.
All these people have their philosophical view about matter, they were never put to experimental tests, nor ever explain any scientific truth.
It was ***John Dalton*** who firstly developed a theory on the structure of matter, later on which is known as ***Dalton's atomic theory.***

DALTON'S ATOMIC THEORY :

- Matter is made up of very small indivisible particles called atoms.
- All the atoms of a given element are identical in all respect i.e. mass, shape, size, etc.
- Atoms cannot be created or destroyed by any chemical process.
- Atoms of different elements are different in nature.

Classification of matter

- **on the basis of physical behaviour**
  - Solids
  - Liquids
  - Gases
- **on the basis of chemical behaviour**
  - Pure substances
    - Element
    - Compound
  - Mixtures

Basic Definitions :

Relative atomic mass :

One of the most important concept come out from Dalton's atomic theory was that of relative atomic mass or relative atomic weight. This is done by expressing mass of one atom with respect to a fixed standard. Dalton used hydrogen as the standard (H = 1). Later on oxygen (O = 16) replaced hydrogen as the reference. Therefore relative atomic mass is given as

On hydrogen scale :

$$\text{Relative atomic mass (R.A.M)} = \frac{\text{Mass of one atom of an element}}{\text{mass of one hydrogen atom}}$$

On oxygen scale :

$$\text{Relative atomic mass (R.A.M)} = \frac{\text{Mass of one atom of an element}}{\frac{1}{16}\times\text{mass of one oxygen atom}}$$

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- The present standard unit which was adopted internationally in 1961, is based on the mass of one carbon-12 atom.

$$\text{Relative atomic mass (R.A.M)} = \frac{\text{Mass of one atom of an element}}{\frac{1}{12}\times \text{mass of one C}-12\text{ atom}}$$

Atomic mass unit (or amu) :

The atomic mass unit (amu) is equal to $\left(\frac{1}{12}\right)^{th}$ mass of one atom of carbon-12 isotope.

$\therefore$ 1 amu $= \frac{1}{12} \times$ mass of one C-12 atom

$\simeq$ mass of one nucleon in C-12 atom.

$= 1.66 \times 10^{-24}$ gm or 1.66×10^{-27} kg

- **one amu is also called one Dalton (Da).**
- **Today, amu has been replaced by 'u' which is known as unified mass**

Atomic & molecular mass :

It is the mass of 1 atom of a substance it is expressed in amu.

- Atomic mass = R.A.M × 1 amu

$$\text{Relative molecular mass} = \frac{\text{mass of one molecule of the substance}}{\frac{1}{12}\times \text{mass of one} - \text{C-12 atom}}$$

- Molecular mass = Relative molecular mass × 1 amu

Note : Relative atomic mass is nothing but the number of nucleons present in the atom.

Solved Examples

Example-1 Find the relative atomic mass of 'O' atom and its atomic mass.

Solution The number of nucleons present in 'O' atom is 16.

$\therefore$ relative atomic mass of 'O' atom = 16.

Atomic mass = R.A.M × 1 amu = 16 × 1 amu = 16 amu

Mole : The Mass / Number Relationship

Mole is a chemical counting SI unit and defined as follows :

A mole is the amount of a substance that contains as many entities (atoms, molecules or other particles) as there are atoms in exactly 0.012 kg (or 12 gm) of the carbon-12 isotope.

From mass spectrometer we found that there are 6.023×10^{23} atoms present in 12 gm of C-12 isotope.

The number of entities in 1 mol is so important that it is given a separate name and symbol known as Avogadro constant denoted by N_A.

i.e. on the whole we can say that 1 mole is the collection of 6.02×10^{23} entities. Here entities may represent atoms, ions, molecules or even pens, chair, paper etc also include in this but as this number (N_A) is very large therefore it is used only for very small things.

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HOW BIG IS A MOLE ?

Amount of water in world's oceans (litres)

Age of earth (seconds)

Population of earth

Avogadro's number 602,200,000,000,000,000,000,000

Distance from earth to sun (centimeters)

❍ **Note :** **In modern practice gram-atom and gram-molecule are termed as mole.**

Gram Atomic Mass :

The atomic mass of an element expressed in gram is called gram atomic mass of the element.

or

It is also defined as mass of 6.02×10^{23} atoms.

or

It is also defined as the mass of one mole atoms.

For example for oxygen atom :

Atomic mass of 'O' atom = mass of one 'O' atom = 16 amu

gram atomic mass = mass of 6.02×10^{23} 'O' atoms

$= 16 \text{ amu} \times 6.02 \times 10^{23}$

$= 16 \times 1.66 \times 10^{-24} \text{ g} \times 6.02 \times 10^{23} = 16 \text{ g}$

$(\because \quad 1.66 \times 10^{-24} \times 6.02 \times 10^{23} \simeq 1)$

Solved Examples

Example-2 How many atoms of oxygen are their in 16 g oxygen.

Solution Let x atoms of oxygen are present

So, $16 \times 1.66 \times 10^{-24} \times x = 16$ g

$$x = \frac{1}{1.66 \times 10^{-24}} = N_A$$

Gram molecular mass :

The molecular mass of a substance expressed in gram is called the gram-molecular mass of the substance.

or

It is also defined as mass of 6.02×10^{23} molecules

or

It is also defined as the mass of one mole molecules.

For example for 'O_2' molecule :

Molecular mass of 'O_2' molecule = mass of one 'O_2' molecule

= 2 × mass of one 'O' atom

= 2 × 16 amu

= 32 amu

gram molecular mass = mass of 6.02×10^{23} 'O_2' molecules = $32 \text{ amu} \times 6.02 \times 10^{23}$

$= 32 \times 1.66 \times 10^{-24} \text{ gm} \times 6.02 \times 10^{23} = 32 \text{ gm}$

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Solved Examples

Example-3 The molecular mass of H_2SO_4 is 98 amu. Calculate the number of moles of each element in 294 g of H_2SO_4.

Solution Gram molecular mass of H_2SO_4 = 98 gm

moles of $H_2SO_4 = \frac{294}{98}$ = 3 moles

H_2SO_4	H	S	O
One molecule	2 atom	one atom	4 atom
$1 \times N_A$	$2 \times N_A$ atoms	$1 \times N_A$ atoms	$4 \times N_A$ atoms
∴ one mole	2 mole	one mole	4 mole
y **3 mole**	**6 mole**	**3 mole**	**12 mole**

Gay-Lussac's Law of Combining Volume :

According to him elements combine in a simple ratio of atoms, gases combine in a simple ratio of their volumes provided all measurements should be done at the same temperature and pressure

$$\begin{array}{cccc} H_2(g) & + \; Cl_2(g) & \longrightarrow & 2HCl \\ 1\text{ vol} & 1\text{ vol} & & 2\text{ vol} \end{array}$$

Avogadro's hypothesis :

Equal volume of all gases have equal number of molecules (not atoms) at same temperature and pressure condition.

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

At S.T.P. condition : temperature = 0°C or 273 K
pressure = 1 atm = 760 mm of Hg

and volume of one mole of gas at STP is found to be experimentally equal to 22.4 litres which is known as molar volume.

Note : **Measuring the volume is equivalent to counting the number of molecules of the gas.**

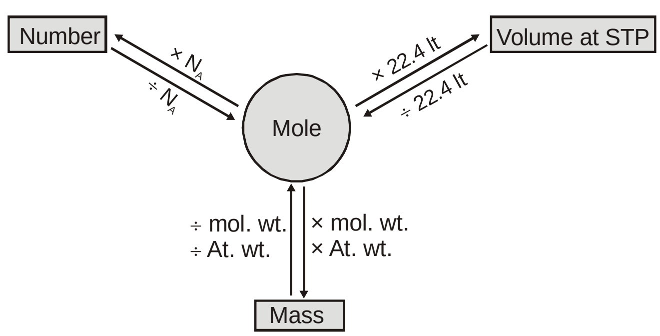
Solved Examples

Example-4 Calculate the volume in litres of 20 g hydrogen gas at STP.

Solution No. of moles of hydrogen gas = $\frac{\text{Mass}}{\text{Molecular mass}} = \frac{20\text{ gm}}{2\text{ gm}}$ = 10 mol

volume of hydrogen gas at STP = 10 × 22.4 lt.

Y-map : Interconversion of mole - volume, mass and number of particles :



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The laws of chemical combination :

Atoine Lavoisier, ***John Dalton*** and other scientists formulate certain law concerning the composition of matter and chemical reactions. These laws are known as the law of chemical combination.

(i) The law of conservation of mass :

In a chemical change total mass remains conserved.
i.e. mass before reaction is always equal to mass after reaction.

Solved Examples

Example-5 $H_2(g) + \frac{1}{2} O_2(g) \longrightarrow H_2O(l)$

Solution

	$H_2(g)$	+ $\frac{1}{2} O_2(g)$	$\longrightarrow H_2O(l)$
Before reaction initially	1 mole	$\frac{1}{2}$ mole	0
After the reaction	0	0	1 mole

mass before reaction = mass of 1 mole $H_2(g)$ + $\frac{1}{2}$ mole $O_2(g)$

= 2 + 16 = 18 gm

mass after reaction = mass of 1 mole water = 18 gm

(ii) Law of constant or Definite proportion :

All chemical compounds are found to have constant composition irrespective of their method of preparation or sources.

Example : In water (H_2O), Hydrogen and Oxygen combine in 2 : 1 molar ratio, this ratio remains constant whether it is tap water, river water or sea water or produced by any chemical reaction.

Solved Examples

Example-6 1.80 g of a certain metal burnt in oxygen gave 3.0 g of its oxide. 1.50 g of the same metal heated in steam gave 2.50 g of its oxide. Show that these results illustrate the law of constant proportion.

Solution In the first sample of the oxide,

Wt. of metal = 1.80 g,
Wt. of oxygen = (3.0 – 1.80) g = 1.2 g

$$\therefore \quad \frac{\text{wt. of metal}}{\text{wt. of oxygen}} = \frac{1.80\text{g}}{1.2\text{g}} = 1.5$$

In the second sample of the oxide,

Wt. of metal = 1.50 g,
Wt. of oxygen = (2.50 – 1.50) g = 1 g.

$$\therefore \quad \frac{\text{wt. of metal}}{\text{wt. of oxygen}} = \frac{1.50\text{ g}}{1\text{g}} = 1.5$$

Thus, in both samples of the oxide the proportions of the weights of the metal and oxygen a fixed. Hence, the results follow the law of constant proportion.

(iii) The law of multiple proportion :

When one element combines with the other element to form two or more different compounds, the mass of one element, which combines with a constant mass of the other, bear a simple ratio to one another.

Note : Simple ratio here means the ratio between small natural numbers, such as 1 : 1, 1 : 2, 1 : 3, later on this simple ratio becomes the valency and then oxidation state of the element.
See oxidation number of carbon also have same ratio 1 : 2 in both the oxide.

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Solved Examples

Example-7 Carbon is found to form two oxides, which contain 42.9% and 27.3% of carbon respectively. Show that these figures illustrate the law of multiple proportions.

Solution **Step-1**

To calculate the percentage composition of carbon and oxygen in each of the two oxides.

	First oxide	Second oxide	
Carbon	42.9 %	27.3 %	(Given)
Oxygen (by difference)	57.1%	72.7 %	

Step-2

To calculate the masses of carbon which combine with a fixed mass i.e., one part by mass of oxygen in each of the two oxides.

In the first oxide, 57.1 parts by mass of oxygen combine with carbon = 42.9 parts.

∴ 1 part by mass of oxygen will combine with carbon = $\frac{42.9}{57.1}$ = 0.751.

In the second oxide. 72.7 parts by mass of oxygen combine with carbon = 27.3 parts.

∴ 1 part by mass of oxygen will combine with carbon = $\frac{27.3}{72.7}$ = 0.376

Step-3.

To compare the masses of carbon which combine with the same mass of oxygen in both the oxides.

The ratio of the masses of carbon that combine with the same mass of oxygen (1 part) is .

0.751 : 0.376 or 2 : 1

Since this is simple whole number ratio, so the above data illustrate the law of multiple proportions.

Percentage Composition :

Here we are going to find out the percentage of each element in the compound by knowing the molecular formula of compound.

We know that according to law of definite proportions any sample of a pure compound always possess constant ratio with their combining elements.

Solved Examples

Example-8 Every molecule of ammonia always has formula NH_3 irrespective of method of preparation or sources. i.e. 1 mole of ammonia always contains 1 mol of N and 3 mole of H. In other words 17 gm of NH_3 always contains 14 gm of N and 3 gm of H. Now find out % of each element in the compound.

Solution **Mass % of N in NH_3** $= \frac{\text{Mass of N in 1 mol } NH_3}{\text{Mass of 1 mol of } NH_3} \times 100 = \frac{14 \text{ gm}}{17} \times 100 = 82.35\ \%$

Mass % of H in NH_3 $= \frac{\text{Mass of H is 1 mol } NH_3}{\text{Mass of 1 mole of } NH_3} \times 100 = \frac{3}{17} \times 100 = 17.65\ \%$

Empirical and molecular formula :

We have just seen that knowing the molecular formula of the compound we can calculate percentage composition of the elements. Conversely if we know the percentage composition of the elements initially, we can calculate the relative number of atoms of each element in the molecules of the compound. This gives us the empirical formula of the compound. Further if the molecular mass is known then the molecular formula can easily be determined.

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The empirical formula of a compound is a chemical formula showing the relative number of atoms in the simplest ratio. An empirical formula represents the simplest whole number ratio of various atoms present in a compound. The molecular formula gives the actual number of atoms of each element in a molecule. The molecular formula shows the exact number of different types of atoms present in a molecule of a compound.

The molecular formula is an integral multiple of the empirical formula.

i.e. molecular formula = empirical formula × n

where $n = \dfrac{\text{molecular formula mass}}{\text{empirical formula mass}}$

Solved Examples

Example-9 Acetylene and benzene both have the empirical formula CH. The molecular masses of acetylene and benzene are 26 and 78 respectively. Deduce their molecular formulae.

Solution ∵ Empirical Formula is CH

Step-1 The empirical formula of the compound is CH

∴ Empirical formula mass = (1 × 12) + 1 = 13.

Molecular mass = 26

Step-2 To calculate the value of 'n'

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{26}{13} = 2$$

Step-3 To calculate the molecular formula of the compound.

Molecular formula = n × (Empirical formula of the compound)

= 2 × CH = C_2H_2

Thus the molecular formula is $\mathbf{C_2H_2}$

Similarly for benzene

To calculate the value of 'n'

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{78}{13} = 6$$

thus the molecular formula is 6 × CH = $\mathbf{C_6H_6}$

Example-10 An organic substance containing carbon, hydrogen and oxygen gave the following percentage composition.

C = 40.684% ; H = 5.085% and O = 54.228%

The molecular weight of the compound is 118 gm. Calculate the molecular formula of the compound.

Solution **Step-1**

To calculate the empirical formula of the compound.

Element	Symbol	Percentage of element	At. mass of element	Relative no. of atoms $= \frac{\text{Percentage}}{\text{At. mass}}$	Simplest atomic ratio	Simplest whole no. atomic ratio
Carbon	C	40.687	12	$\frac{40.687}{12} = 3.390$	$\frac{3.390}{3.389} = 1$	2
Hydrogen	H	5.085	1	$\frac{5.085}{1} = 5.085$	$\frac{5.085}{3.389} = 1.5$	3
Oxygen	O	54.228	16	$\frac{54.228}{16} = 3.389$	$\frac{3.389}{3.389} = 1$	2

∴ Empirical Formula is $C_2H_3O_2$

Step-2 To calculate the empirical formula mass.

The empirical formula of the compound is $C_2H_3O_2$.

∴ Empirical formula mass = (2 × 12) + (3 × 1) + (2 × 16) = 59.

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Step-3 To calculate the value of 'n'

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{118}{59} = 2$$

Step-4 To calculate the molecular formula of the salt.

Molecular formula = n × (Empirical formula) = 2 × $C_2H_3O_2$ = $C_4H_6O_4$

Thus the molecular formula is $C_4H_6O_4$.

Chemical Reaction :

It is the process in which two or more than two substances interact with each other where old bonds are broken and new bonds are formed.

Chemical Equation :

All chemical reaction are represented by chemical equations by using chemical formula of reactants and products. Qualitatively a chemical equation simply describes what the reactants and products are. However, a balanced chemical equation gives us a lot of quantitative information. Mainly the molar ratio in which reactants combine and the molar ratio in which products are formed.

Attributes of a balanced chemical equation:

(a) It contains an equal number of atoms of each element on both sides of equation.(POAC)
(b) It should follow law of charge conservation on either side.
(c) Physical states of all the reagents should be included in brackets.
(d) All reagents should be written in their standard molecular forms (not as atoms)
(e) The coefficients give the relative molar ratios of each reagent.

Solved Examples

Example-11 Write a balance chemical equation for following reaction :
When potassium chlorate ($KClO_3$) is heated it gives potassium chloride (KCl) and oxygen (O_2).

Solution $KClO_3$ (s) $\xrightarrow{\Delta}$ KCl (s) + O_2 (g) (unbalanced chemical equation)

$2KClO_3$ (s) $\xrightarrow{\Delta}$ 2 KCl (s) + 3 O_2 (g) (balanced chemical equation)

Remember a balanced chemical equation is one which contains an equal number of atoms of each element on both sides of equation.

Interpretation of balanced chemical equations :

Once we get a balanced chemical equation then we can interpret a chemical equation by following ways

- Mass - mass analysis
- Mass - volume analysis
- Mole - mole analysis
- Vol - Vol analysis (separately discussed as **eudiometry or gas analysis**)

Now you can understand the above analysis by following example

- **Mass-mass analysis :**

Consider the reaction

$2KClO_3 \longrightarrow 2KCl + 3O_2$ According to stoichiometry of the reaction

mass-mass ratio: 2 × 122.5 : 2 × 74.5 : 3 × 32

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or $$\frac{\text{Mass of } KClO_3}{\text{Mass of KCl}} = \frac{2\times122.5}{2\times74.5}$$

$$\frac{\text{Mass of } KClO_3}{\text{Mass of } O_2} = \frac{2\times122.5}{3\times32}$$

Solved Examples

Example-12 367.5 gram $KClO_3$ (M = 122.5) when heated. How many gram KCl and oxygen is produced.

Solution Balance chemical equation for heating of $KClO_3$ is

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$

mass-mass ratio : 2 × 122.5 gm : 2 × 74.5 gm : 3 × 32 gm

$$\frac{\text{mass of } KClO_3}{\text{mass of KCl}} = \frac{2\times122.5}{2\times74.5} \Rightarrow \frac{367.5}{W} = \frac{122.5}{74.5}$$

W = 3 × 74.5 = 223.5 gm

$$\frac{\text{Mass of } KClO_3}{\text{Mass of } O_2} = \frac{2\times122.5}{3\times32} \Rightarrow \frac{367.5}{W} = \frac{2\times122.5}{3\times32}$$

W = 144 gm

- **Mass - volume analysis :**

Now again consider decomposition of $KClO_3$

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$

mass volume ratio : 2 × 122.5 gm : 2 × 74.5 gm : 3 × 22.4 lt. at STP

we can use two relation for volume of oxygen

$$\frac{\text{Mass of } KClO_3}{\text{volume of } O_2 \text{ at STP}} = \frac{2\times122.5}{3\times22.4 \text{ lt}} \qquad ...(i)$$

and $$\frac{\text{Mass of KCl}}{\text{volume of } O_2 \text{ at STP}} = \frac{2\times74.5}{3\times22.4 \text{ lt}} \qquad ...(ii)$$

Solved Examples

Example-13 367.5 gm $KClO_3$ (M = 122.5) when heated, how many litre of oxygen gas is produced at STP.

Solution You can use here equation (1)

$$\frac{\text{mass of } KClO_3}{\text{volume of } O_2 \text{ at STP}} = \frac{2\times122.5}{3\times22.4 \text{ lt}} \Rightarrow \frac{367.5}{V} = \frac{2\times122.5}{3\times22.4 \text{ lt}}$$

V = 3 × 3 × 11.2 ⇒ **V = 100.8 lt**

- **Mole-mole analysis :**

This analysis is very much important for quantitative analysis point of view. Students are advised to clearly understand this analysis.

Now consider again the decomposition of $KClO_3$.

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$

In very first step of mole-mole analysis you should read the balanced chemical equation like ***2 moles $KClO_3$ on decomposition gives you 2 moles KCl and 3 moles O_2.*** and from the stoichiometry of reaction we can write

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$$\frac{\text{Moles of } KClO_3}{2} = \frac{\text{Moles of } KCl}{2} = \frac{\text{Moles of } O_2}{3}$$

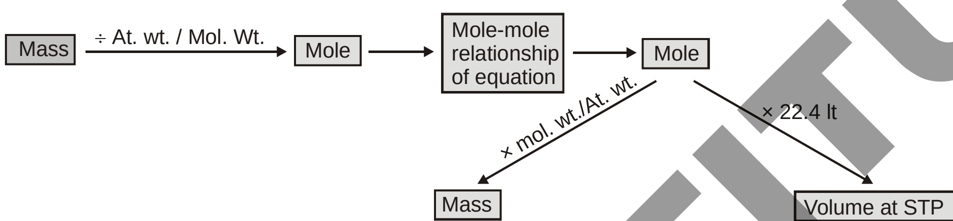
Now for any general balance chemical equation like

$$aA + bB \longrightarrow cC + dD$$

you can write.

$$\frac{\text{Moles of A reacted}}{a} = \frac{\text{moles of B reacted}}{b} = \frac{\text{moles of C produced}}{c} = \frac{\text{moles of D produced}}{d}$$

Note : In fact mass-mass and mass-vol analysis are also interpreted in terms of mole-mole analysis you can use following chart also.



Limiting reagent :

The reactant which is consumed first and limits the amount of product formed in the reaction, and is therefore, called limiting reagent.

Limiting reagent is present in least stoichiometric amount and therefore, controls amount of product.

The remaining or left out reactant is called the excess reagent.

When you are dealing with balance chemical equation then if number of moles of reactants are not in the ratio of stoichiometric coefficient of balanced chemical equation, then there should be one reactant which is limiting reactant.

Solved Examples

Example-14 Three mole of Na_2CO_3 is reacted with 6 moles of HCl solution. Find the volume of CO_2 gas produced at STP. The reaction is

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

Solution

From the reaction :	$Na_2CO_3 + 2HCl \longrightarrow 2\,NaCl + CO_2 + H_2O$
given moles	3 mol 6 mol
given mole ratio	1 : 2
Stoichiometric coefficient ratio	1 : 2

See here given moles of reactant are in stoichiometric coefficient ratio therefore none reactant left over.

Now use Mole-mole analysis to calculate volume of CO_2 produced at STP

$$\frac{\text{Moles of } Na_2CO_3}{1} = \frac{\text{Mole of } CO_2 \text{ Produced}}{1}$$

Moles of CO_2 produced = 3

volume of CO_2 produced at STP = 3 × 22.4 L = 67.2 L

Example-15 6 moles of Na_2CO_3 is reacted with 4 moles of HCl solution. Find the volume of CO_2 gas produced at STP. The reaction is

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

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Solution From the reaction : $Na_2CO_3 + 2HCl \longrightarrow 2\,NaCl + CO_2 + H_2O$

given mole of reactant	6	:	4
given molar ratio	3	:	2
Stoichiometric coefficient ratio	1	:	2

See here given number of moles of reactants are not in stoichiometric coefficient ratio. Therefore there should be one reactant which consumed first and becomes limiting reagent.
But the question is how to find which reactant is limiting, it is not very difficult you can easily find it. According to the following method.

How to find limiting reagent :

Step :-I Divide the given moles of reactant by the respective stoichiometric coefficient of that reactant.
Step :-II See for which reactant this division come out to be minimum. The reactant having minimum value is limiting reagent for you.
Step :-III Now once you find limiting reagent then your focus should be on limiting reagent

From Step I & II Na_2CO_3 HCl

$\frac{6}{1} = 6$ $\frac{4}{2} = 2$ (division is minimum)

$\therefore$ **HCl is limiting reagent**

From Step III

From $\frac{\text{Mole of HCl}}{2} = \frac{\text{Moles of } CO_2 \text{ produced}}{1}$

$\therefore$ mole of CO_2 produced = 2 moles

$\therefore$ volume of CO_2 produced at S.T.P. = 2 × 22.4 = 44.8 lt.

Principle of Atom Conservation (POAC) :

POAC is conservation of mass. Atoms are conserved, moles of atoms shall also be conserved in a chemical reaction (but not in nuclear reactions.)

This principle is fruitful for the students when they don't get the idea of balanced chemical equation in the problem.

The strategy here will be around a particular atom. We focus on a atom and conserve it in that reaction. This principle can be understand by the following example.

Consider the decomposition of $KClO_3$ (s) $\rightarrow$ KCl (s) + O_2 (g) (unbalanced chemical reaction)
Apply the principle of atom conservation (POAC) for K atoms.
Moles of K atoms in reactant = moles of K atoms in products
or moles of K atoms in $KClO_3$ = moles of K atoms in KCl.
Now, since 1 molecule of $KClO_3$ contains 1 atom of K
or 1 mole of $KClO_3$ contains 1 mole of K, similarly,1 mole of KCl contains 1 mole of K.
Thus, moles of K atoms in $KClO_3$ = 1 × moles of $KClO_3$
and moles of K atoms in KCl = 1 × moles of KCl.

$\therefore$ moles of $KClO_3$ = moles of KCl

or $$\frac{\textbf{wt. of } KClO_3 \textbf{ in g}}{\textbf{mol. wt. of } KClO_3} = \frac{\textbf{wt. of KCl in g}}{\textbf{mol. wt. of KCl}}$$

- The above equation gives the mass-mass relationship between $KClO_3$ and KCl which is important in stoichiometric calculations.

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Again, applying the principle of atom conservation for O atoms,
moles of O in $KClO_3$ = 3 × moles of $KClO_3$
moles of O in O_2 = 2 × moles of O_2
$\therefore$ 3 × moles of $KClO_3$ = 2 × moles of O_2

or $$3 \times \frac{\text{wt. of } KClO_3}{\text{mol. wt. of } KClO_3} = 2 \times \frac{\text{vol. of } O_2 \text{ at NTP}}{\text{standard molar vol. (22.4 lt.)}}$$

- The above equations thus gives the mass-volume relationship of reactants and products.

Solved Examples

Example-16 27.6 g K_2CO_3 was treated by a series of reagents so as to convert all of its carbon to $K_2Zn_3[Fe(CN)_6]_2$. Calculate the weight of the product.
[mol. wt. of K_2CO_3 = 138 and mol. wt. of $K_2Zn_3[Fe(CN)_6]_2$ = 698]

Solution Here we have not knowledge about series of chemical reactions but we know about initial reactant and final product accordingly

$$K_2CO_3 \xrightarrow[\text{Steps}]{\text{Several}} K_2Zn_3[Fe(CN)_6]_2$$

Since C atoms are conserved, applying POAC for C atoms,
moles of C in K_2CO_3 = moles of C in $K_2Zn_3[Fe(CN)_6]_2$
1 × moles of K_2CO_3 = 12 × moles of $K_2Zn_3[Fe(CN)_6]_2$
($\because$ 1 mole of K_2CO_3 contains 1 moles of C)

$$\frac{\text{wt. of } K_2CO_3}{\text{mol. wt. of } K_2CO_3} = 12 \times \frac{\text{wt. of the product}}{\text{mol. wt. of product}}$$

wt. of $K_2Zn_3[Fe(CN)_6]_2$ = $\frac{27.6}{138} \times \frac{698}{12}$ = 11.6 g

Miscellaneous :

- **AVERAGE/ MEAN ATOMIC MASS :**

The weighted average of the isotopic masses of the element's naturally occuring isotopes.

Mathematically, average atomic mass of X (A_x) = $\frac{a_1x_1 + a_2x_2 + ... + a_nx_n}{100}$

Where :
a_1, a_2, a_3 .... atomic mass of isotopes.
and x_1, x_2, x_3 .... mole % of isotopes.

Solved Examples

Example-17 Naturally occuring chlorine is 75% Cl^{35} which has an atomic mass of 35 amu and 25% Cl^{37} which has a mass of 37 amu. Calculate the average atomic mass of chlorine -
(A) 35.5 amu (B) 36.5 amu (C) 71 amu (D) 72 amu

Solution (A) Average atomic mass = $\frac{\% \text{ of I isotope x its atoms mass} + \% \text{ of II isotope x its atomic mass}}{100}$

$$= \frac{75 \times 35 + 25 \times 37}{100} = 35.5 \text{ amu}$$

Note : **(a) In all calculations we use this mass. (b) In periodic table we report this mass only.**

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- **MEAN MOLAR MASS OR MOLECULAR MASS:**

The average molar mass of the different substance present in the container = $\dfrac{n_1M_1 + n_2M_2 + ...n_nM_n}{n_1 + n_2 + ...n_n}$.

Where :

M_1, M_2, M_3 .... are molar masses.
n_1, n_2, n_3 .... moles of substances.

Solved Examples

Example-18 The molar composition of polluted air is as follows :

Gas	At. wt.	mole percentage composition
Oxygen	16	16%
Nitrogen	14	80%
Carbon dioxide	-	03%
Sulphurdioxide	-	01%

What is the average molecular weight of the given polluted air ? (Given, atomic weights of C and S are 12 and 32 respectively.

Solution $M_{avg} = \dfrac{\sum_{j=1}^{j=n} n_jM_j}{\sum_{j=1}^{j=n} n_j}$ Here $\sum_{j=1}^{j=n} n_j = 100$

$\therefore M_{avg} = \dfrac{16 \text{ x } 32 + 80 \text{ x } 28 + 44 \text{ x } 3 + 64 \text{ x } 1}{100} = \dfrac{512 + 2240 + 132 + 64}{100} = \dfrac{2948}{100} = \mathbf{29.48}$ **Ans.**

Oxidation & Reduction

Let us do a comparative study of oxidation and reduction :

Oxidation	Reduction
1. Addition of Oxygen e.g. $2Mg + O_2 \rightarrow 2MgO$	1. Removal of Oxygen e.g. $CuO + C \rightarrow Cu + CO$
2. Removal of Hydrogen e.g. $H_2S + Cl_2 \rightarrow 2HCl + S$	2. Addition of Hydrogen e.g. $S + H_2 \rightarrow H_2S$
3. Increase in positive charge e.g. $Fe^{2+} \rightarrow Fe^{3+} + e^-$	3. Decrease in positive charge e.g. $Fe^{3+} + e^- \rightarrow Fe^{2+}$
4. Increase in oxidation number e.g. $\overset{(+2)}{SnCl_2} \rightarrow \overset{(+4)}{SnCl_4}$	4. Decrease in oxidation number e.g. $\overset{(+7)}{MnO_4^-} \rightarrow \overset{(+2)}{Mn^{2+}}$
5. Removal of electron e.g. $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$	5. Addition of electron e.g. $Fe^{3+} + e^- \rightarrow Fe^{2+}$

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Oxidation Number

- It is an imaginary or apparent charge developed over atom of an element when it goes from its elemental free state to combined state in molecules.
- It is calculated on basis of an arbitrary set of rules.
- It is a relative charge in a particular bonded state.
- In order to keep track of electron-shifts in chemical reactions involving formation of compounds, a more practical method of using oxidation number has been developed.
- In this method, it is always assumed that there is a complete transfer of electron from a less electronegative atom to a more electronegative atom.

Rules governing oxidation number

The following rules are helpful in calculating oxidation number of the elements in their different compounds. It is to be remembered that the basis of these rule is the electronegativity of the element .

- **Fluorine atom :**
  Fluorine is most electronegative atom (known). It always has oxidation number equal to –1 in all its compounds
- **Oxygen atom :**
  In general and as well as in its oxides , oxygen atom has oxidation number equal to –2.

In case of (i) peroxide (e.g. H_2O_2 , Na_2O_2) is –1,
(ii) super oxide (e.g. KO_2) is –1/2
(iii) ozonide (e.g. KO_3) is –1/3
(iv) in OF_2 is + 2 & in O_2F_2 is +1

- **Hydrogen atom :**
  In general, H atom has oxidation number equal to +1. But in metallic hydrides (e.g. NaH, KH), it is –1.
- **Halogen atom :**
  In general, all halogen atoms (Cl, Br , I) have oxidation number equal to –1.
  But if halogen atom is attached with a more electronegative atom than halogen atom, then it will show positive oxidation numbers.

  e.g. $K\overset{+5}{Cl}O_3$, $H\overset{+5}{I}O_3$, $H\overset{+7}{Cl}O_4$, $K\overset{+5}{Br}O_3$
- **Metals :**
  (a) Alkali metal (Li , Na, K, Rb, ...) always have oxidation number +1
  (b) Alkaline earth metal (Be , Mg , Ca ....) always have oxidation number +2.
  (c) Aluminium always has +3 oxidation number

Note : Metal may have negative or zero oxidation number

- Oxidation number of an element in free state or in allotropic forms is always zero

  e.g. $\overset{0}{O}_2$, $\overset{0}{S}_8$, $\overset{0}{P}_4$,$\overset{0}{O}_3$
- Sum of the oxidation numbers of atoms of all elements in a molecule is zero.
- Sum of the oxidation numbers of atoms of all elements in an ion is equal to the charge on the ion .
- If the group number of an element in modern periodic table is ***n,*** then its oxidation number may vary from ***(n – 10)*** to ***(n – 18)*** (but it is mainly applicable for p-block elements)
  e.g. N- atom belongs to 15th group in the periodic table, therefore as per rule, its oxidation number may vary from

  –3 to +5 ($\overset{-3}{N}H_3$, $\overset{+2}{N}O$, $\overset{+3}{N}_2O_3$, $\overset{+4}{N}O_2$, $\overset{+5}{N}_2O_5$)
- The maximum possible oxidation number of any element in a compound is never more than the number of electrons in valence shell.(but it is mainly applicable for p-block elements)

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Calculation of average oxidation number :

Solved Examples

Example-19 ***Calculate oxidation number of underlined element*** **:**

(a) $Na_2\underline{S}_2O_3$ **(b)** $Na_2\underline{S}_4O_6$

Solution. (a) Let oxidation number of S-atom is x. Now work accordingly with the rules given before .

$(+1) \times 2 + (x) \times 2 + (-2) \times 3 = 0$

$x = +2$

(b) Let oxidation number of S-atom is x

$\therefore \quad (+1) \times 2 + (x) \times 4 + (-2) \times 6 = 0$

$x = +2.5$

❍ It is important to note here that $Na_2S_2O_3$ have two S-atoms and there are four S-atom in $Na_2S_4O_6$. However none of the sulphur atoms in both the compounds have + 2 or + 2.5 oxidation number, it is the average of oxidation number, which reside on each sulphur atom. Therefore, we should work to calculate the individual oxidation number of each sulphur atom in these compounds.

Calculation of individual oxidation number

It is important to note that to calculate individual oxidation number of the element in its compound one should know the structure of the compound and use the following guidelines.

Formula :

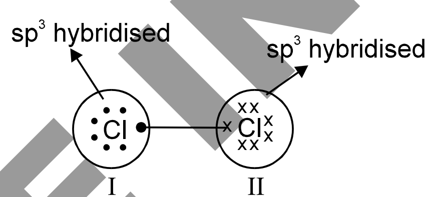
Oxidation Number = Number of electrons in the valence shell – Number of electrons taken up after bonding

Guidelines : It is based on electronegativity of elements.

1. If there is a bond between similar type of atom and each atom has same type of hybridisation, then bonded pair electrons are equally shared by each element.

Example : Calculate oxidation number of each Cl-atom in Cl_2 molecule

Structure :



I : Number of electrons in the valence shell = 7.

Number of electrons taken up after bonding = 7.

$\therefore$ oxidation number = 7 – 7 = 0.

II : similarly, oxidation number = 7 – 7 = 0

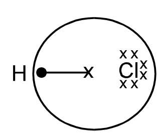
2. If there is a bond between different type of atoms :

e.g. A – B (if B is more electronegative than A)

Then after bonding, bonded pair of electrons are counted with B - atom .

Example : Calculate oxidation number of each atom in HCl molecule

Structure :



Note : Electron of H-atom is now counted with Cl-atom, because Cl-atom is more electronegative than H-atom

H : Number of electrons in the valence shell = 1

Number of electrons taken up after bonding = 0

Oxidation number of H = 1 – 0 = + 1

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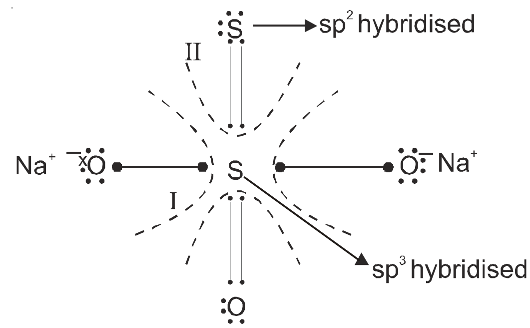
Cl : Number of electrons in the valence shell = 7
Number of electrons taken up after bonding = 8
Oxidation number of Cl = 7– 8 = – 1

Solved Examples

Example-20 Calculate individual oxidation number of each S-atom in $Na_2S_2O_3$ (sodium thiosulphate) with the help of its structure .

Solution. ***Structure :***



Note : I ***(central S-atom) is sp³ hybridised (25% s-character) and*** II ***(terminal S-atom) is sp² hydbridised (33% s-character). Therefore, terminal sulphur atom is more electronegative than central sulphur atom. Now, the shared pair of electrons are counted with terminal S-atom.***

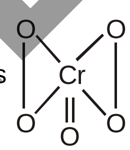
$\therefore$ I, S-atom : Number of electrons in the valence shell = 6
Number of electrons left after bonding = 0
Oxidation number of central S-atom = 6 – 0 = + 6

II, S-atom : Number of electrons in the valence shell = 6
Number of electrons left after bonding = 8
Oxidation number of terminal S-atom = 6 – 8 = – 2

Now, you can also calculate Average Oxidation number of S = $\frac{6+(-2)}{2}$ = + 2 (as we have calculated before)

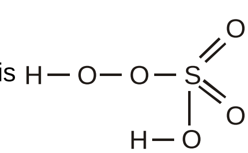
Miscellaneous Examples :

In order to determine the exact or individual oxidation number we need to take help from the structures of the molecules. Some special cases are discussed as follows:

- The structure of CrO_5 is 

From the structure, it is evident that in CrO_5 there are two peroxide linkages and one double bond. The contribution of each peroxide linkage is –2. Let the oxidation number of Cr is x.

$\therefore$ x + (–2)2 + (–2) = 0 or x = 6
$\therefore$ Oxidation number of Cr = + 6 **Ans**

- The structure of H_2SO_5 is 

From the structure, it is evident that in H_2SO_5, there is one peroxide linkage, two sulphur-oxygen double bonds and one OH group. Let the oxidation number of S = x.

$\therefore$ (+ 1) + (– 2) + x + (–2) 2+ (–2) + 1 = 0
or x + 2 – 8 = 0 or x – 6 = 0 or x = 6
$\therefore$ Oxidation number of S in H_2SO_5 is + 6 **Ans.**

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Paradox of fractional oxidation number

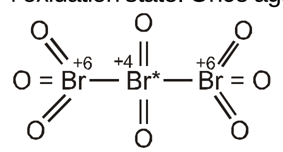
Fractional oxidation number is the average of oxidation state of all atoms of element under examination and the structural parameters reveal that the atoms of element for whom fractional oxidation state is realised a actually present in different oxidation states. Structure of the species C_3O_2, Br_3O_8 and $S_4O_6^{2-}$ reveal the following bonding situations :

- The element marked with asterisk (*) in each species is exhibiting different oxidation number from rest of the atoms of the same element in each of the species. This reveals that in C_3O_2, two carbon atoms are present in +2 oxidation state each whereas the third one is present in zero oxidation state and the average is + 4/3. However, the realistic picture is +2 for two terminal carbons and zero for the middle carbon.

$$O = \overset{+2}{C} = \overset{0}{C}{*} = \overset{+2}{C} = O$$

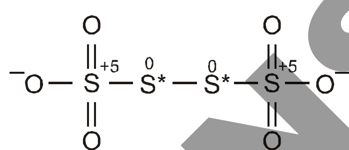
Structure of C_3O_2
(Carbon suboxide)

- Likewise in Br_3O_8, each of the two terminal bromine atoms are present in +6 oxidation state and the middle bromine is present in +4 oxidation state. Once again the average, that is different from reality, is + 16/3.



Structure of Br_3O_8 (Tribromooctaoxide)

- In the same fashion, in the species $S_4O_6^{2-}$, average oxidation number of S is + 2.5, whereas the reality being +5,0,0 and +5 oxidation number respectively for respective sulphur atoms.



Structure of $S_4O_6^{2-}$ (tetrathionate ion)

In general, the conclusion is that the idea of fractional oxidation state should be taken with care and the reality is revealed by the structures only.

Oxidising and reducing agent

- **Oxidising agent or Oxidant :**

  Oxidising agents are those compounds which can oxidise others and reduce itself during the chemical reaction. Those reagents in which for an element, oxidation number decreases or which undergoes gain of electrons in a redox reaction are termed as oxidants.

  e.g. $KMnO_4$, $K_2Cr_2O_7$, HNO_3, conc.H_2SO_4 etc are powerful oxidising agents .

- **Reducing agent or Reductant :**

  Reducing agents are those compounds which can reduce other and oxidise itself during the chemical reaction. Those reagents in which for an element, oxidation number increases or which undergoes loss of electrons in a redox reaction are termed as reductants.

  e.g. KI , $Na_2S_2O_3$ etc are the powerful reducing agents.

Note : ***There are some compounds also which can work both as oxidising agent and reducing agent***
e.g. H_2O_2, NO_2^-

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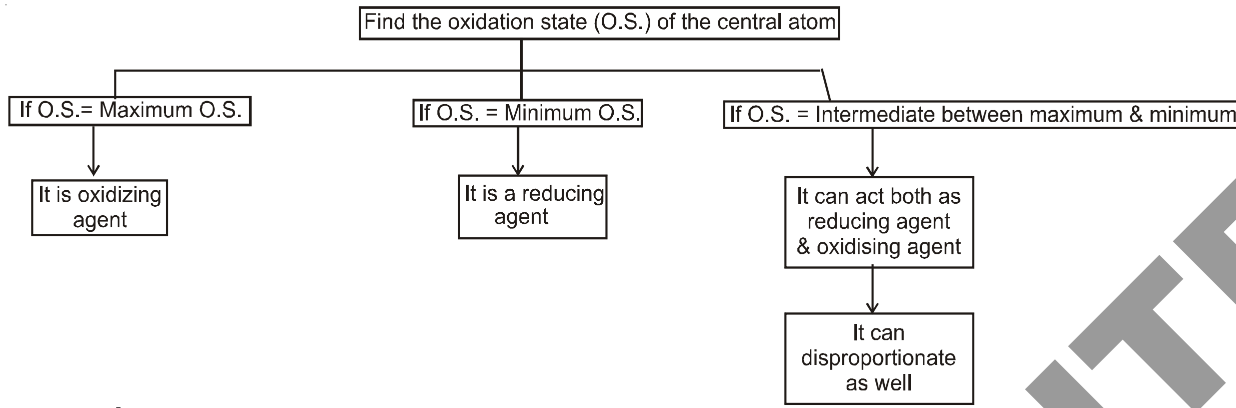
HOW TO IDENTIFY WHETHER A PARTICULAR SUBSTANCE IS AN OXIDISING OR A REDUCING AGENT



Redox reaction

A reaction in which oxidation and reduction simultaneously take place is called a redox reaction
In all redox reactions, the total increase in oxidation number must be equal to the total decrease in oxidation number.

e.g. $10\overset{+2}{Fe}SO_4 + 2K\overset{+5}{Mn}O_4 + 8H_2SO_4 \longrightarrow 5\overset{+3}{Fe_2}(SO_4)_3 + 2\overset{+2}{Mn}SO_4 + K_2SO_4 + 8H_2O$

Disproportionation Reaction :

A redox reaction in which same element present in a particular compound in a definite oxidation state is oxidized as well as reduced simultaneously is a disproportionation reaction.
Disproportionation reactions are a special type of redox reactions. One of the reactants in a disproportionation reaction always contains **an element that can exist in at least three oxidation states**. The element in the form of reacting substance is in the intermediate oxidation state and both higher and lower oxidation states of that element are formed in the reaction. For example :

$$2H_2\overset{-1}{O_2}(aq) \longrightarrow 2H_2\overset{-2}{O}(\ell) + \overset{0}{O_2}(g)$$

$$\overset{0}{S_8}(s) + 12OH^-(aq) \longrightarrow 4\overset{-2}{S^{2-}}(aq) + 2\overset{+2}{S_2}O_3^{2-}(aq) + 6H_2O(\ell)$$

$$\overset{0}{Cl_2}(g) + 2OH^-(aq) \longrightarrow \overset{+1}{Cl}O^-(aq) + \overset{-1}{Cl^-}(aq) + H_2O(\ell)$$

Consider the following reactions :

(a) $2KClO_3 \longrightarrow 2KCl + 3O_2$
$KClO_3$ plays a role of oxidant and reductant both. Here, Cl present in $KClO_3$ is reduced and O present in $KClO_3$ is oxidized. Since same element is not oxidized and reduced, so **it is not a disproportionation reaction,** although it looks like one.

(b) $NH_4NO_2 \longrightarrow N_2 + 2H_2O$
Nitrogen in this compound has -3 and +3 oxidation number, which is not a definite value. So it is not a disproportionation reaction. It is an example of comproportionation reaction, which is a class of redox reaction in which an element from two different oxidation state gets converted into a single oxidation state.

(c) $4K\overset{+5}{Cl}O_3 \longrightarrow 3K\overset{+7}{Cl}O_4 + K\overset{-1}{Cl}$

It is a case of disproportionation reaction and Cl atom is disproportionating.

List of some important disproportionation reactions

1. $H_2O_2 \longrightarrow H_2O + O_2$
2. $X_2 + OH^-(dil.) \longrightarrow X^- + XO^-$ (X = Cl, Br, I)
3. $X_2 + OH^-(conc.) \longrightarrow X^- + XO_3^-$

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F_2 does not undergo disproportionation as it is the most electronegative element.

$F_2 + NaOH(dil.) \longrightarrow F^- + OF_2$

$F_2 + NaOH(conc.) \longrightarrow F^- + O_2$

4. $(CN)_2 + OH^- \longrightarrow CN^- + OCN^-$

5. $P_4 + OH^- \longrightarrow PH_3 + H_2PO_2^-$

6. $S_8 + OH^- \longrightarrow S^{2-} + S_2O_3^{2-}$

7. $MnO_4^{2-} \longrightarrow MnO_4^- + MnO_2$

8. $NH_2OH \longrightarrow N_2O + NH_3$

$NH_2OH \longrightarrow N_2 + NH_3$

9. Oxyacids of Phosphorus (+1, +3 oxidation number)

$H_3PO_2 \longrightarrow PH_3 + H_3PO_3$

$H_3PO_3 \longrightarrow PH_3 + H_3PO_4$

10. Oxyacids of Chlorine(Halogens)(+1, +3, +5 Oxidation number)

$ClO^- \longrightarrow Cl^- + ClO_2^-$

$ClO_2^- \longrightarrow Cl^- + ClO_3^-$

$ClO_3^- \longrightarrow Cl^- + ClO_4^-$

11. $HNO_2 \longrightarrow NO + HNO_3$

- Reverse of disproportionation is called **Comproportionation.** In some of the disproportionation reactions, by changing the medium (from acidic to basic or reverse), the reaction goes in backward direction and can be taken as an example of **Comproportionation reaction.**

$$I^- + IO_3^- + H^+ \longrightarrow I_2 + H_2O$$

Balancing of redox reactions

All balanced equations must satisfy two criteria.

1. Atom balance (mass balance) :
There should be the same number of atoms of each kind on reactant and product side.

2. Charge balance :
The sum of actual charges on both sides of the equation must be equal.
There are two methods for balancing the redox equations :
1. Oxidation - number change method
2. Ion electron method or half cell method

❍ Since First method is not very much fruitful for the balancing of redox reactions, students are advised to use second method (Ion electron method) to balance the redox reactions

Ion electron method : By this method redox equations are balanced in two different medium.
(a) Acidic medium (b) Basic medium

- **Balancing in acidic medium**
Students are adviced to follow the following steps to balance the redox reactions by Ion electron method in acidic medium

Solved Examples

Example-21 Balance the following redox reaction :

$FeSO_4 + KMnO_4 + H_2SO_4 \longrightarrow Fe_2(SO_4)_3 + MnSO_4 + H_2O + K_2SO_4$

Solution. **Step–f** Assign the oxidation number to each element present in the reaction.

$$\overset{+2}{Fe}\,\overset{+6}{S}\,\overset{-2}{O_4} + \overset{+1}{K}\,\overset{+7}{Mn}\,\overset{-2}{O_4} + \overset{+1}{H_2}\,\overset{+6}{S}\,\overset{-2}{O_4} \longrightarrow \overset{+3}{Fe_2}(\overset{+6}{S}\,\overset{-2}{O_4})_3 + \overset{+2}{Mn}\,\overset{+6}{S}\,\overset{-2}{O_4} + \overset{+1}{H_2}\overset{-2}{O}$$

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Step ff :
Now convert the reaction in Ionic form by eliminating the elements or species, which are not undergoing either oxidation or reduction.

$$Fe^{2+} + \overset{+7}{Mn}O_4^- \longrightarrow Fe^{3+} + Mn^{2+}$$

Step fff :
Now identify the oxidation / reduction occuring in the reaction

undergoes reduction.

$$Fe^{2+} + MnO_4^- \rightarrow Fe^{3+} + Mn^{2+}$$

undergoes oxidation.

Step f**V :** Spilt the Ionic reaction in two half, one for oxidation and other for reduction.

$$Fe^{2+} \xrightarrow{\text{oxidation}} Fe^{3+} \quad \Big| \quad MnO_4^- \xrightarrow{\text{Reduction}} Mn^{2+}$$

Step V :
Balance the atom other than oxygen and hydrogen atom in both half reactions

$$Fe^{2+} \longrightarrow Fe^{3+} \quad \Big| \quad MnO_4^- \longrightarrow Mn^{2+}$$

Fe & Mn atoms are balanced on both side.

Step Vf **:**
Now balance O & H atom by H_2O & H^+ respectively by the following way : For one excess oxygen atom, add one H_2O on the other side and two H^+ on the same side.

$Fe^{2+} \longrightarrow Fe^{3+}$ (no oxygen atom) ....(i)

$8H^+ + MnO_4^- \longrightarrow Mn^{2+} + 4H_2O$ ....(ii)

Step Vff **:**
Equation (i) & (ii) are balanced atomwise. Now balance both equations chargewise. To balance the charge, add electrons to the electrically positive side.

$Fe^{2+} \xrightarrow{\text{oxidation}} Fe^{3+} + e^-$ ....(1)

$5e^- + 8H^+ + MnO_4^- \xrightarrow{\text{Reduction}} Mn^{2+} + 4H_2O$ ....(2)

Step Vfff **:**
The number of electrons gained and lost in each half -reaction are equalised by multiplying both the half reactions with a suitable factor and finally the half reactions are added to give the overall balanced reaction.
Here, we multiply equation (1) by 5 and (2) by 1 and add them :

$Fe^{2+} \longrightarrow Fe^{3+} + e^-$ ....(1) × 5

$5e^- + 8H^+ + MnO_4^- \longrightarrow Mn^{2+} + 4H_2O$ ....(2)×1

$5Fe^{2+} + 8H^+ + MnO_4^- \longrightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$

(Here, at his stage, you will get balanced redox reaction in Ionic form)

Step f**X :**
Now convert the Ionic reaction into molecular form by adding the elements or species, which are removed in step (2).
Now, by some manipulation, you will get :

$$5\,FeSO_4 + KMnO_4 + 4H_2SO_4 \longrightarrow \frac{5}{2}\,Fe_2(SO_4)_3 + MnSO_4 + 4H_2O + \frac{1}{2}\,K_2SO_4 \quad \text{or}$$

$$10FeSO_4 + 2KMnO_4 + 8H_2SO_4 \longrightarrow 5Fe_2(SO_4)_3 + 2MnSO_4 + 8H_2O + K_2SO_4.$$

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- **Balancing in basic medium :**
  In this case, except step VI, all the steps are same. We can understand it by the following example:

Solved Examples

Example-22 Balance the following redox reaction in basic medium :

$$^{-} + CrO_2^- + OH^- \longrightarrow Cl^- + CrO_4^{2-} + H_2O$$

Solution. By using upto step V, we will get :

$$\overset{+1}{Cl}O^- \xrightarrow{\text{Reduction}} Cl^- \quad \Big| \quad \overset{+3}{Cr}O_2^- \xrightarrow{\text{Oxidation}} \overset{+6}{Cr}O_4^{2-}$$

Now, students are advised to follow step VI to balance 'O' and 'H' atom.

$2H^+ + ClO^- \longrightarrow Cl^- + H_2O$ | $2H_2O + CrO_2^- \longrightarrow CrO_4^{2-} + 4H^+$

- Now, since we are balancing in basic medium, therefore add as many as OH^- on both side of equation as there are H^+ ions in the equation.

$2OH^- + 2H^+ + ClO^- \longrightarrow Cl^- + H_2O + 2OH^-$ | $4OH^- + 2H_2O + CrO_2^- \longrightarrow CrO_4^{2-} + 4H^+ + 4OH^-$

Finally you will get | Finally you will get

$H_2O + ClO^- \longrightarrow Cl^- + 2OH^-$ ....(i) | $4OH^- + CrO_2^- \longrightarrow CrO_4^{2-} + 2H_2O$ .... (ii)

Now see equation (i) and (ii) in which O and H atoms are balanced by OH^- and H_2O

Now from step VIII

$2e^- + H_2O + ClO^- \longrightarrow Cl^- + 2OH^-$ .... (i) ×3

$4OH^- + CrO_2^- \longrightarrow CrO_4^{2-} + 2H_2O + 3e^-$ .... (ii) ×2

Adding : $3ClO^- + 2CrO_2^- + 2OH^- \longrightarrow 3Cl^- + 2CrO_4^{2-} + H_2O$

Solutions :

A mixture of two or more substances can be a solution. We can also say that "a solution is a homogeneous mixture of two or more substances,'' 'Homogeneous' means 'uniform throughout'. Thus a homogeneous mixture, i.e., a solution, will have uniform composition throughout.

Properties of a solution :

- A solution is clear and transparent. For example, a solution of sodium chloride in water is clear and tranparent.
- The solute in a solution does not settle down even after the solution is kept undisturbed for some time.
- In a solution, the solute particle cannot be distinguished from the solvent particles or molecules even under a microscope. In a true solution, the particles of the solute disappear into the space between the solvent molecules.
- The components of a solution cannot be separated by filtration.

Concentration terms :

The following concentration terms are used to expressed the concentration of a solution. These are

- Molarity (M)
- Molality (m)
- Mole fraction (x)
- % calculation
- Normality (N)
- ppm

- Remember that all of these concentration terms are related to one another. By knowing one concentration term you can also find the other concentration terms. Let us discuss all of them one by one.

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Molarity (M) :

The number of moles of a solute dissolved in 1 L (1000 ml) of the solution is known as the molarity of the solution.

i.e., $\text{Molarity of solution} = \dfrac{\text{number of moles of solute}}{\text{volume of solution in litre}}$

Let a solution is prepared by dissolving w gm of solute of mol.wt. M in V ml water.

$\therefore$ Number of moles of solute dissolved $= \dfrac{w}{M}$

$\therefore$ V ml water have $\dfrac{w}{M}$ mole of solute

$\therefore$ 1000 ml water have $\dfrac{w \times 1000}{M \times V_{ml}}$ $\therefore$ $\text{Molarity (M)} = \dfrac{w \times 1000}{(\text{Mol. wt of solute}) \times V_{ml}}$

Some other relations may also useful.

$\text{Number of millimoles} = \dfrac{\text{mass of solute}}{(\text{Mol. wt. of solute})} \times 1000 = (\text{Molarity of solution} \times V_{ml})$

- Molarity of solution may also given as :

  $\dfrac{\text{Number of millimole of solute}}{\text{Total volume of solution in ml}}$

- Molarity is a unit that depends upon temperature. It varies inversely with temperature .

  Mathematically : Molarity decreases as temperature increases.

  $\text{Molarity} \propto \dfrac{1}{\text{temperature}} \propto \dfrac{1}{\text{volume}}$

- If a particular solution having volume V_1 and molarity = M_1 is diluted upto volume V_2 mL than

  $M_1V_1 = M_2V_2$

  M_2 : Resultant molarity

- If a solution having volume V_1 and molarity M_1 is mixed with another solution of same solute having volume V_2 mL & molarity M_2

  then $M_1V_1 + M_2V_2 = M_R (V_1 + V_2)$

  M_R = Resultant molarity = $\dfrac{M_1V_1 + M_2V_2}{V_1 + V_2}$

Solved Examples

Example-23 149 gm of potassium chloride (KCl) is dissolved in 10 Lt of an aqueous solution. Determine the molarity of the solution (K = 39, Cl = 35.5)

Solution Molecular mass of KCl = 39 + 35.5 = 74.5 gm

$\therefore$ Moles of KCl $= \dfrac{149 \text{ gm}}{74.5 \text{ gm}} = 2$

$\therefore$ Molarity of the solution $= \dfrac{2}{10} = 0.2$ M

Molality (m) :

The number of moles of solute dissolved in1000 gm (1 kg) of a solvent is known as the molality of the solution.

i.e., $\text{molality} = \dfrac{\text{number of moles of solute}}{\text{mass of solvent in gram}} \times 1000$

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Let Y gm of a solute is dissolved in X gm of a solvent. The molecular mass of the solute is M_0. Then Y/M_0 mole of the solute are dissolved in X gm of the solvent. Hence

$$\text{Molality} = \frac{Y}{M_0 \times X} \times 1000$$

- ***Molality is independent of temperature changes.***

Solved Examples

Example-24 225 gm of an aqueous solution contains 5 gm of urea. What is the concentration of the solution in terms of molality. (Mol. wt. of urea = 60)

Solution Mass of urea = 5 gm

Molecular mass of urea = 60

Number of moles of urea = $\frac{5}{60}$ = 0.083

Mass of solvent = (255 – 5) = 250 gm

$$\therefore \quad \text{Molality of the solution} = \frac{\text{Number of moles of solute}}{\text{Mass of solvent in gram}} \times 1000 = \frac{0.083}{250} \times 1000 = 0.332.$$

Mole fraction (x) :

The ratio of number of moles of the solute or solvent present in the solution and the total number of moles present in the solution is known as the mole fraction of substances concerned.

Let number of moles of solute in solution = n

Number of moles of solvent in solution = N

$$\therefore \quad \text{Mole fraction of solute } (x_1) = \frac{n}{n+N}$$

$$\therefore \quad \text{Mole fraction of solvent } (x_2) = \frac{N}{n+N}$$

also $x_1 + x_2 = 1$

- Mole fraction is a pure number. It will remain independent of temperature changes.

% calculation :

The concentration of a solution may also expressed in terms of percentage in the following way.

- **% weight by weight (w/w) :** It is given as mass of solute present in per 100 gm of solution.

  i.e. $\quad \% \text{ w/w} = \frac{\text{mass of solute in gm}}{\text{mass of solution in gm}} \times 100$

- **% weight by volume (w/v) : It is given as mass of solute present in per 100 ml of solution.**

  i.e., $\quad \% \text{ w/v} = \frac{\text{mass of solute in gm}}{\text{volume of solution in ml}} \times 100$

- **% volume by volume (v/v) : It is given as volume of solute present in per 100 ml solution.**

  i.e., $\quad \% \text{ v/v} = \frac{\text{volume of solute in ml}}{\text{volume of solution in ml}} \times 100$

Solved Examples

Example-25 0.5 g of a substance is dissolved in 25 g of a solvent. Calculate the percentage amount of the substance in the solution.

Solution Mass of substance = 0.5 g

Mass of solvent = 25 g

$$\therefore \quad \text{percentage of the substance (w/w)} = \frac{0.5}{0.5+25} \times 100 = 1.96$$

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Example-26 20 cm^3 of an alcohol is dissolved in80 cm^3 of water. Calculate the percentage of alcohol in solution.

Solution Volume of alcohol = 20 cm^3

Volume of water = 80 cm^3

$\therefore$ Percentage of alcohol = $\frac{20}{20+80} \times 100$ = 20.

Hydrogen peroxide (H_2O_2)

H_2O_2 can behave both like oxidising and reducing agent in both the mediums (acidic and basic).

$$H_2\overset{-1}{O}_2 \xrightarrow{\text{Reduce}} H_2\overset{-2}{O}$$

$$H_2\overset{-1}{O}_2 \xrightarrow{\text{Oxidise}} \overset{0}{O}_2$$

- **Oxidising agent : (H_2O_2 ì H_2O)**

  (a) Acidic medium : $2e^- + 2H^+ + H_2O_2 \longrightarrow 2H_2O$

  v.f. = 2

  (b) Basic medium : $2e^- + H_2O_2 \longrightarrow 2OH^-$

  v.f = 2

- **Reducing agent : (H_2O_2 ì O_2)**

  (a) Acidic medium : $H_2O_2 \longrightarrow O_2 + 2H^+ + 2e^-$

  v.f = 2

  (b) Basic medium : $2OH^- + H_2O_2 \longrightarrow O_2 + 2H_2O + 2e^-$

  v.f = 2

Volume strength of H_2O_2 : Strength of H_2O_2 is represented as 10V , 20 V , 30 V etc.

20V H_2O_2 means **one litre** of this sample of H_2O_2 on decomposition gives **20L of O_2** gas at **STP.**

Decomposition of H_2O_2 is given as :

$$H_2O_2 \longrightarrow H_2O + \frac{1}{2}O_2$$

1 mole $\qquad \frac{1}{2} \times 22.4$ L O_2 at STP

= 34g $\qquad$ = 11.2 L O_2 at STP

- **Molarity of H_2O_2 (M) =** $\dfrac{\text{Volume strength of } H_2O_2}{11.2}$

  Strength (in g/L) : Denoted by S

  Strength = Molarity × Mol. wt = Molarity × 34

Hardness of water (Hard water does not give lather with soap)

Temporary hardness - due to bicarbonates of Ca & Mg

Permanent hardness - due to chlorides & sulphates of Ca & Mg. There are some method by which we can soften the water sample.

(a) By boiling : $2HCO_3^- \longrightarrow H_2O + CO_2 + CO_3^{2-}$ **or**

By Slaked lime : $Ca(HCO_3)_2 + Ca(OH)_2 \longrightarrow CaCO_3 + 2H_2O$

$Ca^{2+} + CO_3^{2-} \longrightarrow CaCO_3$

(b) By Washing Soda : $CaCl_2 + Na_2CO_3 \longrightarrow CaCO_3 + 2NaCl$

(c) By ion exchange resins : $Na_2R + Ca^{2+} \longrightarrow CaR + 2Na^+$

(d) By adding chelating agents like $(PO_3^-)_3$ etc.

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Parts Per Million (ppm)

When the solute is present in very less amount, then this concentration term is used. It is defined as the number of parts of the solute present in every 1 million parts of the solution. ppm can both be in terms of mass or in terms of moles. If nothing has been specified, we take ppm to be in terms of mass. Hence, a 100 ppm solution means that 100 g of solute is present in every 1000000 g of solution.

$$ppm_A = \frac{\text{mass of A}}{\text{Total mass}} \times 10^6 = \text{mass fraction} \times 10^6$$

Measurement of Hardness :

Hardness is measured in terms of ppm (parts per million) of $CaCO_3$ or equivalent to it.

$$\text{Hardness in ppm} = \frac{\text{mass of } CaCO_3}{\text{Total mass of solution}} \times 10^6$$

Solved Examples

Example-27 0.00012% $MgSO_4$ and 0.000111% $CaCl_2$ is present in water. What is the measured hardness of water and millimoles of washing soda required to purify water 1000 L water ?

Solution. Basis of calculation = 100 g hard water

$$MgSO_4 = 0.00012g = \frac{0.00012}{120} \text{ mole}$$

$$CaCl_2 = 0.000111g = \frac{0.000111}{111} \text{ mole}$$

$\therefore$ equivalent moles of $CaCO_3 = \left(\frac{0.00012}{120} + \frac{0.000111}{111}\right)$ mole

$\therefore$ mass of $CaCO_3 = \left(\frac{0.00012}{120} + \frac{0.000111}{111}\right) \times 100 = 2 \times 10^{-4}$ g

Hardness (in terms of ppm of $CaCO_3$) $= \frac{2 \times 10^{-4}}{100} \times 10^6 = 2$ ppm

$$CaCl_2 + Na_2CO_3 \longrightarrow CaCO_3 + 2NaCl$$

$$NaSO_4 + Na_2CO_3 \longrightarrow MgCO_3 + Na_2SO_4$$

$\therefore$ Required Na_2CO_3 for 100g of water $= \left(\frac{0.00012}{120} + \frac{0.000111}{111}\right)$ mole

$= 2 \times 10^{-6}$ mole

$\therefore$ Required Na_2CO_3 for 1000 litre water $= \frac{2 \times 10^{-6}}{100} \times 10^6 = \frac{2}{100}$ mole ($\because$ d = 1g/mL)

$= \frac{20}{1000}$ mole = 20 m mole

Strength of Oleum :

Oleum is SO_3 dissolved in 100% H_2SO_4. Sometimes, oleum is reported as more than 100% by weight, say y% (where y > 100). This means that (y – 100) grams of water, when added to 100 g of given oleum sample, will combine with all the free SO_3 in the oleum to give 100% sulphuric acid.

Hence, weight % of free SO_3 in oleum = 80(y –100)/18

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Solved Examples

Example-28 What volume of water is required (in mL) to prepare 1 L of 1 M solution of H_2SO_4 (density = 1.5g/mL) by using 109% oleum and water only (Take density of pure water = 1 g/mL).

Solution. 1 mole H_2SO_4 in 1L solution = 98 g H_2SO_4 in 1500 g solution = 98 g H_2SO_4 in 1402 g water.
Also, in 109% oleum, 9 g H_2O is required to form 109 g pure H_2SO_4 & so, to prepare 98 g H_2SO_4, water needed is 9/109 × 98 = 8.09 g.
Total water needed = 1402 + 8.09 = 1410.09 g = **1410.09 mL**

MISCELLANEOUS SOLVED PROBLEMS (MSPS)

1. Find the relative atomic mass, atomic mass of the following elements.
(i) Na (ii) F (iii) H (iv) Ca (v) Ag

Sol. (i) 23, 23 amu (ii) 19, 19 amu (iii) 1, 1.008 amu , (iv) 40, 40 amu, (v) 108, 108 amu.

2. A sample of (C_2H_6) ethane has the same mass as 10^7 molecules of methane. How many C_2H_6 molecules does the sample contain ?

Sol. Moles of $CH_4 = \frac{10^7}{N_A}$

Mass of $CH_4 = \frac{10^7}{N_A} \times 16$ = mass of C_2H_6

So Moles of $C_2H_6 = \frac{10^7 \times 16}{N_A \times 30}$

So No. of molecules of $C_2H_6 = \frac{10^7 \times 16}{N_A \times 30} \times N_A$ = **5.34×10^6.**

3. From 160 g of SO_2 (g) sample, 1.2046×10^{24} molecules of SO_2 are removed then find out the volume of left over SO_2 (g) at STP.

Sol. Given moles = $\frac{160}{64}$ = 2.5.

Removed moles = $\frac{1.2046 \times 10^{24}}{6.023 \times 10^{23}}$ = 2.

so left moles = 0.5.
volume left at STP = 0.5 × 22.4 = 11.2 lit.

4. 14 g of Nitrogen gas and 22 g of CO_2 gas are mixed together. Find the volume of gaseous mixture at STP.

Sol. Moles of $N_2 = \frac{14}{28}$ = 0.5.

moles of $CO_2 = \frac{22}{44}$ = 0.5.

so total moles = 0.5 + 0.5 = 1.
so vol. at STP = 1 × 22.4 = 22.4 lit.

5. Show that in the reaction $N_2 (g) + 3H_2(g) \rightarrow 2NH_3 (g)$, mass is conserved.

Sol.

	N_2 (g) +	$3H_2$(g) $\rightarrow$	$2NH_3$ (g)
moles before reaction	1	3	0
moles after reaction	0	0	2

Mass before reaction = mass of 1 mole N_2(g) + mass of 3 mole H_2(g)
= 14 x 2 + 3 x 2 = 34 g

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mass after reaction = mass of 2 mole NH_3
= 2 x 17 = 34 g.

6. When x gram of a certain metal brunt in 1.5 g oxygen to give 3.0 g of its oxide. 1.20 g of the same metal heated in a steam gave 2.40 g of its oxide. shows the these result illustrate the law of constant or definite proportion

Sol. Wt. of metal = 3.0 – 1.5 = 1.5 g
so wt. of metal : wt of oxygen = 1.5 : 1.5 = 1 : 1
similarly in second case ,
wt. of oxygen = 2.4 – 1.2 = 1.2 g
so wt. of metal : wt of oxygen = 1.2 : 1.2 = 1 : 1
so these results illustrate the law of constant proportion.

7. Find out % of O & H in H_2O compound.

Sol. % of O = $\frac{16}{18} \times 100 = 88.89\%$

% of H = $\frac{2}{18} \times 100 = 11.11\%$

8. Acetylene & butene have empirical formula CH & CH_2 respectively. The molecular mass of acetylene and butene are 26 & 56 respectively deduce their molecular formula.

Ans. C_2H_2 & C_4H_8

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

For Acetylene :

$$n = \frac{26}{13} = 2$$

$\therefore$ Molecular formula = C_2H_2

For Butene :

$$n = \frac{56}{14} = 4$$

$\therefore$ Molecular formula = C_4H_8.

9. An oxide of nitrogen gave the following percentage composition :
N = 25.94
and O = 74.06
Calculate the empirical formula of the compound.

Ans. N_2O_5

Sol.

Element	% / Atomic mass	Simple ratio	Simple intiger ratio
N	$\frac{25.94}{14} = 1.85$	1	2
O	$\frac{74.06}{16} = 4.63$	2.5	5

So empirical formula is N_2O_5.

10. Find the density of CO_2(g) with respect to N_2O(g).

Sol. R.D. = $\frac{\text{M.wt. of } CO_2}{\text{M.wt. of } N_2O} = \frac{44}{44} = 1.$

11. Find the vapour density of N_2O_5

Sol. V.D. = $\frac{\text{Mol. wt. of } N_2O_5}{2} = 54.$

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12. Write a balance chemical equation for following reaction :
When ammonia (NH_3) decompose into nitrogen (N_2) gas & hydrogen (H_2) gas.

Sol. $NH_3 \rightarrow \frac{1}{2}N_2 + \frac{3}{2}H_2$ or $2NH_3 \rightarrow N_2 + 3H_2$.

13. When 170 g NH_3 (M =17) decomposes how many grams of N_2 & H_2 is produced.

Sol. $NH_3 \rightarrow \frac{1}{2}N_2 + \frac{3}{2}H_2$

$$\frac{\text{moles of } NH_3}{1} = \frac{\text{moles of } N_2}{1/2} = \frac{\text{moles of } H_2}{3/2}.$$

So moles of $N_2 = \frac{1}{2} \times \frac{170}{17} = 5.$ So wt. of $N_2 = 5 \times 28 = 140$ g.

Similarly moles of $H_2 = \frac{3}{2} \times \frac{170}{17} = 15.$

So wt. of $H_2 = 15 \times 2 = 30$ g.

14. 340 g NH_3 (M = 17) when decompose how many litres of nitrogen gas is produced at STP.

Sol. $NH_3 \rightarrow \frac{1}{2}N_2 + \frac{3}{2}H_2$

moles of $NH_3 = \frac{340}{17} = 20.$

So moles of $N_2 = \frac{1}{2} \times 20 = 10.$

$\therefore$ vol. of N_2 at STP = $10 \times 22.4 = 224$ lit.

15. 4 mole of $MgCO_3$ is reacted with 6 moles of HCl solution. Find the volume of CO_2 gas produced at STP, the reaction is

$$MgCO_3 + 2HCl \rightarrow MgCl_2 + CO_2 + H_2O.$$

Sol. Here HCl is limiting reagent. So moles of CO_2 formed = 3.
So vol. at STP = $3 \times 22.4 = 67.2$ lit.

16. 117 gm NaCl is dissolved in 500 ml aqueous solution. Find the molarity of the solution.

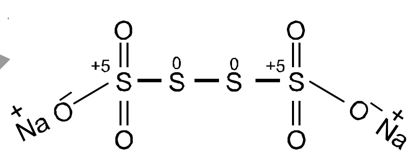
Sol. Molarity = $\frac{117/58.5}{500/1000}$ = **4M.**

17. 0.32 mole of $LiAlH_4$ in ether solution was placed in a flask and 74 g (1 moles) of t-butyl alcohol was added. The product is $LiAlHC_{12}H_{27}O_3$. Find the weight of the product if lithium atoms are conserved.
[Li = 7, Al = 27, H = 1, C = 12, O = 16]

Sol. Applying POAC on Li
$1 \times$ moles of $LiAlH_4 = 1\times$ moles of LiAlH $C_{12}H_{27}O_3$
$254 \times 0.32 = 1 \times$ wt. of LiAlH $C_{12}H_{27}O_3$.
wt. of LiAlH $C_{12}H_{27}O_3$ = 81.28 gm.

18. Calculate individual oxidation number of each S-atom in $Na_2S_4O_6$ (sodium tetrathionate) with the help of its structure .

Sol.



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19. Find the average and individual oxidation number of Fe & Pb in Fe_3O_4 & Pb_3O_4, which are mixed oxides.

Sol. (i) Fe_3O_4 is mixture of FeO & Fe_2O_3 in 1 : 1 ratio
so, individual oxidation number of Fe = +2 & +3

& average oxidation number = $\frac{1(+2)+2(+3)}{3}$ = 8/3

(ii) Pb_3O_4 is a mixture of PbO & PbO_2 in 2 : 1 molar ratio
so, individual oxidation number of Pb are +2 & +4

& average oxidation number of Pb = $\frac{2(+2)+1(+4)}{3} = 8/3$

20. Balance the following equations :

(a) $H_2O_2 + MnO_4^- \longrightarrow Mn^{+2} + O_2$ (acidic medium)

(b) $Zn + HNO_3(dil) \longrightarrow Zn(NO_3)_2 + H_2O + NH_4NO_3$

(c) $CrI_3 + KOH + Cl_2 \longrightarrow K_2CrO_4 + KIO_4 + KCl + H_2O.$

(d) $P_2H_4 \longrightarrow PH_3 + P_4$

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

Ans. (a) $6H^+ + 5H_2O_2 + 2MnO_4^- \longrightarrow 2Mn^{+2} + 5O_2 + 8H_2O$

(b) $4Zn + 10HNO_3(dil) \longrightarrow 4Zn(NO_3)_2 + 3H_2O + NH_4NO_3$

(c) $2CrI_3 + 64KOH + 27Cl_2 \longrightarrow 2K_2CrO_4 + 6KIO_4 + 54KCl + 32H_2O.$

(d) $6P_2H_4 \longrightarrow 8PH_3 + P_4$

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

21. Calculate the resultant molarity of following :

(a) 200 ml 1M HCl + 300 ml water
(b) 1500 ml 1M HCl + 18.25 g HCl
(c) 200 ml 1M HCl + 100 ml 0.5 M H_2SO_4
(d) 200 ml 1M HCl + 100 ml 0.5 M HCl

Ans. (a) 0.4 M (b) 1.33 M (c) 1 M (d) 0.83 M.

Sol. (a) Final molarity = $\frac{200\times1+0}{200+300}$ = 0.4 M.

(b) Final molarity = $\frac{1500\times1+\frac{18.25\times1000}{36.5}}{1500} = 1.33$ M

(c) Final molarity of H^+ = $\frac{200\times1+100\times0.5\times2}{200+100}$ = 1 M.

(d) Final molarity = $\frac{200\times1+100\times0.5}{200+100}$ = 0.83 M.

22. 518 gm of an aqueous solution contains 18 gm of glucose (mol.wt. = 180). What is the molality of the solution.

Sol. wt. of solvent = 518 – 18 = 500 gm. $\Rightarrow$ so molarity = $\frac{18/180}{500/1000}$ = 0.2.

23. 0.25 of a substance is dissolved in 6.25 g of a solvent. Calculate the percentage amount of the substance in the solution.

Sol. wt. of solution = 0.25 + 6.25 = 6.50.

so % (w/w) = $\frac{0.25}{6.50}$ × 100 = 3.8%.

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24. A fresh H_2O_2 solution is labelled 11.2 V. This solution has the same concentration as a solution which is :
(A) 3.4% (w / w) (B) 3.4% (v / v) (C) 3.4% (w / v) (D) None of these

Sol. Molarity of $H_2O_2 = \frac{\text{vol.strength}}{11.2} = \frac{11.2}{11.2} = 1$

Now, $\%(w/v) = \frac{\text{wt. of solute in g}}{\text{wt. of solution in mL}} \times 100$

$= \text{Molarity} \times \text{Mol. wt. of solute} \times \frac{1}{10}$

$= 1 \times 34 \times \frac{1}{10} = 3.4\%$ **Ans. (C)**

25. Calculate the resultant molarity of following :
(a) 200 ml 1M HCl + 300 ml water
(b) 1500 ml 1M HCl + 18.25 g HCl
(c) 200 ml 1M HCl + 100 ml 0.5 M H_2SO_4
(d) 200 ml 1M HCl + 100 ml 0.5 M HCl

Ans. (a) 0.4 M (b) 1.33 M (c) 1 M (d) 0.83 M.

Sol. (a) Final molarity = $\frac{200 \times 1 + 0}{200 + 300}$ = 0.4 M.

(b) Final molarity = $\frac{1500 \times 1 + \frac{18.25 \times 1000}{36.5}}{1500} = 1.33$ M

(c) Final molarity of H^+ = $\frac{200 \times 1 + 100 \times 0.5 \times 2}{200 + 100}$ = 1 M.

(d) Final molarity = $\frac{200 \times 1 + 100 \times 0.5}{200 + 100}$ = 0.83 M.

26. 518 gm of an aqueous solution contains 18 gm of glucose (mol.wt. = 180). What is the molality of the solution.

Sol. wt. of solvent = 518 – 18 = 500 gm. $\Rightarrow$ so molarity = $\frac{18/180}{500/1000}$ = 0.2.

27. 0.25 of a substance is dissolved in 6.25 g of a solvent. Calculate the percentage amount of the substance in the solution.

Sol. wt. of solution = 0.25 + 6.25 = 6.50.

so % (w/w) = $\frac{0.25}{6.50} \times 100 = 3.8\%$.

28. A fresh H_2O_2 solution is labelled 11.2 V. This solution has the same concentration as a solution which is :
(A) 3.4% (w / w) (B) 3.4% (v / v) (C) 3.4% (w / v) (D) None of these

Sol. Molarity of $H_2O_2 = \frac{\text{vol.strength}}{11.2} = \frac{11.2}{11.2} = 1$

Now, $\%(w/v) = \frac{\text{wt. of solute in g}}{\text{wt. of solution in mL}} \times 100$

$= \text{Molarity} \times \text{Mol. wt. of solute} \times \frac{1}{10}$

$= 1 \times 34 \times \frac{1}{10} = 3.4\%$ **Ans. (C)**

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