Some Basic Concepts of Chemistry

- Chemistry is the scientific study of the properties and behaviour of matter. It is the science of molecules and their transformation. It is an important part of our life. Formation of curd from milk, rusting of iron, formation of vinegar from sugarcane juice these are some examples of changes in chemical and physical properties of matter.

- The branch of science which deals with the preparation, properties, structure and chemical reactions of different matter is known as chemistry. Matter is made of atoms and molecules so we can say that chemistry is the science of atoms and molecules.

Importance and scope of chemistry

- Chemistry is the important branch of science, it plays key role in various areas like weather patterns, production of acid, alkalis, fertilizers, salt, dyes, polymers, medicines, detergents, metal alloys, etc.
- Chemistry provides extraction of life saving drugs from natural sources. Cisplatin and taxol are used for cancer therapy. AZT (Azidothymidine) is used for AIDS treatment.
- Chemistry majorly contributes to the development of a nation. With a better understanding of chemistry and its principles we can design and synthesize new environment friendly superconductors, polymers, optical fibres etc. with higher efficiency. Finding new and safer alternatives for CFCs (chlorofluorocarbon), which is major cause of ozone depletion.
- A developing country like India needs good chemist who can accept challenges like global warming, climate change and scarcity of natural resources like patrol, diesel, etc. To become a good chemist, one needs to understand basic concept of chemistry.

Nature of matter

- Anything which occupies space and has mass is known as matter.
- Matter exists in three physical states i.e., solid, liquid and gas on the basis of arrangement of the particles.
- Solids have definite shape and volume. Liquids have definite volume but lacking definite shape. They take the shape of container in which they are transferred. Gases have neither definite shape nor definite volume.
- All the three states of matter are inter-convertible under different temperature and pressure conditions.

$$\text{Solid} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{liquid} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{Gas}$$

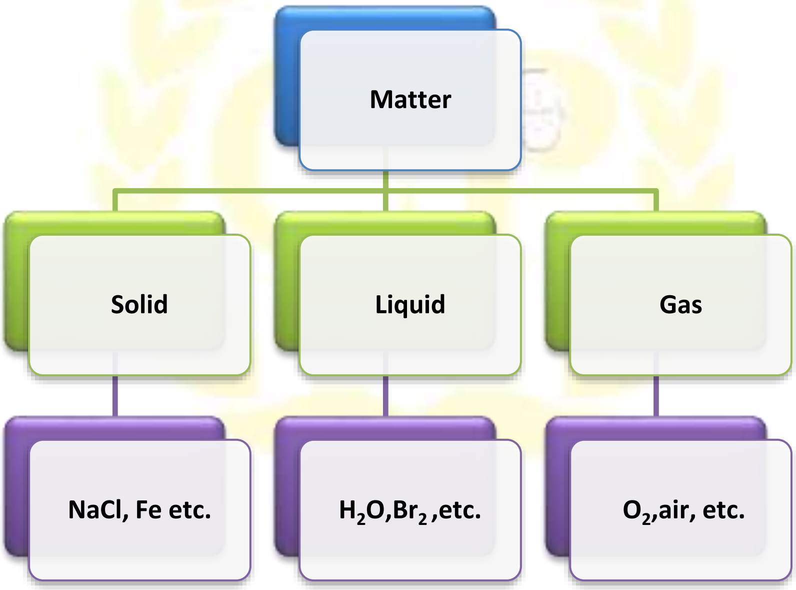


Figure: Physical classification of matter

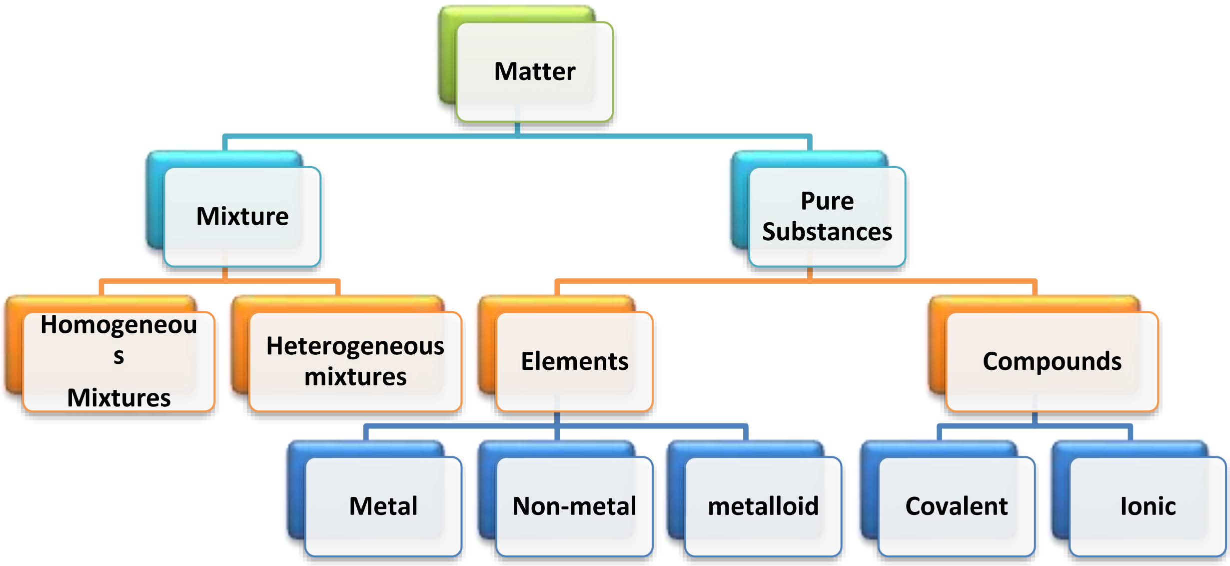


Figure: Chemical classification of matter

Properties of matter

- The properties of matter are classified in two categories i.e. physical and chemical properties.
- Physical properties are colour, odour, boiling point, melting point and density. Chemical properties like reactivity with acid and bases, combustion, composition etc.

Measurement of physical properties:

Table: Base physical quantities and their units in SI system

Base Physical Quantity	Symbol	SI unit	Symbol for SI unit
Length	*l*	Meter	m
Mass	m	Kilogram	Kg
Time	t	Second	s
Electric current	I	Ampere	A
Temperature	T	Kelvin	K
Luminous intensity	I_v	Candela	cd
Amount of substance	N	mole	mol

Laws of chemical combination

1. **Law of conservation of mass:**
   - This law was propounded by Antoine Lavoisier in 1789.
   - This law is experimented on combustion reactions.
   - It stated that in all chemical and physical changes there is no change in mass during the reaction. Matter cannot be created or destroyed. This is law of conservation of mass.

2. **Law of definite proportions**:
   - This law was propounded by Joseph Proust.
   - This law stated that a given compound always contains exactly the same proportion of elements by weight.
   - Proust worked on two samples of cupric carbonate, one of them was natural and other one was synthetic. He found that both samples contain same percentage composition of elements as given in table.
   - This law is also known as Law of definite composition.

	% of copper	**% of carbon**	**% of oxygen**
Natural Sample	**51.35**	**9.74**	**38.91**
Synthetic Sample	**51.35**	**9.74**	**38.91**

3. **Law of multiple proportions:**
   - This law was propounded by Dalton in 1803.
   - This law stated that if two elements can combine and form more than one compound then mass of one element that combine with a fixed mass of the other element will be in the ratio of whole numbers. For example, hydrogen react oxygen to form two compounds water and hydrogen peroxide. The masses of oxygen i.e., 16g and 32g combines with fixed mass of hydrogen in ratio of 16:32 or 1:2.

Hydrogen + Oxygen $\rightarrow$ Water

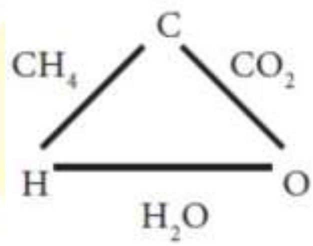
2g 16g 18g

Hydrogen + Oxygen $\rightarrow$ Hydrogen Peroxide

2g 32g 34g

4. **Law of reciprocal proportions:**
    - Law of reciprocal proportions is propounded by Jeremias Richter.
    - It is one of the basic laws of stoichiometry. It is also known as law of equivalent proportion or law of permanent ratio.
    - This law stated that if two different elements combine separately with a fixed mass of a third element, the ratio of the masses in which they combine are either same or in simple multiple ratio of the masses in which they combine with each other.
    - For example, oxygen and hydrogen reacts with carbon separately and forms carbon dioxide and methane respectively.

      The ratio of different weight of oxygen [32] and hydrogen [4] are combining with fixed weight of carbon [12] is 32:4 that is 8:1.



Sr. No.	Compounds	Combining elements		Combining weight	
1.	CH_4	C	H	12	4
2.	CO_2	C	O	12	32

5. **Gay Lussac's law of gaseous volumes:**
    - This law was propounded by Gay Loussac in 1808.
    - This law states that at constant volume the pressure of a gas increase as its temperature increases.

$$\mathbf{P \propto T}$$

$$\mathbf{P = KT}$$

Here P – Pressure exerted by gas

T – Absolute temperature of gas

K – Constant

- As we increase the temperature, the kinetic energy of gas increases. Eventually the collision of gas particle increases on the walls of container and pressure increase.
- For the same substance under two different conditions this law can be written as:

$$P_1/T_1 = P_2/T_2$$

- In our everyday life examples of application of Gay Lussac's law are pressure cooker, automobile tires, etc.
- This law is similar to the law of proportion.
- This law is for volume relationship.
- When gases combine, they combine in simple whole number ratio.
- These simple numbers are the coefficients of the balanced equation.
- For example:

1) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

Here three volumes of hydrogen will produce two volumes of ammonia.

2) $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

Here two volume of hydrogen reacts with one volume of oxygen and forms two volumes of water.

Dalton's Atomic Theory:

- John Dalton propounded this theory in 1808.
- According to his theory -
  1. Matter is consisting of small indivisible particles called atoms.
  2. All the atoms of an element are identical in mass, size and shape. Different elements have different atoms of different mass and size.
  3. When atoms of different elements combine in a fix ratio they form compounds.
  4. Chemical reaction involves rearrangement of reacting atoms. Atoms are neither created nor destroyed in a chemical reaction.
- **Limitations:**

1. It was failed to explain the laws of gaseous volumes, but it explained the laws of chemical combination.
2. According to this theory atom is indivisible but atom is made up of electrons, protons and neutrons.
3. According to Dalton atoms of an element are identical in mass and size but in case of isotopes, an atom with different mass exists.
4. Dalton theory could not explain the existence of allotropes.

- **Merits:**
  1. The law of conservation of mass, the law of multiple proportions and the law of constant proportions are not violated by Dalton's atomic theory.
  2. This theory helps to differentiate between elements and compounds.

Atomic mass

- Atomic mass means the mass of an atom.
- One atomic mass unit is defined as a mass equal to the one twelfth of the mass of one carbon-12 atom.

  Mass of one atom of an element =

  Atomic mass X (1/12) of the mass of C-12 atom

- It is measured in unit called amu (atomic mass unit).

$$1 \text{ amu} = 1.66056 \times 10^{-24} \text{ g}$$

- The mass of an atom of hydrogen = $1.6736\text{x } 10_{-24}$ g

  In terms of amu, the mass of hydrogen atom = $\dfrac{1.6736\text{x } 10_{-24}\text{ g}}{1.66056 \times 10^{-24}\text{ g}}$

  $= 1.0078 \text{ amu} \sim 1.0080 \text{ amu}$

- At present **amu** has been replaced by **u** which is known as **unified mass**.

Average atomic mass

- **Average atomic mass of an element** = (Fractional abundance of Isotope1 x atomic mass of Isotope1) + (Fractional abundance of Isotope2 x atomic mass of Isotope2)
- For example:

Isotope	Relative Abundance (%)	Atomic Mass (amu)
^{12}C	98.892	12
^{13}C	1.108	13.00335
^{14}C	$2x10^{-10}$	14.00317

Average atomic mass of carbon= (0.98892) (12) + (0.01108) (13.00335) + ($2x\ 10^{-12}$) (14.00317) = 12.011 amu

Molecular mass

- The sum of atomic masses of the elements present in a molecule is called molecular mass.
- Molecular mass= (Atomic mass of an element) (Number of atoms of that element) + (Atomic mass other element) (Number of atoms of that element)
- Molecular mass of **Methane (CH_4)**

  (CH_4) = 1 x atomic mass of carbon + 4 x atomic mass of hydrogen
  = 1 (12.011 amu) + 4 (1.008amu)
  = 16.043 amu
- Molecular mass of **water (H_2O)**

  (H_2O) = 2 x atomic mass of hydrogen +1 x atomic mass of oxygen
  = 2 (1.008amu) + 1 (16.00 amu)
  = 18.02 amu

Formula mass

- The mass of an individual molecule or formula unit is called formula mass.
- The formula mass of **Sodium chloride (NaCl)**

= Atomic mass of sodium + atomic mass of chloride

= 23.0 amu + 35.5amu = 58.5 amu

- The formula mass of **glucose ($C_6H_{12}O_6$)**

= 6 x (12.011amu) + 12 x (1.008 amu) + 6 x (16.00amu)

= 72.006amu + 12.096amu + 96.00amu = 180.162 amu

Mole concept and Molar mass

- Mole is the amount of a substance. This is a SI unit.
- 1 mole = $6.02214076 \times 10^{23} = N_A$

  Here N_A is Avogadro constant
- We can say that

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

  1 mol of NaCl = 6.022×10^{23} formula unit of sodium chloride
- The number of moles of a substance = n =

$$\mathbf{n = N/ N_A}$$

  Here n = number of moles

  N = number of particles

  N_A = Avogadro number = 6.022×10^{23}
- Number of moles of molecules (n) = $\dfrac{\text{Weight of substance}}{\text{Molecular Mass}}$

- Number of moles of atoms (n) = $\dfrac{\text{Weight of substance}}{\text{Atomic mass}}$

- Number of moles of gases (n) = $\dfrac{\text{Volume of gas in L at S.T.P.}}{22.4}$

  Here S.T.P. stands for standard temperature and pressure

- The mass of one mole of a substance in grams is called its molar mass. The molar mass in grams is numerically equal to atomic or molecular or formula mass in amu.

  For example –

  Molar mass of **Ethyne (C_2H_2)**

  = 2 x molar mass of carbon + 2 x molar mass of hydrogen

  = 2 x 12 + 2 x 1 = 24 + 2 = 26 g mol^{-1}

Molar mass of **Water** = H_2O = 2 x 1.008 + 1 x 16.00 = 18.02 g mol^{-1}

Molar mass of **Sodium chloride** = 23.0 + 35.5 = 58.5 g mol^{-1}

Percentage Composition

Mass % of an element =

$$\frac{\textbf{Mass of that element in the compound x 100}}{\textbf{Molar mass of the compound}}$$

Examples:

1. **Mass % of Hydrogen and Oxygen in Water (H_2O)**

Molar mass of water = 18.02 g

Mass % of hydrogen = $\frac{2 \times 1.008}{18.02} \times 100 = 11.18\%$

Mass % of oxygen = $\frac{16.00}{18.02} \times 100 = 88.79\%$

2. **What is the percentage of carbon, hydrogen and oxygen in ethanol?**

Molar mass of C_2H_5OH = (2x12.01 + 6 x 1.008 + 16.00) g = 46.068 g

Mass percent of carbon = $\frac{24.02}{46.068} \times 100 = 52.14\ \%$

Mass percent of hydrogen = $\frac{6.048}{46.068} \times 100 = 13.13\%$

Mass percent of oxygen = $\frac{16.00}{46.068} \times 100 = 34.73\%$

Empirical formula for molecular formula

- Empirical formula denotes the whole number ratio of atoms present in a compound where as molecular formula denotes the exact number of atoms present in a molecule of compound.

- If the mass percent of elements present in a compound are given, we can calculate its empirical formula. For example:

1. A compound's molar mass is 98.96g and contains

Hydrogen - 4.07%

Carbon – 24.27%

Chlorine – 71.65%

What are its empirical and molecular formulas?

Solution-

Convert mass percent to grams - since mass % is given in the question so 100 g of the compound is taken. Thus Hydrogen - 4.07g, Carbon – 24.27g and Chlorine – 71.65g.

Convert into number of moles-

Moles of hydrogen = $\frac{4.07g}{1.008\ g}$ = 4.04

Moles of carbon = $\frac{24.27g}{12.01g}$ = 2.021

Moles of chlorine = $\frac{71.65g}{35.453\ g}$ = 2.021

Divide mole values by the smallest number amongst them

Here 2.021 value is the smallest, so we get ratio of 2:1:1 foe H: C: Cl

Write down the empirical formula with the help of ratio

CH_2Cl

Write molecular formula

I. Empirical Formula mass of CH_2Cl = 12.01 + (2x 1.008) + 35.453 = 49.48g

II. Divide molar mass by Empirical Formula mass = $\frac{98.96g}{49.48g}$ = 2 = (n)

III. Multiply empirical formula by n =

Empirical formula CH_2Cl, n is 2.

Hence Molecular formula = $C_2H_4Cl_2$

Limiting reagent

- The reagent which consumes completely during the chemical reaction is known as limiting reagent. In other words, the reactant which gets consumed first and limits the amount of product formed is limiting reagent.

Concentration terms – The concentration of a solution can be expressed in any of the following ways:

1. Mass percent or weight percent (w/W%)

$$\text{Mass percent (w/W \%)} = \frac{\text{mass of solute}}{\text{Mass of solution}} \times 100$$

Example – if a solution is prepared by adding 2 g of a substance A to 18 g of water. Calculate the mass % of solute?

Solution- $\text{mass \% of A} = \frac{\text{Mass of A}}{\text{Mass of solution}} \times 100$

$$= \frac{2}{2+18} \times 100$$

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

2. Weight by volume percent (w/V%)

It is weight of solute dissolved in 100 ml of solution.

$$(w/V\%) = \frac{W_B}{V_A + V_B} \times 100$$

3. Mole fraction

- It is defined as the number of moles of a component present in the total number of moles of solution.
- If Substance A is dissolved in substance B. Their numbers of moles are n_A and n_B respectively.
- Then mole fraction of A = X_A

$$= \frac{n_A}{n_A + n_B}$$

- Mole fraction of B = X_B

$$= \frac{n_B}{n_A + n_B}$$

4. Molarity

- It is denoted by 'M'.
- It is the number of moles of solute dissolved in one litre of the solution.

$$\text{Molarity (M)} = \frac{\text{No. of moles of solute}}{\text{Volume of solution in litres}}$$

$$M = \frac{W_B}{M_B \times V\ (ml)} \times 1000$$

Here W_B is the weight of solute
M_B is the molar mass of solute
V is the volume of solution

- Example – 2 g of NaOH are dissolved in water and the solution is made to 500cm^3 in a volumetric flask. Find the molarity of the solution?
  Solution-

$$M = \frac{2}{40 \times 500} \times 1000 = 0.1\ M$$

5. Molality

- It is denoted by 'm'.
- It is the number of moles of solute present in 1 kilogram of solvent.

$$m = \frac{\text{no. of moles of solute}}{\text{Mass of solvent in Kg}}$$

$$m = \frac{n_B}{W_A\ (gm)} \times 1000$$

$$m = \frac{W_B}{M_B \times W_A\ (gm)} \times 1000$$

Here W_A =weight of solution in gm
W_B = weight of solute in g
M_B = molar mass of solute

- Example – What is the molality of solution which contains 18 gm of glucose in 250g of water?
  Molar mass of $C_6H_{12}O_6$= 180
  Solution-
  W_A = 250g
  W_B = 18g
  M_B = 180

$$m = \frac{18}{180 \times 250} \times 1000 = 0.4\ m$$

6. Normality

- It is denoted by "N".
- It is the number of gram equivalent of the solute dissolved in 1 litre of the solution.

$$N = \frac{\text{Number of gram equivalent of solute}}{\text{Volume of solution in litres}}$$

$$\text{Gram equivalent of solute} = \frac{\text{Mass of solute}}{\text{Equivalent mass of solute}}$$

$$\text{Equivalent mass} = \frac{\text{Molecular mass}}{\text{No. of } H^+ \text{ and } OH^- \text{ (Acidity or basicity)}}$$

Or

$$\text{Equivalent mass} = \frac{\text{Gram molecular mass}}{\text{Valency}}$$

$$N = \frac{W_B}{(\text{Gram equivalent})_B \times V\ (\text{ml})} \times 1000$$

- Here W_B stands for weight of solute
- The unit of normality is gram equivalent per litre.
- **Example**- 0.63 g of oxalic acid is dissolved in 250ml of a solution. Calculate the normality.

  Solution-

  Oxalic acid = $(COOH)_2 .2H_2O$, Molecular mass = 126

  $$\text{Equivalent mass} = \frac{\text{Molecular mass}}{\text{Basicity}} = \frac{126}{2} = 63$$

  $$\text{Gram equivalent} = \frac{0.63}{63} = 0.01$$

  $$N = \frac{0.01}{250} \times 1000 = 0.04\ N$$

Variable equivalent weight

- Few elements show variable valencies. The equivalent weight will be different for different valencies. Sulphur shows 2, 4, and 6 as valencies.

- Variable equivalent weight of sulphur

1. SCl_2

   $Cl_2 + S \rightarrow SCl_2$

   Here valency of S is 2.

   $$\text{Equivalent weight of S} = \frac{\text{Atomic mass}}{\text{Valency}} = \frac{32}{2} = 16\ g/eq$$

2. S_2Cl_2

Here valency of S is 1.

$$\text{Equivalent weight} = \frac{\text{Atomic mass}}{\text{Valency}} = \frac{32}{1} = 32\ g/eq$$

Equivalent weight of acid

1. H_2SO_4

   Molar mass of H_2SO_4 = 98 g mol^{-1}

$H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$

Basicity of H_2SO_4 (n_b) = 2

Equivalent mass = $\frac{M}{n_b} = \frac{98}{2}$ = 49 g/eq.

2. H_3PO_4
   Molar mass of H_3PO_4 = 98 g mol^{-1}

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

   H_3PO_4 gives 2 H^+ and these 2 H^+ reacts with 2 OH^- to form water molecule.

   So basicity of H_3PO_4 (n_b) = 2

   Equivalent mass = $\frac{M}{n_b} = \frac{98}{2}$ = 49 g/eq

Equivalent weight of base

1. NaOH

   $NaOH \rightarrow Na^+ + OH^-$
   So, acidity of NaOH = (n_a) = 1

   Equivalent weight of NaOH = $\frac{M}{n_a} = \frac{23+16+1}{1}$ = 40 g equi^{-1}
2. Ca (OH)$_2$

   $Ca(OH)_2 \rightarrow Ca^{+2} + 2OH^-$

   So, acidity of Ca $(OH)_2$ = (n_a) = 2

   Equivalent weight of Ca $(OH)_2$ = $\frac{M}{n_a} = \frac{40+(2x16)+2x1}{2} = \frac{74}{2}$ =37g equi^{-1}

Equivalent weight of salt

1. NaCl

   $NaCl \rightarrow Na^+ + Cl^-$

   $+1 \quad -1$

   Valency factor = total positive or negative charge of salt = 1

   $$\text{Equivalent weight} = \frac{\text{Molar mass}}{\text{Valency factor}} = \frac{58.5}{1} = 58.5 \text{ g/eq}$$

2. K_2SO_4

   $K_2SO_4 \rightarrow 2K^+ + SO_4^{2-}$

   $2x\,(+1) \quad -2$

   $+2$

   Valency factor = total positive or negative charge of salt = 2

   $$\text{Equivalent weight} = \frac{\text{Molar mass}}{\text{Valency factor}} = \frac{174}{2} = 87 \text{ g/eq}$$

Equivalent weight of oxidising and reducing agent

1. $2HNO_3 + 3H_2S \rightarrow 4H_2O + 2NO + 3S$

   $+5 \quad -2 \quad +2 \quad 0$

Here oxidising agent is HNO_3, it is reduced from +5 to +2 so net change is 3

Equivalent weight of oxidising agent

$$= \frac{\text{Molar mass}}{\text{Change in oxidation number}}$$

$$= \frac{63}{3} = 23 \text{ g/eq}$$

Here reducing agent is H_2S, it is oxidised from -2 to 0 so net changes is 2.

Equivalent weight of reducing agent i.e., $H_2S = \frac{34}{2} = 17$ g/eq

2. $10\ FeSO_4 + 2\ KMnO_4 + 8\ H_2SO_4 \rightarrow 5Fe_2(SO_4) + 2MnSO_4 + K_2SO_4 + 8\ H_20$

   +2 +7 +3 +2

Here oxidizing agent is $KMnO_4$, Mn reduced from +7 to +2 so net changes is 5.

Equivalent weight of oxidising agent $KMnO_4$

$$= \frac{\text{Molar mass}}{\text{Change in oxidation number}}$$

$$= \frac{158.04}{5} = 31.61 \text{ g/eq}$$

Here $FeSO_4$ is a reducing agent, Fe oxidised from +2 to +3 so net changes is 1.

Equivalent weight of reducing agent i.e., $FeSO_4 = \frac{151.92}{1} = 151.92$ g/eq