I PUC
SOME BASIC CONCEPTS OF CHEMISTRY

LAWS OF CHEMICAL COMBINATION

(i) **Law of conservation of mass:** In all physical and chemical changes, the total mass of the reactants is equal to that of the products. In other words, matter can neither be created nor be destroyed.

$$\text{e.g. } HgO(s) \xrightarrow[\text{close tube}]{\Delta} Hg(\ell) + \frac{1}{2}O_2(g)$$
$$100\,g \qquad\qquad 92.6\,g \quad 7.4\,g$$

(ii) **Law of constant composition or definite proportion:** A chemical compound is always found to be made up of the same elements combined together in the same fixed proportion by weight.

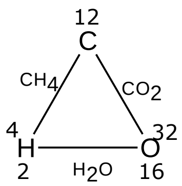
For example, pure water obtained from whatever source or any country will always be made up of only hydrogen and oxygen elements combined together in the same fixed ratio of 1: 8 by weight.

(iii) **Law of multiple proportions:** When two elements combine to form two or more chemical compounds then the weights of one of elements which combine with a fixed weight of the other bear a simple ratio to one another.

e.g. the weights of nitrogen and oxygen forming compound N_2O, NO, N_2O_3, N_2O_4, N_2O_5, if we fix the weight of nitrogen as 14 then weight of oxygen bears a simple ratio 1 : 2 : 3 : 4 : 5 to one another.

(iv) **Law of reciprocal proportions:** The ratio of the weights of two elements A and B which combine separately with a fixed weight of the third element C is either the same or some simple multiple of the ratio of the weights in which A and B combine directly with each other. For example : CO_2, SO_2 and CS_2, H_2S, H_2O and SO_2.

Eg:



(v) **Gay Lussac's law of gaseous volumes:** When gases react together they always do so in volumes which bears a simple ratio to one another and to the volumes of the products, if these are also gases, provided all measurements of volumes are done under similar conditions of temperature and pressure. For example

$\underset{1\text{ vol}}{H_2} + \underset{1\text{ vol}}{Cl_2} = \underset{2\text{ vol}}{2HCl}$	ratio	1: 1: 2
$\underset{2\text{ vol}}{2H_2} + \underset{1\text{ vol}}{O_2} = \underset{2\text{ vol}}{2H_2O}$	ratio	2: 1: 2
$\underset{2\text{ vol}}{2CO} + \underset{1\text{ vol}}{O_2} = \underset{2\text{ vol}}{2CO_2}$	ratio	2: 1: 2
$\underset{1\text{ vol}}{N_2} + \underset{3\text{ vol}}{3H_2} = \underset{2\text{ vol}}{2NH_3}$	ratio	1: 3: 2

Atomic Mass

$$\text{Atomic mass of an element} = \frac{\text{Average mass of an atom of the element}}{\frac{1}{12}^{th}\text{ the mass of an atom of C} - 12}$$

One gram of atomic mass of any substance (Eg: For water, 18g) always contains 6.022×10^{23} atoms.

$1/12^{th}$ mass of an atom of carbon – 12 is called one a. m. u (atomic mass unit).

1 amu = 1.66×10^{-27} Kg

Molecular mass

$$\text{Molecular mass} = \frac{\text{Mass of one molecule of the substance}}{\frac{1}{12}^{th} \text{ the mass of an atom of } C-12}$$

One gram molecular weight of any substance (Eg: For water, 18g) always contains 6.022×10^{23} Molecules.

Molecular mass of an element (like O_2, H_2, O_3 molecule etc)
Mol. Mass = Atomic mass x Atomicity

Atomicity:

is the number of atoms present in a molecule of an elementary gas.

a) Monoatomic molecules: Ar, He, Ne etc — Atomicity is 1
b) Diatomic molecules: O_2, H_2, N_2 etc — Atomicity is 2
c) Triatomic molecules: O_3 — Atomicity is 3
d) Tetra atomic molecules: P_4 — Atomicity is 4

METHODS FOR THE DETERMINATION OF ATOMIC MASS

(i) Dulong and Petit's Law: According to this law, the product of atomic mass and specific heat of a solid element is approximately equal to 6.4. The product of atomic mass and specific heat is called atomic heat. Thus,

Atomic mass x Specific heat = 6.4

Or Atomic mass (approximate) = $\frac{6.4}{\text{Specific heat}}$

The law is applicable only to solid elements except Be ,B, C and Si.

Methods of determining Molecular Masses

1. Vapour density method:

$$\text{V.D.} = \frac{1}{2}\ \frac{\text{Mass of volatile substance}}{\text{Volume of vapour of its chloride at STP}} \times 22400$$

Molecular mass = 2 x V.D.

For volatile chloride, Valency $= \dfrac{2xV.D}{Eq\ mass\ of\ element + 35.5}$

2. Diffusion method:

According to Graham's law of diffusion, rate of diffusion of a gas is inversely proportional to the square root of its molecular mass.

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

3. Victor Meyer method:

This method applies to volatile organic liquids. Suppose vapour of an organic liquid having mass 'W' g occupies a volume of 'V' mL at STP. Then its molecular mass is,

Mol. mass = $\frac{W}{V}$ x 22, 400.

Empirical and molecular formulae of an organic compound

Empirical formula: Empirical formula of a compound is the simplest formula deduced from its percentage composition. It indicate the relative number of atoms of different elements present in a molecule of the compound

Molecular formula: Molecular formula of a compound is the formula which gives exact number of atoms of different elements present in the molecule

Steps involved in calculating the empirical formula and molecular formula

Step 1: The percentage composition by mass of each element present in the compound is determined

Step 2: The percentage composition of each element is divided by the atomic mass of individual elements

Step 3: The values obtained in step 2 are divided by the least value

The values obtained in step 3 are rounded off the nearest integer. This gives the relative number of atoms of different elements present in the molecule. It is the empirical formula.

Step 4: Molecular formula=n x Empirical formula

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

AVOGADRO'S HYPOTHESIS

Equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules.

Applications of Avogadro's law

(i) In the calculation of atomicity of elementary gases.

(ii) To find the relationship between molecular weight and vapour density of a gas.

$$\text{Molecular weight} = 2 \times \text{Vapour density}$$

(iii) To find the relationship between weight and volume of a gas.

Molecular weight = Weight of 22.4 L of the gas at STP

Vapour density is the ratio of mass of certain volume of a gas to the mass of equal volume of hydrogen both measured under similar condition of temperature and pressure

EQUIVALENT MASS:

Equivalent mass of an element is the mass of it which combines or displaces 8 parts by mass of oxygen or 35.45 parts by mass of chlorine or 1.008 parts by mass of hydrogen.

Always one equivalent mass of one element combines with one equivalent of any other element. That is, one equivalent of one substance always reacts with one equivalent of any other substance.

Atomic mass = Equivalent mass × Valency

Note: Elements with variable valencies have variable equivalent masses

Determination of equivalent mass:

OXIDE METHOD:

In the oxide method of determination of equivalent mass of the element a known mass of the element is converted to its oxide and mass of the oxide formed is determined. From this mass of the element which combines with 8 g of oxygen is calculated as

$$\frac{\textbf{Mass of element}}{\textbf{Mass of oxygen}} \times \mathbf{8}$$

This is equivalent mass of the element.

CHLORIDE METHOD:

In the chloride method a known mass of the element is converted to its chloride. The mass of chloride formed is determined. From this the mass of the element which combines with 35.45g of chlorine is calculated.

Equivalent mass = $\dfrac{\text{mass of the element}}{\text{mass of chlorine}} \times 35.45$

HYDROGEN DISPLACEMENT METHOD:

In hydrogen displacement method a certain mass of the metal is made to react with excess of dilute acid. The volume of hydrogen liberated is measured. The volume of hydrogen liberated is converted to STP using the formula

$$\frac{P_0V_0}{T_0} = \frac{(P_1 - f) \times V_1}{T_1}$$

From this the mass of the metal which displaces 11,200 cm^3 of hydrogen at STP is calculated using the expression

$$\frac{\textbf{mass of metal}}{\textbf{volume of hydrogen displaced at STP}} \times \mathbf{11{,}200}$$

This is the equivalent mass of the metal.

Since hydrogen is collected over water the pressure is taken as (P_1- f) where f is the aqueous tension.

INTERCONVERSION METHOD:

Inter conversion method a known mass of the compound AX is converted to another compound AY. Then the equivalent mass of A is calculated as

$$\frac{\textbf{mass of AX}}{\textbf{mass of AY}} = \frac{\textbf{Equivalent mass of AX}}{\textbf{Equivalent mass of AY}} = \frac{\textbf{Equivalent mass of A} + \textbf{Eq.mass of X}}{\textbf{Equivalent mass of A} + \textbf{Eq.mass of Y}}$$

Knowing the equivalent masses of X and Y the equivalent mass of A can be calculated.

ELECTROLYTIC METHOD:

When 1 Faraday = 96500 coulomb of electricity is passed through the solution containing a metallic salt, one gram equivalent mass of the metal gets deposited.

When 96500 coulombs of electricity is passed through acidified water solution, the amount of hydrogen liberated is 1.008 g or 11200cm^3 at STP.

Coulombs of electricity= current flows in amps$\times$time in secs.

<u>Faraday's 1st law</u>: Quantity of substance discharged $\propto$ the quantity of electricity passed

Mass of Na discharged when 10 ampere current flows for 10 hours is calculated as follows

Number of coulombs flowing = It = 10 $\times$ 10 $\times$ 60 $\times$ 60 = 3,60,000 coulomb

96500 coulombs deposit 23 g of Na

$\therefore$ 3, 60, 000 coulombs deposit 85.8 g of Na

Faraday's second law

When same amount of electricity is passed through different solution, the amount of elements discharged will be proportional to their equivalent masses.

E α m

Note: **Equivalent mass of a radical** = $\dfrac{\textbf{Formula mass}}{\textbf{Valency}}$

Equivalent mass of SO_4^{2-} = $\dfrac{96}{2}$ = 48

Equivalent mass of the substance can be obtained by adding the equivalent masses of constituent parts.

EQUIVALENT MASS OF COMPOUNDS

EQUIVALENT MASS OF ACID:

Basicity of the acid is the number of replaceable hydrogen atoms present in a molecule of the acid.

HCl has one displaceable hydrogen atom. So its basicity is 1

Acid	Basicity (maximum)
HCl, HNO_3, H_3BO_3	1
H_2SO_4, H_3PO_3, $H_2C_2O_4.2H_2O$	2
H_3PO_4	3
$H_4P_2O_7$	4

Equivalent mass of the acid is the number of parts by mass of the acid which contains 1.008 parts by mass of replaceable hydrogen

It is related to molecular mass as

Equivalent mass = $\frac{\text{Molecular mass}}{\text{Basicity}}$.

Ex: For H_2SO_4, Equivalent mass = $\frac{98}{2} = 49$

To find the equivalent mass of H_3PO_4 in the reaction

$\mathbf{Ca(OH)_2 + H_3PO_4 \rightarrow CaHPO_4 + 2H_2O}$

From the equation it is clear that only two hydrogens are replaced.

Hence equivalent mass= $\frac{\text{Molecular mass}}{2} = \frac{98}{2} = 49$

EQUIVALENT MASS OF BASE

Acidity of a base is the number of replaceable OH groups present in a molecule of the base.

NaOH has one – OH group. Hence its acidity is one

Base	Acidity (maximum)
NaOH, KOH	1
$Ca(OH)_2$, $Ba(OH)_2$	2
$Al(OH)_3$	3

Equivalent mass of a base is the number of parts by mass of the base which combines with one equivalent of the acid.

It is related to the molecular mass as

Equivalent mass = $\frac{\text{Molecular mass}}{\text{Acidity}}$

Ex: $\mathbf{Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O}$

$\therefore$ Equivalent mass of $Ca(OH)_2$ = $\frac{74}{2} = 37$

EQUIVALENT MASS OF SALT:

Equivalent Mass of a salt = $\frac{\text{Molecular mass}}{\text{Total no. of positive charges}}$

Ex: Equivalent mass of $Na_2SO_4 = \frac{\text{Mol. mass}}{2} = \frac{142}{2} = 71$

Eq. mass of sodium carbonate ($Na_2CO_3 \rightarrow 2Na^+ + CO_3^{-2}$)

Mol. Mass of Na_2CO_3 is 106 , total number of '+ 've charges is 2

Equivalent mass of $Na_2CO_3 = \frac{\text{Mol. mass}}{\text{Total positive valencies}} = 106/2 = 53$

EQUIVALENT MASS OF OXIDISING AGENT:

During oxidation, oxidation number of one of the element of the oxidizing agent decreases. The equivalent mass is related to molecular mass as

$$\text{Equivalent mass} = \frac{\text{Molecular mass}}{\text{Decrease in oxidation number}}$$

Equivalent Mass of oxidizing agent is a variable

Example:

$KMnO_4$ acts as an oxidizing agent in different medium.

Acid medium: $KMnO_4$ changes to $MnSO_4$

Change in the oxidation state of the Mn = 7 - 2 = 5

$\therefore$ Equivalent mass $= \frac{\text{Molecular mass}}{5} = \frac{158}{5} = 31.6$

Basic medium: $KMnO_4$ changes to K_2MnO_4

Change in oxidation state of Mn = 7 – 6 = 1

$\therefore$ Equivalent mass $= \frac{\text{Molecular mass}}{1} = 158$

Neutral medium: $KMnO_4$ changes to MnO_2

$\therefore$ Change in the oxidation state of Mn = 7 – 4 = 3

$\therefore$ Equivalent mass $= \frac{\text{Molecular mass}}{3} = \frac{158}{3} = 52.67$

EQUIVALENT MASS OF REDUCING AGENT:

Equivalent mass of a reducing agent is the number of parts by mass of reducing agent which combines with 8 parts by mass of oxygen or which gives 1.008 parts by mass of hydrogen for reduction

Equivalent masses of reducing agent can also be calculated as

$$\text{Equivalent mass} = \frac{\text{Molecular mass}}{\text{Change in oxidation state}}$$

Equivalent mass of Mohr's salt:

Change in oxidation number = Fe^{+2} to Fe^{+3} = 1

$\therefore$ Equivalent mass of Mohr's salt $= \frac{\text{Mol. mass}}{1} = 392$

STP is standard temperature and pressure. The standard temperature is 273 K and standard pressure is 101.3 kPa (or 1 atm or 760 mm)

Different ways of calculating Equivalent Weight and n-factor

- Equivalent Weight of Element (atomic state) $= \frac{\text{Atomic weight}}{n}$

  where n = Valency of element

  n = Change in oxidation number per atom

  = Number of e^- lost or gained per atom.

- Equivalent Weight of Compound/Molecule $= \frac{\text{Molecular weight}}{n}$

  where n= Acidity of the base

n = Basicity of the acid
= Number of electrons lost or gained per molecule.
= Change in oxidation number per molecule of the substance
= Total charge on cation or anion.

Mole concept

One mole of the any substance contains Avogadro number of particles.
The number of atoms present in 12g of C-12 isotope is called Avogadro number.
Its value is 6.022×10^{23}.
One mole of any gas occupies $22.4 dm^3$ at STP.
A mole of any substance is related to:
mass of a substance
number of particles
volume of the gaseous substance

Mole-particle relationship

A mole is a collection of 6.023×10^{23} particles, ions, atoms etc.
i.e. i) 6.023×10^{23} atoms of Na constitute one mole atom of Na.
ii) 6.023×10^{23} molecules of oxygen constitutes 1 mole of oxygen molecules.
iii) 6.023×10^{23} electrons constitute one mole of electrons.

Mole-weight relationship

One mole of every substance weighs equal to the gram atomic weight of the substance or to the gram molecular weight of the substance.
e.g. i) 1 mole of sodium weighs 23 g of Na.
ii) 1 mole of $CaCO_3$ weighs 100 g.

Mole-Volume relationship

One mole of every gas occupies 22.4 lit. of volume at STP.
i.e. 1 mole of O_2 occupies 22400 ml of volume at STP.
1 mole of He occupies 22400 ml of volume at STP.

Note: Hence 1 gram molecular mass (1 mole) = 6.023×10^{23} particles = $22.4 dm^3$ at STP.

$$\text{No. of moles} = \frac{\text{Mass}}{\text{Molecular mass}}$$

$$\text{No. of moles} = \frac{\text{Mass}}{\text{Atomic mass}}$$

Mass of an atom = $\frac{\text{Gram at. wt}}{6.022 \times 10^{23}}$

Mass of a molecule = $\frac{\text{Gram mol. wt}}{6.022 \times 10^{23}}$

$$\text{No. of moles} = \frac{\text{No. of atoms}}{6.022 \times 10^{23}}$$

$$\text{No. of moles} = \frac{\text{No. of molecules}}{6.022 \times 10^{23}}$$

$$\text{No. of moles of a gas} = \frac{\text{Volume of the gas at S.T.P in cm}^3}{22400\ \text{cm}^3}$$

Mass of 1 dm^3 of a gas = $\frac{\text{Gram. mol. wt}}{22.4}$

Relations based on chemical equations

Consider an example:

$2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$

2 × 1 atomic mass of Aluminium (= 54 g) reacts with 6 molecular mass of HCl (36.5 × 6 g) to form 2 molecular mass of $AlCl_3$ liberating 3 × 1 molecular mass of hydrogen.

Limiting reagent:

It is defined as the reactant which is consumed completely during the reaction. In stoichiometric calculations limiting reagent determines the amount of product formed.

Ex: 2 moles each of hydrogen and oxygen is ignited to form water.

Balanced equation of formation water

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

$$1\text{ mol} \quad \frac{1}{2}\text{mol} \quad 1\text{ mol}$$

From the equation it is clear that for 1 mole of hydrogen half mole of oxygen is required. Hence 2 mole of hydrogen requires one mole of oxygen. Thus for this process, oxygen is in excess and hydrogen is limited so hydrogen is the limiting reagent.

Concentration of a solution can be expressed as

1) ppm 2) normality 3) molarity 4) mole fraction

i) **ppm** is the parts by mass of a substance present in 10^6 parts by mass of the bulk.

Hence 1g of a substance present in 10^6g of solution gives the concentration 1 ppm.

ii) **Normality** is the number of gram equivalent mass of the solute present in 1000 cm^3 of the solution.

Mass can be converted to gram equivalents using the relation:

No. of gram equivalent = $\dfrac{\text{Mass}}{\text{Equivalent mass}}$

Hence normality of a solution = $\dfrac{\text{Mass of solute in 1 dm}^3}{\text{Equivalent mass}}$

Or W = N × E

$$N = \frac{Mass\ per\ dm^3}{Equivalent\ mass}$$

$$N = \frac{W\ x 1000}{E\ x V}$$

$$N = \frac{Wt.\%\ x density x 10}{Equivalent\ mass}$$

$$N = \frac{Strentgh\ of\ the\ solution\ in\ g\ \overline{L}^{-1}}{Gram\ Equivalent\ weight\ of\ solute}$$

$$N = \frac{\%\ of\ solute\ x 10}{gram\ eq.wt\ of\ solute}$$

iii) **Molarity** is the number of gram molecular mass of the solute present in 1 dm^3 of the solution.

$$\text{Molarity} = \frac{\text{Mass in 1 dm}^3}{\text{Molecular mass}}$$

$$M = \frac{Number\ of\ moles\ of\ solute(n)}{Volume\ of\ solution\ in\ litres}$$

$$M = \frac{Weight\ of\ solute\ per\ litre\ \textbf{of solution}}{Molecular\ weight\ of\ solute}$$

$$M = \frac{W\ x 1000}{Mol.mass\ x V}$$

For HCl; 1N = 1M $\because$ Equivalent mass = Mol. Mass

For H_2SO_4; 2N = 1M $\because$ Eq.mass = $\frac{\text{Mol. mass}}{2}$

For H_3PO_4; 3N = 1M $\because$ Eq.mass= $\frac{\text{Mol. mass}}{3}$

Representation	Molarity	Common name
1 M	1	Molar solution
0.5 M or M/2	1/2	Semimolar solution
0.1 M or M/10	1/10	Decimolar solution
0.01 M or M/100	1/100	Centimolar solution
0.010 M or M/1000	1/1000	Millimolar solution

Relationship between normality and molarity :

We know, Molarity x Molecular mass = Strength of the solution (g/L)
Similarly, Normality x Equivalent mass = strength of the solution (g/L)
Hence, Molarity x Molecular mass = Normality x Equivalent mass

$$\frac{\text{Normality}}{\text{Molarity}} = \frac{\text{Molecular mass}}{\text{Equivalent mass}} = n$$

Normality = n x Molarity

$$NORMALITY = Molarity\ x \frac{Molecular\ mass}{Equivalent\ mass}$$

For acids

n = Basicity of acid, i.e., number or replaceable H^+ ions.
Thus, Normality of an acid = Molarity x Basicity of acid.

For bases:

n = Acidity of base, i.e. number of replaceable OH^- ions.
Thus, normality of a base = Molarity x Acidity of base

For salts:

n = Total positive or negative valency of salt.
Thus, normality of salt = Molarity x total positive or negative valency

For oxidising and reducing agents :

n = Change in oxidation number of oxidising and reducing agent.

Thus, Normality of an oxidising or reducing agent = Molarity x change in oxidation number.

iv) **Molality** (m)is the number of gram molecular mass of a solute present in 1000g of solvent

$$m = \frac{Number\ of\ moles\ of\ solute}{Weight\ of\ solvent\ in\ kg}$$

$$m = \frac{Weight\ of\ solute\ x\ 1000}{Molecular\ weight\ of\ solute\ x\ Weight\ of\ solvent\ in\ gram}$$

v) **Mole fraction** is the ratio of number of moles of a substance to the total number of moles present in the solution.

In a solution containing a liquid, the mole fraction of the components can be calculated by calculating the number of moles of solute and number of moles of solvent.

If n_A is the number of solute and n_B is the number of moles of solvent, then

Mole fraction of solute $X_A = \frac{n_A}{n_A + n_B}$

Mole fraction of solvent $X_B = \frac{n_B}{n_A + n_B}$

Then, $X_A + X_B = \frac{n_A}{n_A + n_B} + \frac{n_B}{n_A + n_B} = 1$

Strength of H_2O_2 solution

Strength of H_2O_2 solution is expressed in terms of volumes of O_2 that can be produced by 1 vol of H_2O_2 solution e.g. 20 vol H_2O_2 solution means 1 vol of H_2O_2 will give 20 vol of O_2 at NTP.

Vol. strength of H_2O_2 = 5.6 x Normality

Vol. strength of H_2O_2 = 11.2 x molarity

% strength = 0.303 x volume strength

(vi) Formality: Formality is similar to Molarity. It is the number of formula weight in gm dissolved per litre of solution It is especially used for ionic crystals like Na^+Cl^-.

$$\text{Formality} = \frac{\text{moles of crystal}}{\text{volume in L of solutions}}$$

$$= \frac{\text{Weight of ionic solute}}{\text{Formula weight of solute}} \times \frac{1}{\text{V (in L) of solution}}$$

(vii)

a)Percentage by mass % w/w: The amount of the solute present in 100 g of solution is called percentage by weight.

$$\%\text{ by weight} = \frac{\text{weight of solute}}{\text{weight of solution}} \times 100$$

b)Percentage weight by volume (% w/v): The amount of solute present in 100 ml of solution is called percentage weight volume.

$$\%\text{ by volume} = \frac{\text{weight of solute}}{\text{volume of solution}} \times 100$$

c)Percentage volume by volume (% v/v): The amount of solute present in 100 ml of solution is called percentage volume by volume.

% by strength $= \frac{\text{volume of solute}}{\text{volume of solution}} \times 100$

(viii) Specific gravity $= \frac{\text{weight of solution}}{\text{volume of solution}}$

Titration:

The experimental procedure involved in the determination of volume of one solution required to completely react with known volume of another solution is called titration.

In volumetric analysis are based on the equation $V_1N_1 = V_2N_2$

During dilution of a solution since the no. of gram equivalents present in both solutions are same, the normality of dilute solution can be calculated using the formula

$$\underset{\text{Before dilution}}{V_1N_1} = \underset{\text{After dilution}}{V_2N_2}$$

When two solutions of different normalities are mixed the normality of resulting solution can be calculated using the formula

$V_1N_1+V_2N_2 = V_3N_3$

- Let V_1 cm^3 of an acid of normality N_1 be mixed with V_2 cm^3 of base of normality N_2
- If $V_1N_1 > V_2N_2$ then the normality of the left over acidic solution is $\frac{V_1N_1 - V_2N_2}{V_1 + V_2}$
- If base is strong then the normality of the left over basic solution is $\frac{V_2N_2 - V_1N_1}{V_1 + V_2}$

Let V_1 cm^3 of an acid of Molarity M_1 be mixed with V_2 cm^3 of base of Molarity M_2

$V_1M_1=V_2M_2$

When two solutions of different Molarities are mixed the molarity of resulting solution can be calculated using the formula

$V_1M_1+V_2M_2 = V_3M_3$

If $V_1M_1 > V_2M_2$ then the Molarity of the left over acidic solution is $\frac{V_1M_1 - V_2M_2}{V_1 + V_2}$

Calculation of Percentage Composition form Formula:

% of an element or constituent in a compound = $\frac{\text{No. of parts by mass of that element or constituent}}{\text{Molecular mass of the compound}}$ x 100