Some Basic Concepts of Chemistry

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CHAPTER

Matter and its Nature, Dalton's Atomic Theory; Concepts of Atom, Molecule, Element and Compound; Physical Quantities and their Measurements in Chemistry, Precision and Accuracy, Significant Figures, S.I. Units, Dimensional Analysis; Laws of Chemical Combination; Atomic and Molecular Masses, Mole Concept, Molar Mass, Percentage Composition, Empirical and Molecular Formulae; Chemical Equations and Stoichiometry.

THIS CHAPTER INCLUDES

- Laws of chemical combination
- Mole concept
- Measurement of concentration
- Limiting reagent
- Equivalent Weight
- n-factor or valence factor
- Laws of equivalence
- Empirical and molecular formula

LAWS OF CHEMICAL COMBINATION

There are five laws of chemical combination.

1. **Law of Conservation of Mass (Lavoisier 1774)**. It deals with the mass of reactants and products and states that in a chemical change the total mass of the products is equal to the total mass of the reactants. e.g.

$$\underset{12g}{C} + \underset{32g}{O_2} \rightarrow \underset{44g}{CO_2}$$

12g + 32g = 44g

2. **Law of Constant Composition** (**Proust 1799**). A chemical compound always contains same elements combined together in same proportion by mass e.g H_2O prepared from any source contains H and O in the ratio of 1 : 8 by mass.

3. **Law of Multiple Proportion** (**Dalton 1804**). When two elements combine to form two or more compounds then the masses of one of which combines with a fixed mass of the other element bears a simple whole number ratio to one another.

4. **Law of Reciprocal Proportion** (**Richter 1792**). It states that when two elements combine separately with a fixed mass of the third element then the ratio in which they do so is either the same or whole number multiple of the ratio in which they combine with each other.

5. **Gay Lussac's Law of Combining Volumes**. It states that at a given temperature and pressure, when the gases combine they do so in volumes which bear a simple ratio to each other and also to the volume of gaseous product e.g.

$$\underset{(1L)}{H_2(g)} + \underset{(1L)}{Cl_2(g)} \rightarrow \underset{(2L)}{2HCl(g)}$$. The ratio of their volumes is 1 : 1 : 2

MOLE CONCEPT

A mole is nothing but a collection of 6.023×10^{23} particles (atoms or molecules or ions) as well as it is equal to atomic weight (in g), molecular weight (in g) and ionic weight (in g) for atoms, molecules and ions respectively.

1 mole is represented in the forms of atoms, molecules and ions as:-

For atoms $\rightarrow$ 1 gm atom

For molecules $\rightarrow$ 1 gm molecule or 1 gm mole

For ions $\rightarrow$ 1 gm ion

Moles can be calculated by the following ways :

(a) Number of moles of molecules = $\dfrac{\text{Weight of substance (in g)}}{\text{Molecular weight}}$

(b) Number of moles of atoms = $\dfrac{\text{Weight of substance (in g)}}{\text{Atomic weight}}$

(c) Number of moles of gases = $\dfrac{\text{Volume of gas at NTP (in litres)}}{22.4}$

(1 mole of any gas occupies a volume of 22.4 litres at N.T.P., N.T.P. Corresponds to 0ºC and 1 atm pressure)

(d) Number of moles of atoms/molecules/ions

$$= \frac{\text{Number of atoms / molecules / ions}}{\text{Avogadro constant}}$$

(Avogadro constant is equal to 6.023×10^{23}).

MEASUREMENT OF CONCENTRATION

The concentration of a solution reflects the relative proportion of solute and solvent present in the solution. The various concentration terms are

(a) **% w/W** (weight percent or Mass percent)

x % w/W means that x g solute is present in 100 g of solution.

(b) **% w/V** — x % w/V means that x g of solute is present in 100 ml solution.

(c) **% v/V** — x % v/V means that x ml of solute is present in 100 ml solution.

(d) **Molality (m)** – It is defined as number of moles of solute present in 1 kilogram of solvent.

(e) **Molarity (M)** – It is defined as number of moles of solute present in 1 litre of solution.

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

Suppose, w gram solute is dissolved in V (in ml) solution and molecular weight of solute is m.

$$M = \frac{w/m}{V(\text{in ml})/1000} = \frac{w}{m} \times \frac{1000}{V(\text{in ml})}$$

$\dfrac{w}{m}$ = Moles of solute.

$\dfrac{w}{m} \times 1000 =$ Millimoles of solute.

So, millimoles of solute = M × V (in ml)

Moles of solute = M × V (in litre).

LIMITING REAGENT

In the given reaction if number of quantities (either in gm/mole/molecules) if are present with exact co-efficients then it is referred as, reactants are present in exact molar proportions required by chemical equation.

However if exact proportion is not present then the one which gets totally consumed is known as **limiting reagent** (Limiting reagent decides the product quantity for given information).

e.g., $2H_2 + O_2 \rightarrow 2H_2O$

In above *e.g.* 2 moles of H_2 reacts exactly with 1 mole of O_2 to give 2 moles of H_2O. If given moles of H_2 are 4 moles and that of O_2 are 0.5, then 0.5 O_2 will act as limiting reagent as it is in minimum amount and product formation is given w.r.t. O_2 *i.e.*, 1 mole of H_2O.

EQUIVALENT WEIGHT

Equivalent weight of substance is defined as number of parts by weight of given substance which combines or displaces 1 part by weight of hydrogen (11.2 L of H_2 at STP), 8 parts by weight of oxygen (5.6 L at STP), 35.5 parts by weight of chlorine (11.2 L at STP).

$$\text{Equivalent weight of element} = \frac{\text{Atomic weight}}{\text{Valency}}$$

$$\text{Equivalent weight of acids} = \frac{\text{Molecular weight of acid}}{\text{Basicity}}$$

$$\text{Equivalent weight of bases} = \frac{\text{Molecular weight of base}}{\text{Acidity}}$$

$$\text{Equivalent weight of salts} = \frac{\text{Molecular weight of salt}}{\text{Total} + \text{ve or} - \text{ve charge}}$$

n-FACTOR OR VALENCE FACTOR

n-factor is very important for both redox and non redox reactions through which we predict the following two informations:

(a) It predicts the molar ratio of the species taking part in reactions i.e. reactants. The reciprocal of n-factor's ratio of the reactants is the molar ratio of the reactants.

For example : If X (having n-factor = a) reacts with Y (having n-factor = b) then its n-factor's ratio is a : b, so molar ratio of X to Y is b : a.

It can be represented as $\underset{(nf=a)}{bX} + \underset{(nf=b)}{aY} \longrightarrow \text{Products}$

(b) $\text{Equivalent weight} = \dfrac{\text{Molecular weight}}{n-\text{factor}}$ or $\dfrac{\text{Atomic weight}}{n-\text{factor}}$

LAWS OF EQUIVALENCE

According to law of equivalence, for each and every reactant and product,

Equivalents of each reactant reacted = Equivalents of each product formed.

For example :

Suppose, the reaction is taking place as under

$A + B \rightarrow C + D$.

Then according to law of equivalence,

Equivalents of A reacted = Equivalents of B reacted

= Equivalents of C produced

= Equivalents of D produced

$$\text{Equivalents of any substance} = \frac{\text{Weight of substance (in g)}}{\text{Equivalent weight}}$$

= Normality (N) × Volume (V) (In litre)

Normality (N) = n-factor × Molarity (M)

Normality and molarity are temperature dependent. Since on changing the temperature, the volume of solution changes, so normality and molarity change.

EMPIRICAL AND MOLECULAR FORMULA

(a) **Empirical Formula of a compound** is the simplest whole number ratio of the atoms of elements constituting its one molecule. The sum of atomic masses of the atoms representing empirical formula is called **empirical** formula mass.

(b) **Molecular Formula** of a compound shows the actual number of the atoms of the elements present in its one molecule. The sum of atomic masses of the atoms representing molecule is called **molecular mass**.

(c) **Relationship between Empirical Formula and Molecular Formula**

Molecular formula = n × empirical formula where n is a simple whole number having values of 1, 2, 3... etc.

Also, n = Molecular formula mass/Empirical formula mass.

Weigh sample → Weigh CO_2 and H_2O → Calculate moles of C and H → Calculate C : H mole ratio → Calculate empirical formula