Basic Concept in Chemistry

Prepared by

Shreemoyee Phukan

Silapathar College

1.1 The International system of Units(SI Units)

A Unit is defined as standard reference chosen to measure any physical quantity. Different types of units to measure the same physical quantity are used in different parts of world. In order to maintain the uniformity in the measurements, an internationally accepted system of units called as International system of units or SI is established. The SI system has two types of units. **Fundamental or basic, units and derived units.**

SI Base units are as follows:

Physical Quantity	Unit	Unit Symbol
Length	Meter	m
Mass	Kilogram	kg
Time	second	s
Temperature	Kelvin	K
Amount of substance	mole	mol
Electric current	ampere	A
Luminous intensity	candela	Cd
Plane angle*	radian	rad
Solid angle*	steradian	sr

*These are two dimensionless supplementary units.

The units; which are derived from fundamental SI units are called derived units.
Derived units are as follows:

Physical Quantity	Unit	Symbol
Area	Square meter	m^2
Volume	Cubic meter	m^3
Density	Kilogram per cubic meter	kgm^{-3}
Molar mass	Kilogram per Mole	$kgmol^{-1}$
Molar concentration	Mole per cubic meter	$mol\ m^{-3}$

SI Prefixes are given below:

Multiplex	Prefix	Symbol	Submultiplex	Prefix	Symbol
10^{15}	Peta	P	10^{-1}	deci	d
10^{12}	tera	T	10^{-2}	centi	c
10^{9}	giga	G	10^{-3}	milli	m
10^{6}	mega	M	10^{-6}	micro	µ
10^{3}	kilo	k	10^{-9}	nano	n
10^{2}	hector	h	10^{-12}	pico	p
10^{1}	deca	da	10^{-15}	femto	f

1.2 Precision and Accuracy

While measuring physical quantities the data must be precise and accurate. A measurement is said to be precise when the values of different measurements are close to each other and hence closer to their average value. It means precision refers to closeness of set of values obtained from identical measurements of quantity. At the same time; a measurement is accurate when the average values of different measurements are close to the correct value. It means accuracy refers to the closeness of a single measurement to its true value.

1.3 Significant Figures (S.F)

The total number of digits in the number with last one, having uncertain value, is called the significant figures.

Rules for counting significant figures.

1. All digits are significant except zero at the beginning of the number.
2. The zero to the right of the decimal point are significant.
3. In case of multiplication and/or division, the result, may carry no more S.F than the least precisely known quantity in the calculation e.g.

   (a) 14.79 x 12.11 x 5.05

   (4 S.F) (4 S.F) (3 S.F)

   = 904.48985 = 9.04 x 10^2 (3 S.F) (least S.F)

   (b) $\frac{5.28 \text{ x } 0.156 \text{ x } 3.00}{0.0428}$

   = 57.73458 = 57.7 (3 S.F)
4. In case of addition/subtraction, the result must be expressed with the same number of decimal places as the term carrying the smallest number of decimal places.
   - 22.2 + 2.22 + 0.222 = 24.642 = 24.6

     One(least) Two Three Three One

5. Exact numbers can be considered to have an unlimited number of significant figures. Thus, for counting significant figures for exact number, first rounding off the exact numbers. Various rules are used for rounding off a number.
    (i) If the first digit removed is less than 5, round down by dropping it and all following digits. Thus, 5.663507 becomes 5.66 when rounded off the three S.F because first of the dropped digits (3) is less than 5.
    (ii) If the first digit removed is 6 or greater than 6 round off by adding 1 to the digit on the left. Thus, 5.663507 becomes 5.7 when rounded off to two S.F.
    (iii) If the first digit removed is 5 and there are more non-zero digits following round up.Thus , 5.663507 becomes 5.664 when rounded off to four S.F.
    (iv) If the digit removed is 5 and there is no digit after, then and one to the preceding digit if it is odd, otherwise write as such if it is even.

    Thus, 4.7475 becomes 4.748 when rounded off to four S.F (odd digit before 5) and 4.7465 becomes 4.746 when rounded off to four S.F. (even digit before 5)

1.4 Laws of chemical Combination

There are six laws of chemical combination.

(i) Law of conservation of mass (Given by Lavoisier)

According to this law during any physical or chemical change, total mass of reactants equals to the total mass of products.

This law is not applicable to nuclear reactions because in these reactions mass is converted into energy.

(ii) Law of definite proportions or constant proportions (Given by Proust)

According to this law, a compound always has same elements combine together in same fixed ratio by weight.

This law is not true to non-stoichiometric compounds.

(iii) Law of multiple proportions (Given by Dalton)

According to this law, when two elements combine to form two or more compounds, then the weight of one element which combines with the fixed weight of other has a simple ratio to one another.

(iv) Law of reciprocal proportions or equivalent proportions (Given by Pichter)

According to this law, when two elements combine with fixed weight of the third element, then it is either the same or simple multiple ratio of weights of two elements which combine directly with each other.

(v) Law of combining volumes (Given by Gay-Lussac)

According to this law, when gases react together, they always do so in volume which bears a simple ratio to one another and to volume of products, when all measurements are made under same conditions of temperature and pressure. The law is applicable to gases only.

(vi) **Avogadro's law (Mole concept)**

This law is based on Berzeilius hypothesis.

According to this law, all gases contain equal number of particles under similar conditions of temperature and pressure.

1 mole = atomic/ molecular wt in g

= 22.4 L at STP

= 6.02×10^{23} atoms/molecules /ions

Equal volumes of gases or vapours obeying gas laws under similar conditions of pressure and temperature contain equal number of molecules.

Molecular weight (for gaseous phase only)

= 2 x vapour density

1.4 Atomic Weight, Molecular Weight and Equivalent Weight

Atomic weight: It is a relative number not the exact weight of atom and it tells how many times an atom is heavier than 1/12 of one C-12 atom.

- Atomic mass unit (1 u) or unified mass unit (u)
  = 1 Aston = 1 Dalton =1.66×10^{-24} g
- Gram atomic mass is atomic masses of element expressed in grams.
  Average atomic mass = $\frac{x \times a + y \times b}{x+y}$
  (where x:y is the ratio of atomic masses of isotopes a and b).

Molecular weight: It is an additive property. It is calculated by adding atomic weights of all atoms present in the molecule.

Equivalent weight: It is the parts of a substance that combines with or displaces 1.008 parts by mass of hydrogen or 8 parts by mass of oxygen or 35.5 parts by mass of chlorine.

- Equivalent wt. (Eq. wt.)

$$= \frac{atomic\ wt.\ or\ molecular\ wt.}{'n'\ factor}$$

- 'n' factor for various compounds can be obtained as:
  1. 'n' factor for acids ie, basicity
     Number of ionisable H^+ per molecule is the basicity of acid. Eg.
     (a) Basicity of HCL = 1
     (b) Basicity of H_2SO_4 = 2
     (c) Basicity of H_3PO_4 = 3
     (d) Basicity of $H_2C_2O_4$ = 2

  2. 'n' factor for bases ie, acidity
     Number of ionisable OH^- per molecule is the acidity of a base. Eg.
     (a) Acidity of NaOH = 1

(b) Acidity of $Mg(OH)_2$ =2
(c) Acidity of $Al(OH)_3$ =3

3. 'n' factor for salt

Total positive or negative charge of ions.

(a) $Na_2CO_3 \longrightarrow 2Na^+ + CO_3^{2-}$ n=2
(b) $NaHCO_3 \longrightarrow Na^+ + HCO_3^-$ n=1
(c) $Al_2(SO_4)_3 \longrightarrow 2Al^{3+} + 3SO_4^{2-}$ n=6

4. 'n' factor for ion

In case of ion 'n' factor is equal to charge of that ion . eg.

$E_{Cl-} = \frac{35.5}{1} = 35.5$

$E_{CO_3^{2-}} = \frac{60}{2} = 30$

$E_{Al^{3+}} = \frac{27}{3} = 9.0$

5. 'n' factor for redox titration

(a) $FeSO_4$

⇨ As reducing agent

$Fe^{2+} \longrightarrow Fe^{3+} + e^-$ 'n' factor = 1

⇨ As an oxidizing agent

$Fe^{2+} + 2e^- \longrightarrow Fe(s)$ 'n' factor = 2

(b) $H_2C_2O_4$ or $C_2O_4^{2-}$

⇨ As reducing agent only

$C_2O_4^{2-} \longrightarrow 2CO_2 + 2e^-$

$H_2C_2O_4 \longrightarrow 2CO_2 + 2H^+ + 2e^-$ 'n' factor = 2

(c) HI

⇨ As reducing agent only

$HI \longrightarrow \frac{1}{2}I_2 + H^+ + e^-$ 'n' factor = 1

(d) $K_2Cr_2O_7$

⇨ As oxidizing agent only (acidic)

$Cr_2O_7^{2-} + 6e^- + 14H^+ \longrightarrow 2Cr^{3+} + 2H_2O$ 'n' factor = 6

(e) $KMnO_4$

⇨ As oxidizing agent in acidic medium

$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$ 'n' factor = 5

⇨ As oxidizing agent in alkaline

$MnO_4^- + 2H_2O + 3e^- \longrightarrow MnO_2 + 4OH^-$ 'n' factor = 3

(f) $Na_2S_2O_3$ (sodium thiosulphate)

⇨ As reducing agent in acidic medium

$S_2O_3^{2-} \longrightarrow \frac{1}{2} S_4O_6^{2-} + 1e^-$ 'n' factor = 1

⇨ As reducing agent in alkaline

$S_2O_3^{2-} + 10\ OH^- \longrightarrow 2SO_4^{2-} + 5H_2O + 8e^-$ 'n' factor = 8

(i) Equivalent weight of element

$= \frac{atomic\ weight}{valency}$

(ii) Equivalent weight of acid/base

$= \frac{molecular\ weight}{Basicity/acidity}$

(Where, Basicity = no. of replaceable H^+ and acidity = no. of replaceable OH^-

(iii) Equivalent weight of salt

$= \frac{molecular\ weight}{total\ positive\ or\ negative\ charges}$

(iv) Equivalent weight of oxidizing agent or reducing agent

$= \frac{Molecular\ weight\ of\ oxidising\ agent\ or\ reducing\ agent}{No\ of\ electrons\ gained/ost\ by\ one\ molecule}$

Methods for Determining Atomic Weight

- Atomic weight = equivalent weight × valency
- **Dulong and Petits method**

  This law is valid for metals only.

  Atomic weight x specific heat (in cal/g) ≈ 6.4

  Approximate atomic weight

  = 6.4 / Specific heat (in cal/g)

  The exact atomic weight is calculated with the help of valency as follows;

  (i) Valency = approximate atomic weight/Eq. wt.

  (ii) Exact atomic weight = Eq. wt. × valency.

  (iii) The method is used to determine atomic weight of solids except Be, B, C and Si.

- **Volatile chloride method :** This method is used to determine the atomic weights of elements which form volatile chloride Valency (of metal which = 2 × vapour density of metal chloride / Eq. wt. of metal forms volatile chloride)

(iv) **Specific heat method:** This method is used to determine the atomic weights of gases only.
Atomic weight of gas = molecular weight of gas / Atomicity

Methods for determining Molecular Weight

(i) Molecular weight = 2 × vapour density

(ii) $\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$ **(**Where, r_1 and r_2 are rate of diffusion of gases; M_1 and M_2 are molecular weights of gases.)

(iii) **Victor Meyer method:** The method is used to determine the molecular weight of volatile organic compounds only.
Molecular wt. of volatile organic compound = $\frac{\text{mass of volatile organic compound} \times 22400}{\text{volume of volatile organic compound}}$

(iv) **Ebullioscopic method:**
Molecular weight = $\frac{k_b \times wt.of\ solute \times 1000}{wt.of\ solvent \times elevation\ in\ boiling\ point}$

(v) **Cryoscopic method:**
Molecular weight = $\frac{K_f \times Wt.of\ solute \times 1000}{wt.of\ solvent \times depression\ in\ freezing\ point}$

Methods for Determining Equivalent Weight

(i) Eq. wt. = strength/ normality

(ii) **Hydrogen displacement method**
Eq. wt. of metal = $\frac{mass\ of\ metal}{mass\ of\ hydrogen\ displaced}$ × 1.008

(iii) **Oxide formation method**
Eq. wt. of metal = $\frac{mass\ of\ metal}{mass\ of\ oxygen\ combined}$ × 8.0

(iv) **Chloride formation method**
Eq. wt. of metal = $\frac{mass\ of\ metal}{mass\ of\ chlorine\ combined}$ × 35.5

(v) **Electrolytic method**
Eq. wt. = wt. deposited by 1F (= 96500 C)

$$\frac{W_1}{W_2} = \frac{E_1}{E_2}$$

(Where, W_1 and W_2 are weights of metal displaced, E_1 and E_2 are equivalent weights)

(vi) Metal displacement method

$$\frac{Mass\ of\ metal\ added}{Mass\ of\ metal\ displaced} = \frac{eq.wt.of\ metal\ added}{eq.wt.of\ metal\ displaced}$$

(vii) Double decomposition method

For a chemical reaction

AB + XY → AY↓+BX

$$\frac{wt.of\ comp.AB}{wt.of\ comp.AY} = \frac{eq.wt.of\ A + Eq.wt.of\ B}{eq.wt.of\ AY}$$

(viii) Silver salt method

The method is used to determine the equivalent weight of organic acids only

Eq. wt. of organic acid = $\frac{108 \times wt.\ of\ silver\ salt}{wt.\ of\ silver\ metal}$ - 107

1.5 Mole Concept

- A mole is defined as the number equal to the number of carbon atoms in exactly 12g of pure carbon-12 i.e. 6.023×10^{23}

  1 mole = 6.023×10^{23} atoms/molecules/ions

  1 mole = atomic/molecular wt. in g = 22.4 L at STP

  Number of moles = $\frac{number\ of\ particles}{6.023\times10^{23}}$
- The number 6.023×10^{23} was not determined by Avogadro but is called Avogadro's number to honour the contribution of this great chemist in chemistry.

Methods of Calculations of mole:

(a) If no. of some species is given, then no. of moles = $\frac{Given\ no.}{N_A}$

(b) If weight of a given species is given, then no. of moles = $\frac{Given\ wt.}{Atomic\ wt.}$ (for atoms.),

Or = $\frac{Given\ wt.}{Molecular\ wt.}$ (for molecules).

(c) If volume of a gas is given along with the temperature (T) and pressure (P) use

n = $\frac{PV}{RT}$

where R = 0.0821 lit-atm/mol-K (when P is in atmosphere and V is in liltre.)

1 mole of any gas as STP occupies 22.4 litre

1.6 Percent Composition, Empirical Formula and Molecular Formula

(i) **Percent composition** of a compound is expressed by identifying the elements present and giving the mass percent of each.

Percent composition of an element =

$$\frac{mass\ of\ that\ element}{total\ mass\ of\ compound} \times 100 \text{ (in 1 mol)}$$

Empirical formula is the simplest whole number ratio of all the elements present in one molecule of the substance.

(ii) Molecular formula: is the actual number of all the atoms of different elements present in one molecule of the substance.

- From percent composition, the empirical formula and then molecular formula is derived as follows

Element	% of each element	Atomic mass	Mol = $\frac{\%}{atomic\ mass}$ (relative no. of atoms)	Divided By lowest no.	Simple ratio

Molecular formula = (empirical formula)

Where,

$$n = \frac{molecular\ weight}{empirical\ formula\ weight}$$

Ex.Problem:

(i) A carbohydrate contains 40.0 % C, 6.70 % H and 53.3 % O and has a molecular weight of 180u. Calculate its molecular formula.

(Ans. $C_6H_{12}O_6$)

(ii) Calculate the empirical formula of the compound whose % composition is C = 21.9 %, H = 4.6 % and Br = 73.4 %.C = 12, H = 1, Br = 80u

(Ans.C_2H_5Br)

- 1 mole of water ≠22400cc (because water is liquid not gas)
- Loschmidt number = Number of molecules in 1 cm^3 of gas at STP = 2.688×10^{19}

1.7 Chemical Stoichiometry

Stoichiometry is the quantitative study of reactants and products in a chemical reaction. e.g.

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

For stoichiometric calculations, we would read this equation as "2 moles of $KClO_3$ decomposes to form 2 moles of KCl and 3 moles of O_2".

Limiting Reagent

It is the reactant which is present in least quantity and is consumed completely during the reaction. The amount of product formed depends on the limiting reagent.

1.8 Stoichiometry of Reaction in Solutions

Strength of solution is generally expressed in terms of molarity, normality, molality etc.

(i) Molarity is the number of moles of solute present in one litre of solution.

Molarity (M) = $\frac{moles\ of\ solute}{volume\ of\ solution\ in\ litre}$

Molarity equation: $M_1V_1 = M_2V_2$

(ii) Normality is the number of gram equivalents of solute present in one litre of solution.

Normality (N) = $\frac{gram\ equivalents\ of\ solute}{volume\ of\ solution\ in\ litre}$

Normality (N) = $N_1V_1 = N_2V_2$

(iii) Molality is the number of moles of solute present in 1 kg of solvent.

Molality (m) = $\frac{number\ of\ moles\ of\ solute}{mass\ of\ solvent\ (kg)}$

(iv) Mole fraction is the ratio of number of moles of one component to the total number of moles of all components in the solution.

$$X_A = \frac{n_A}{n_A + n_B}\ ;\ X_b = \frac{n_B}{n_A + n_B}\ ;\ n_A + n_B = 1$$

Parts per million, ppm (A)

(A) = $\frac{mass\ of\ A \times 10^6}{total\ mass\ of\ solution}$

Multiple Choice Questions

1. 10g $CaCO_3$ on heating leaves behind a residue weighing 5.6g. Carbon dioxide released into the atmosphere at STP will be
   (a) 2.24 L (b) 4.48 L
   (c) 1.12 L (d) 0.56 L
2. The equivalent weight of a metal is 4.5 and the molecular weight of its chloride is 80. The atomic weight of the metal is
   (a) 18 (b) 9
   (c) 4.5 (d) 36
3. 5.6 litre of a gas at NTP are found to have a mass of 11g. The molecular mass of the gas is
   (a) 22 (b) 44
   (c) 88 (d) 32
4. If 0.5 mol of $BaCl_2$ is mixed with 0.2 mol of Na_3PO_4, the maximum number of moles of $Ba_3(PO_4)_2$ that can be formed is
   (a) 0.7 (b) 0.5
   (c) 0.30 (d) 0.10

5. Boron has two stable isotopes, $^{10}B(19\%)$ and $^{11}B(81\%)$. The atomic mass that should appear for boron in the periodic table is

(a) 10.8 (b) 10.2

(c) 11.2 (d) 10.0

6. How many significant figures should be used for the answer to the following calculation?

$$\frac{(0.082056)(298.15)(0.379)}{(0.9480)}$$

(a) 2 (b) 3

(c) 4 (d) 5

7. A 5 molar solution of H_2SO_4 is diluted from 1 litre to a volume of 10 litres, the normality of the solution will be :

(a) 1 N (b) 0.1 N

(c) 5 N (d) 0.5 N

8. What is the weight of oxygen required for the complete of 2.8 kg of ethylene?

(a) 2.8 kg (b) 6.4 kg

(c) 9.6 kg (d) 96 kg