CHAPTER-2

CHEMICAL COMBINATIONS

You will learn in this chapter about :

* Laws of chemical combinations.
* Atomic mass.
* Chemical formula.
* Empirical formula.
* Molecular formula.
* Molecular mass and formula mass.
* Mole.
* Avogadro's number.
* Chemical reactions and chemical changes.
* Chemical reactions and types of chemical reactions.
* Chemical equations.
* Writing of chemical equations.
* The balancing of chemical equations.
* Concept of mole ratios and balanced chemical equations.
* Calculations based on chemical equations.

2.1 LAWS OF CHEMICAL COMBINATIONS

Introduction:-

Chemistry deals with the matter and the changes occurring in it, Chemists are particularly interested in those changes, where one or more substances are changed into quite different substances. They had found that these chemical changes are governed by some empirical laws known as laws of chemical combinations.

These laws are:

1. Law of conservation of mass

2. Law of constant composition (or)
   Law of definite proportions
3. Law of multiple proportions
4. Law of reciprocal proportions.

1. Law of Conservation of Mass:

Matter under goes changes. However, it has been found that in all chemical changes, there is no change in the mass of the substances being changed. For example, in iron (Fe) increase in weight on rusting is because of its combination with oxygen from the air and the increase in weight is exactly equal to the weight of oxygen combined. The French Chemist Lavosier, (1785) tried to learn about chemical changes by weighing the quantities of substances used in chemical reactions. He found that when a chemical reaction was carried out in a closed system, the total weight of the system was not changed. The most important chemical reaction that Lavosier performed was the decomposition of the red oxide of mercury to form metallic mercury and a gas, he named this gas as oxygen. Lavosier summarised his findings by formulating a law, which is known as **law of conservation of mass**. It states that mass is neither created nor destroyed during a chemical reaction. In other words , *In any chemical reaction the initial weight of reacting substances is equal to the final weight of the products.*

The law of conservation of mass may be demonstrated by the union of hydrogen (H_2) and oxygen (O_2) to form water. If the H_2 and O_2 are weighed before they unite, it will be found that their combined weight is equal to the weight of water (H_2O) formed. This is shown below in fig.2.1.

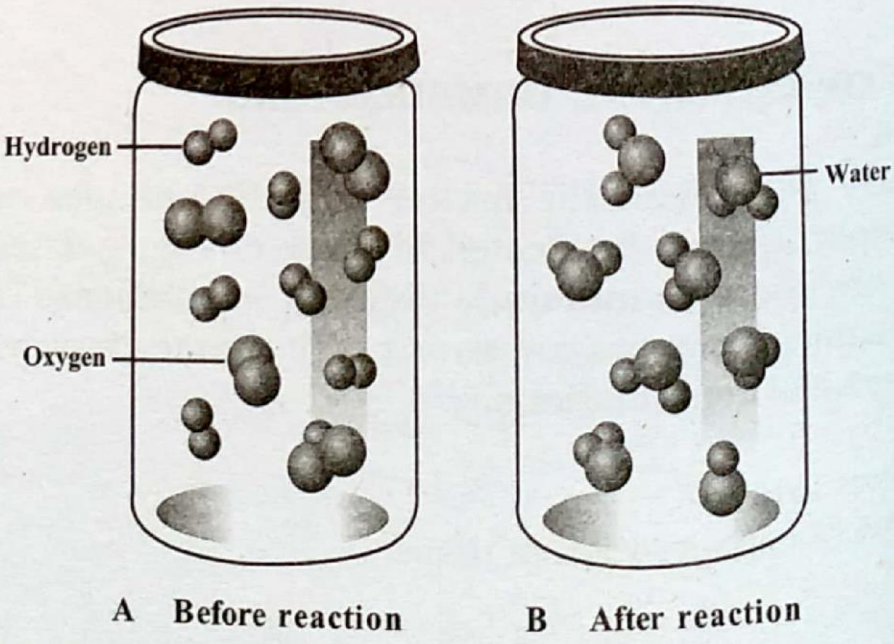

Fig: 2.1

Practical Verification: (Landolt Experiment)

German Chemist H. Landolt, studied about fifteen different chemical reactions with a great skill, to test the validity of the law of conservation of mass. For this, he took H-shaped tube as shown in fig. 2.2 and filled the two limbs A and B, with silver nitrate (Ag NO_3) in limb A and hydrochloric acid (HCl) in limb B. The tube was sealed so that the material could not escape outside. The tube was weighed initially in a vertical position so that the solutions should not intermix with each other. The reactants were mixed by inverting and shaking the tube. The tube was weighed after mixing (on the formation of white precipitate of (AgCl). He observed that weight remains same.

$$Ag\,NO_3(aq) + HCl(aq) \longrightarrow \underset{\text{white ppt:}}{AgCl(s)} + HNO_3(aq)$$

Thus total mass of the substance before the reaction is equal to the total mass of the substances after the reaction.

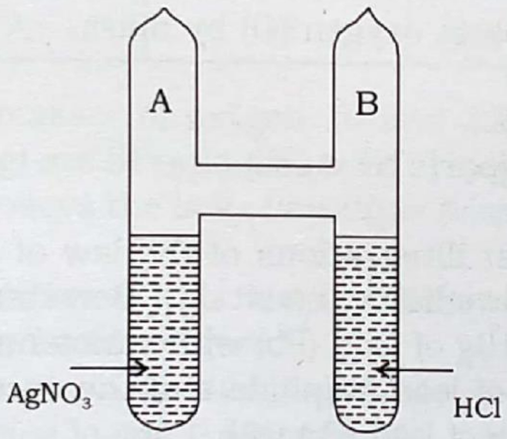

Fig: 2.2

In ordinary chemical changes, relatively small amount of energy is released. But in nuclear changes where uranium atoms undergo fission (break up) into smaller atoms plus neutrons, the total mass of products is noticeably less than that of starting material. This clearly indicates that some mass of uranium has been converted into energy, which is evident to us as heat and radiation.

The relationship between mass that is lost and the energy that is released is given by the equation.

$$E = mc^2$$

Where (E) is the energy in ergs, (m) is the mass in grams and (C) is the velocity of light in centimeters per second, ($3x10^{10}$ cm/sec). This relationship between mass and energy was first proposed in (1906), by the famous Physicist and Mathematician, Albert Einstien.

It follows that for every chemical change, there will be a mass change.

But this mass change is too small that no one has yet been able to detect through weighing techniques.

Hence the law of conservation of mass is, therefore, still valid from practical view point for ordinary chemical reactions i.e., *"there is no detectable gain or loss of mass in a chemical reaction"*.

2. Law of Constant Composition or Law of Definite (fixed) Proportion

By the end of Eighteenth century, chemists showed that a given compound has a definite (constant) composition. French chemist Louis Proust in (1799) summarized this result in the form of the law of definite proportion (also known as constant composition) which states, that different samples of the same compound always contain the same elements combined together in the same proportions by mass.

For instance every sample of pure water, though prepared in the laboratory or obtained from rain, river or water pump contains one part hydrogen (H) and (8) parts oxygen (O) by mass

eg: H_2O
2:16
1:8 (parts by mass)

One of the earliest illustrations of the law of definite proportions is found in the work of Swedish chemist J. J Berzelius (1779-1848).

Berzelius heated 10g of lead (Pb) with various amounts of sulphur (S). He got exactly 11.56g of lead sulphide and the excess of sulphur was left over, when he used 18g of lead (Pb) with 1.56g of sulphur (S), he got exactly 11.56g of lead sulphide (PbS) and the 8g of lead (Pb) remained unused.

These reactions are shown diagrammatically in Fig. 2.3

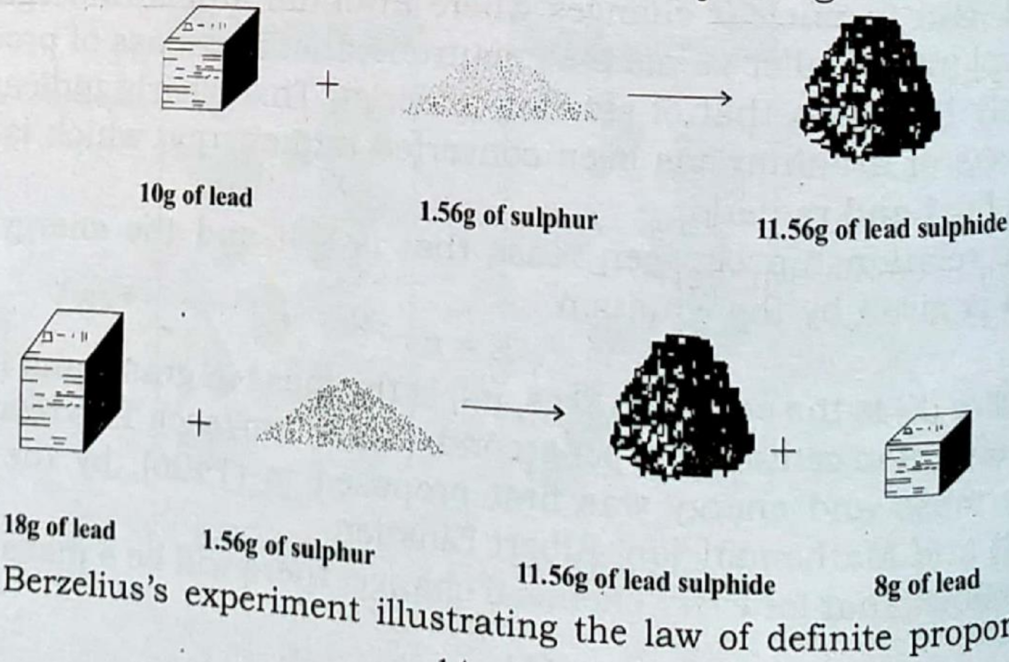

Fig. 2.3 Berzelius's experiment illustrating the law of definite proportions.

3. Law of Multiple Proportion:

The fact that the same element, can combine in more than one ratio to form different compounds was published by, John Dalton, (1803) in the form of law of multiple proportion. "*It states that if two elements combine to form more than one compounds, the masses of one element that combine with a fixed mass of the other element are in the ratio of small whole numbers or simple multiple ratio.*"

For example: Carbon (C) forms two stable compounds with oxygen (O) namely carbon monoxide (CO) and carbon dioxide (CO_2).

Compound	Mass of Carbon (C)	Mass of Oxygen (O)	Ratio of Oxygen (O)
Carbon monoxide CO Carbon dioxide CO_2	12 12	16 32	1 2

The different masses of oxygen 16 and 32 which combine with the fixed mass of C (12g) are in ratio of [16:32], i.e. 1:2, which is simple whole number ratio, and obeys the law of multiple proportion.

Another illustration of this law is the formation of water (H_2O) and (H_2O_2) from hydrogen and oxygen.

Compound	Mass of Hydrogen (H)	Mass of Oxygen (O)	Ratio of Oxygen (O)
Water H_2O Hydrogen peroxide (H_2O_2)	2 2	16 32	1 2

The different masses of oxygen 16:32, which combine with the fixed mass of hydrogen (2g) are in ratio of 16:32 i. e. 1:2 which is again in a ratio of simple whole numbers.

The excellent illustration of law of multiple proportion is furnished, when the elements nitrogen (N) and oxygen (O) combine together to form a series of five oxides of nitrogen, in which these two elements are present in different proportions.

S.No.	Name of Oxides	Mass of (N)	Mass of (O)	Fixed mass of (N)	Variable mass of (O)	Ratio of (O)
1.	Nitrous oxide (N_2O)	28	16	14	8 (1x8)	1
2.	Nitric oxide (NO)	14	16	14	16 (2x8)	2
3.	Nitrogen trioxide (N_2O_3)	28	48	14	24 (3x8)	3
4.	Nitrogen tetra oxide (N_2O_4)	28	64	14	32 (4x8)	4
5.	Nitrogen penta oxide (N_2O_5)	28	80	14	40 (5x8)	5

By fixing the mass of (N), the mass of (O) in different oxides varies.

i.e. 8 : 16 : 24 : 32 : 40
1 : 2 : 3 : 4 : 5

These figures (in multiple ratios), are according to the law of multiple proportion.

4. Law of Reciprocal Proportion:

This law was enunciated by Ritcher in (1792-94). *It states that "when two different elements separately combine with the fixed mass of third element, the proportions in which they combine with one another shall be either in the same ratio or some simple multiple of it".*

For instance, when two elements C and O separately combine with H to form methane (CH_4) and water (H_2O) respectively it is very clear, that in methane 3g of C combine with 1g of hydrogen, and in water (H_2O) 8g of O combine with the same (fixed) mass i.e (1g) of H. Now when C and O combine with each other to form carbon dioxide (CO_2), they do so in the same proportion i. e. 12:32 = 3:8 parts by mass.

H

CH_4 1g H_2O

12 : 4   2 : 16

3 : 1   1 : 8

3g   8g

C.................O

CO_2

12 : 32

i.e. 3 : 8

Another illustration of law of reciprocal proportion is provided when, 12g of C combine with 32g of O to form carbon dioxide (CO_2) and 32g of sulphur (S) combine with the same (fixed) mass of oxygen (O) i. e. 32 g to form sulphur dioxide.

O

CO_2 1g SO_2

12 : 32   32 : 32

3 : 8   8 : 8

12g   32g

C.................S

CS_2

12 : 64

i.e. 3 : 16

The above example, shows that the mass of C and S that combine with the same mass of O are in the proportion of 12:64 i. e. 3:16.

According to the statement of law of reciprocal proportion, the proportion in which C and S combine with one another shall be either in the same ratio (3:8) or some simple multiple of it i. e (3:16).

It is very clear that in the formation of carbon disulphide (CS_2), C and S combine in the ratio of (12:64) i.e (3:16) which is simple multiple of (3:8).

2.2 ATOMIC MASS

The mass of an atom depends upon the number of protons and neutrons present in it. As the atoms are extremely small particles, it is difficult to weigh them directly. For example the mass of single hydrogen (H) atom, is 1.6×10^{-24}g (0.000 000 000 000 000 000 000 0016g). Clearly we cannot weigh a hydrogen atom or any other kind of atom, by placing it on a balance pan. Scientist need special method to obtain the mass of an atom by comparison to a standard mass.

In 1961, by an international agreement, an atom of C-12, that has 6 protons and 6 neutrons has a mass of exactly 12 atomic mass units (a.m.u.) is taken as a standard. So one atomic mass unit (1 a.m.u.) is defined as a mass exactly equal to one-twelfth the mass of C-12 atom. Since most elements consist of a mixture of isotopes (Isotopes are the atoms of same element, having same atomic number but different atomic masses). For example, naturally occurring carbon is composed of 98.889 percent C-12 and 1.111 percent C-13. Thus average atomic mass of C-atom becomes 12.011 a.m.u.

The atomic mass of an element is now taken as, the average mass of natural mixture of isotopes which is compared to the mass of one atom of C-12 a.m.u.

Thus the atomic mass of oxygen (O)=16 a.m.u. and that of sulphur (S) = 32 a.m.u.

2.2.1 Empirical Formula and Molecular Formula: (E.F and M.F)

A formula is a combination of symbols for atoms or ions, that are held together chemically in a compound. By formula we mean not only the elements present but also ratios in which the atoms are combined. Hence we will discuss two types of formulas i.e. Empirical formula and Molecular formula.

2.2.1(a) Empirical Formula (E.F): (Simplest formula).

A formula that gives only the relative number of each type of atoms present in a molecule. In other words, the empirical formula does not necessarily give the actual number of atoms in a molecule. For example, the molecular formula of benzene is C_6H_6. This formula indicates that benzene molecule consists of (6) carbon atoms and (6) hydrogen atoms. The ratio of carbon (C) to hydrogen (H) atoms in this molecule is 6:6 or 1:1. The empirical formula of benzene is, therefore written as (CH).

Thus empirical formula tells us which elements are present and their simplest atomic ratio, but not necessarily the actual number of atoms present in the molecule.

Consider another example, the molecule of glucose ($C_6H_{12}O_6$) in which the ratio of C, H and O atoms is 6:12:6 i.e. 1:2:1. The empirical formula of glucose is, therefore, (CH_2O).

2.2.1 (b) Molecular Formula (M. F):

Molecular formula indicates the actual number and type of atoms in a molecule. It can either be same as empirical formula or some simple multiple of it. Mathematically, M. F = (E.F) n, where (n) is the whole number. For example the molecule of CO_2 consists of one atom of carbon in combination with two atoms of oxygen. The formula (CO_2) is the molecular formula of carbon dioxide. It represents the true composition of a molecule of the compound. The molecular formula may be same as empirical formula as in the case of CO_2 or some simple multiple of empirical formula. Thus the molecule of glucose which shows that the molecule of glucose, consists of (6) carbon, (12) hydrogen and (6) oxygen atoms and its simplest atomic ratio i.e empirical formula is (CH_2O). Thus the molecular formula of glucose ($C_6H_{12}O_6$) is equal to $(CH_2O)_6$ or six times to empirical formula (CH_2O).

It follows that molecular formula of glucose is six times the empirical formula, which is obtained by M. F = $(E.\ F)_n$ where (n) is the whole number, and in glucose n=6.

Mathematically,

$$n = \frac{\text{M.F. weight}}{\text{E.F. weight}}$$

For many molecules, the molecular formula and empirical formula are the same, some examples are formaldehyde (CH_2O), ammonia (NH_3) and methane (CH_4).

2.2.2 Molecular Formula Mass:

The molecular formula mass (molecular mass) of a substance is the sum of the atomic masses of all atoms present in the molecular formula of a substance or molecule. Taking as an examples, let us calculate the molecular formula mass of CO_2. The molecule of CO_2 contains one atom of C and two atoms of O. The atomic masses of C and O are 12 a. m. u. and 16 a. m. u, respectively.

$$\begin{aligned} C &= 12 \times 1 &&= 12 \text{ a.m.u.} \\ O &= 16 \times 2 &&= \underline{32 \text{ a.m.u.}} \\ \text{Molecular formula mass of } CO_2 & &&= 44 \text{ a.m.u.} \end{aligned}$$

For example, Compute the molecular formula mass of ozone (O_3). Molecular formula mass is calculated by adding together the atomic masses of the

constituent atoms. The ozone (O_3) molecule contains three oxygen atoms each of which has mass equal to 16a.m.u.
Therefore molecular formula mass of ozone (O_3) = 3 x 16 = 48 a.m.u.

2.2.3 Formula Mass:

Formula mass of substances is the sum of the atomic masses of all atoms in a formula unit of the substance. For example, we can calculate the formula mass of sodium chloride (NaCl), a common salt, by adding the atomic masses of all atoms in the formula unit, expressed in (a.m.u.). the atomic masses of (Na) and (Cl) are 23 a.m.u. and 35.5 a.m.u., respectively.

$$\begin{aligned} Na = 23 \times 1 &= 23 \text{ a.m.u.} \\ Cl = 35.5 \times 1 &= \underline{35.5 \text{ a.m.u.}} \\ \therefore \text{Formula mass of NaCl} &= 58.5 \text{ a.m.u.} \end{aligned}$$

Remember

That the term molecular mass applies to molecular compounds. The term formula mass can be used with either molecular compounds or ionic compounds. The term molecular mass can not be used with ionic compounds because there are no discrete molecules in ionic compounds.

2.2.4 Molar Mass:

Molar mass of a substance is its relative molecular mass expressed in grams. Thus molar mass of a substance has a fixed unit. For example, 1 mole of (C) is equal to its atomic mass expressed in grams.

Molecular mass of C = 12 a.m.u. and therefore the molar mass of carbon would be 12g.

For example, calculate the molar mass of ammonia (NH_3).

The molar mass, is obtained by adding the atomic masses of component atoms.

$$\begin{aligned} N &= 1 \times 14g = 14g \\ H &= 3 \times 1g = \underline{3g} \\ \text{Molar mass of } NH_3 &= 17g \end{aligned}$$

Remember, that relative molecular mass of NH_3 = 17 a.m.u.

2.3 MOLE

(A practical chemical unit for handling atoms and molecules).

Since atoms and molecules are so small, it is impossible to handle and count atoms and molecules individually. Therefore, the Chemists devised a special unit to describe very large number of atoms, ions and molecules.

It is called mole and is abbreviated as mol. A mole can be defined as *"the molecular mass, atomic mass and formula mass of a substance expressed in grams. Thus, 12g of carbon is equal to 1 mol of carbon atoms. 24g of C is equal to 2 mol of carbon atoms"*. The mole concept tries to give a practical meaning to the mass of reactants and products in chemical reactions in terms of the number of particles (atoms, molecules or ions) involved. The actual countable number of particles in one mole of a substance is 6.02×10^{23} particles, which is referred as Avogadro number.

The (S.I.) definition of mole is the amount of substance, containing as many elementary particles (units) as there are atoms in exactly 12g of C-12 a.m.u.

It is also defined as the mass of any substance equal to its atomic mass, molecular mass or formula mass in grams.

Thus.

1 mole of C	=	12 g
1 mole of Mg	=	24 g
1 mole of H_2O	=	18 g
1 mole of CO_2	=	44 g
1 mole of $CaCO_3$	=	100g
1 mole of Fe_2O_3	=	160g

By formula, number of moles = $\dfrac{\text{Given mass of substances}}{\text{Atomic mass or formula mass}}$

As, formula mass, represents the both covalent and electrovalent compounds.

2.3.1 Avogadro's Number (N_A):

A mole of substance always contains the same number of particles (*atoms, ions, molecules or formula units*) irrespective of its state, solid, liquid or gaseous, that is 6.02×10^{23} particles. This constant number has been determined by several methods, called Avogadro's number (symbol N_A),in the honour of Avogadro, the scientist who gave chemistry a method for finding atomic and molecular masses.

Thus.

1 mole of C	=	12 g	=	6.02×10^{23} atoms of carbon
1 mole of Mg	=	24 g	=	6.02×10^{23} atoms of magnesium
1 mole of H_2O	=	18 g	=	6.02×10^{23} molecules of water
1 mole of CO_2	=	44 g	=	6.02×10^{23} molecules of CO_2

1 mole of $NaCl$	=	58.5g	=	6.02×10^{23} F-units of $NaCl$
1 mole of $CaCO_3$	=	100g	=	6.02×10^{23} F-units of $CaCO_3$
1 mole of Na^+	=	23g	=	6.02×10^{23} ions of Na^+
1 mole of Cl^-	=	35.5g	=	6.02×10^{23} ions of Cl^-

Conversion of Mass into Moles and Moles into Mass of Substance

Problem 1. Calculate the number of moles, in 50 g of each.

(a) Na (b) H_2O

Solution:

$$\text{Number of moles} = \frac{\text{Given mass of substance}}{\text{Atomic mass or Formula mass}}$$

(a) Given,

i) Number of moles = ?
ii) Given mass of Na = 50g
iii) Atomic mass of Na = 23 a.m.u.

$$\therefore \text{Number of moles of Na} = \frac{50}{23} = 2.173 \text{ moles of Na}$$

(b) Given,

i) Number of moles of H_2O = ?
ii) Formula mass of H_2O = 18 a.m.u.
iii) Given mass of H_2O = 50g

$$\therefore \text{Number of moles of } H_2O = \frac{50}{18} = 2.777 \text{ moles of } H_2O$$

Problem 2. What is the mass of 3 moles of each:

a) Al b) CO_2.

$$\text{Number of moles} = \frac{\text{Given mass of substance}}{\text{Atomic mass or Formula mass}}$$

$\therefore$ Mass of substance = number of moles x atomic mass or formula mass in grams

(a) Given,

i) Number of moles of Al = 3 moles
ii) Atomic mass of Al = 27 grams
iii) Mass of Al = Mole x atomic mass of Al
= 3 x 27=81g

(b) Given,

i) Number of moles of CO_2 = 3 moles

ii) Formula mass of CO_2 = 44g

iii) Mass of CO_2 = Mole x Formula mass of CO_2

= $3 \times 44 = 132$ g

Use of Avogadro's number:

a) To calculate the number of atoms or molecules in a given sample of substance.

b) To calculate the mass of single atom or molecule of any substance.

Problem: 1. Calculate the number of atoms in 9g of Al.

Solution: According to Avogadro's number.

1 mole of Al= 27g = 6.02×10^{23} atoms.

This shows that:

27 g of Al contain 6.02×10^{23} atoms of Al

1 g of Al will contain $\frac{6.02 \times 10^{23}}{27g}$

9g of Al will contain $\frac{6.02 \times 10^{23} \times 9}{27g} = 2.006 \times 10^{23}$ atoms of Al.

This numerical can also be solved, by using the formula.

$$\text{Number of atoms} = \frac{N_A \times \text{Mass of substance}}{\text{Atomic mass}}$$

$$\text{Number of (Al)atoms} = \frac{6.02 \times 10^{23} \times 9g}{27g}$$

$= 2.006 \times 10^{23}$ atoms of Al.

Problem: 2. Calculate the number of molecules in 9g of CO_2.

Solution: According to Avogadro's number

1 mole of CO_2 = 44g = 6.02×10^{23} molecules

This shows that:

44 g of CO_2 contain 6.02×10^{23} molecules of CO_2

$$1 \text{ g of } CO_2 \text{ contain } \frac{6.02\times10^{23}}{44g} \text{ molecules}$$

$$9 \text{ g of } CO_2 \text{ will contain } \frac{6.02\times10^{23}\times9g}{44g}$$

$$= 1.231\text{x}10^{23} \text{ molecules of } CO_2$$

By formula

$$\text{Number of molecules} = \frac{N_A \times \text{Mass of substance}}{\text{Formula mass}}$$

$$\text{Number of } (CO_2) \text{ molecules} = \frac{6.02\times10^{23}\times9g}{44g}$$

$$= 1.231 \text{ x } 10^{23} \text{ molecules of } CO_2$$

Problem: 3.

Calculate the mass of one atom of carbon in grams

Solution: According to Avogadro's number

1 mole of C = 12g = $6.02 \text{ x } 10^{23}$ atoms

this indicates that

$6.02\text{x}10^{23}$ atoms of C weigh 12 g

$$1 \text{ atom of C} = \frac{12 \text{ g}}{6.02\times10^{23}} = \frac{12 \text{ g}\times10^{-23}}{6.02}$$

$$= 1.993 \text{ x } 10^{-23} \text{ g}$$

Mass of one C-atom $= 1.993 \text{ x } 10^{-23}$ g

By formula

$$\text{Mass of one atom} = \frac{\text{Atomic mass in grams}}{N_A}$$

$$\text{Mass of one (C)atom} = \frac{12 \text{ g}}{6.02\times10^{23}} = 1.993 \text{ x } 10^{-23} \text{ g}$$

Problem: 4.

Calculate the mass of one molecule of water (H_2O) in gram.

Solution:

According to Avogadro's number

1 mole of H_2O = 18g = 6.02×10^{23} molecules

This indicates that

6.02×10^{23} molecules of (H_2O) weigh 18g

1 molecule of (H_2O) weighs $\frac{18\ g}{6.02 \times 10^{23}} = \frac{18\ g \times 10^{-23}}{6.02} = 2.90 \times 10^{-23} g$

By formula

$$\text{Mass of one molecule} = \frac{\text{Formula mass in grams}}{N_A}$$

$$\text{Mass of one } (H_2O) \text{ molecule} = \frac{18\ g}{6.02 \times 10^{23}} = 2.90 \times 10^{-23}\ g$$

Mass of one (H_2O) molecule in grams = 2.90×10^{-23} g

2.4 CHEMICAL REACTION OR CHEMICAL CHANGE

Any change, which alters the composition of a substances, is a chemical change. In this type of change one. or more new substances are formed from the original substances, for example, when iron (Fe) rusts, it reacts with oxygen (O) of air in presence of moisture to form red brown iron oxide (rust). Similarly, when coal burns, it forms smoke, gaseous products and ashes. The burning of coal is a chemical reaction (change)in which it combines with oxygen in air to form entirely new substances.

2.4.1 Types of Chemical Reactions:

Chemical reactions can be divided commonly into five different types.

1. Decomposition reactions.
2. Addition reaction (combination reaction).
3. Single displacement reaction.
4. Double displacement reaction.
5. Combustion reaction.

1. Decomposition Reactions:

A reaction in which a chemical substance breaks down to form two or more simpler substances is called a decomposition reaction. These reactions require some energy for decomposition.

For example: Calcium carbonate decomposes into calcium oxide and carbondioxide in presence of heat.

$$CaCO_3(s) \xrightarrow{Heat} CaO(s) + CO_2(g)$$

Similarly, potassium chlorate ($KClO_3$) on heating produces two simpler substances, potassium chloride and oxygen gas.

$$2KClO_3(s) \xrightarrow{Heat} 2KCl(s)+3O_2(g)$$

(2) Addition or Combination Reaction:

A reaction in which two or more substances combine to form a single substance is called an addition or combination reaction. These reactions are reverse of decomposition reactions.

For example: Calcium oxide (CaO) reacts with carbon dioxide (CO_2) to form calcium carbonate ($CaCO_3$).

$$CaO(s)+CO_2(g) \longrightarrow CaCO_3(s)$$

Another example is, when sodium reacts with chlorine gas it gives a new substance called common salt (NaCl).

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

(3) Single Replacement (Displacement) Reaction:

A reaction in which one atom or group of atoms of a compound is replaced by another atom or group of atoms is defined as displacement reaction. Some metals react with acids, bases or even water to displace hydrogen (H_2) gas.

For example: Zinc replaces hydrogen in hydrogen chloride (HCl) to give zinc chloride.

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2(g)$$

Similarly sodium (Na) metal reacts with water to give sodium hydroxide and hydrogen gas.

$$2Na + 2(H - OH) \longrightarrow 2NaOH + H_2(g)$$

When chlorine reacts with a solution of potassium bromide, chlorine replaces bromine to form KCl and Br_2 vapours.

$$Cl_2(g) + 2K\,Br(aq) \longrightarrow 2KCl(aq) + Br_2(g)$$

(4) Double Displacement Reaction:

It is a reaction in which two compounds exchange their partners, so that two new compounds are formed. In double displacement reaction usually there is an exchange of ionic radicals.

For example: When sodium chloride (NaCl) is reacted with silver nitrate ($AgNO_3$) solution, they exchange their partners to form two different compounds silver chloride (AgCl) and sodium nitrate ($NaNO_3$).

$$NaCl(aq) + AgNO_3(aq) \longrightarrow NaNO_3(aq) + AgCl(s) \text{ (insoluble white precipitate)}$$

Consider another example, when, calcium chloride ($CaCl_2$) is reacted with sodium carbonate (Na_2CO_3) they exchange their partners to form two new compounds, sodium chloride and calcium carbonate ($CaCO_3$)(s).

$$CaCl_2(aq) + Na_2CO_3(aq) \longrightarrow 2NaCl(aq) + CaCO_3(s) \text{ (insoluble white precipitate)}$$

Remember that neutrilization and hydrolysis reactions are also double displacement reactions. These reactions will be discussed in the 9th chapter on Acids, Bases and Salts.

(5) Combustion Reaction:

A reaction in which substances react with either free oxygen or oxygen of the air, with the rapid release of heat and flame, is called combustion reaction.

For example, when methane (CH_4), gas burns in air, it forms carbon dioxide gas (CO_2), water (H_2O) and heat.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g) + \Delta H \text{ (heat)}$$

Similarly, when carbon (C) burns in air, it produces carbon dioxide (CO_2) gas and heat

$$C(s) + O_2(g) \longrightarrow CO_2(g) + \Delta H \text{ (Heat)}$$

2.4.2 Chemical Equation:

Chemical equation is short hand method of describing (expressing) the chemical reaction, in terms of symbols and formulae of the substances involved in a chemical reaction.

The starting substances are called reactants and are always written on the left hand side, where as the substances which are produced (formed) are known as products and are always written on the right hand side of an equation. The reactants and products are separated from one another by using the single arrow ($\rightarrow$) or double headed arrow ($\rightleftharpoons$), depending upon the kind (type) of reaction.

For illustration, when two molecules of hydrogen (H_2) combine (react) with one molecule of oxygen (O_2) to give two molecules of water (H_2O), instead of writing the full names of substances, chemist represents this chemical reaction in the form of following equation.

eg. $2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$

Here, hydrogen and oxygen are called "Reactants" (substances, which are present before the chemical reaction) and water is the product (substances that result from the chemical reaction).

The numbers, in front of the formulas in a chemical equation are called co-efficients (they show the number of molecules that react with each other) where no co-efficient appears, only one number is considered.

one coefficient

$$2H_2 + O_2 \longrightarrow 2H_2O$$

coefficients coefficients

The expressions (g), (l) and (s) are placed some times as subscript after the formulas of the reactants and products, indicating the state, gaseous, liquid and solid of reactants and products.

eg. $C(s) + O_2(g) \longrightarrow CO_2(g)$

Another expression, frequently used is (aq) for aqueous, showing that the substance is in the form of water solution.

eg. $NaCl(s) \xrightarrow{H_2O} NaCl(aq)$

Sometimes, reaction conditions are written over the arrow, for example, when reactants are heated, a capital Greek letter delta (Δ) may be placed over the arrow.

eg. $CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$

Similarly, if catalyst is used, this catalyst is shown over the arrow.

eg. $2SO_2(g)\ O_2(g) \xrightarrow{Pt} + 2SO_3(g)$

If the reactants are heated in the presence of catalyst, both the symbols are placed on the arrow.

eg. $2KClO_3(s) \xrightarrow{\Delta,\ MnO_2} 2KCl(s) + 3O_2(g)$

2.4.3 Writing of Chemical Equation:

Before we write a chemical equation we must know the composition of all reactants and products. Otherwise, we cannot introduce the proper formulae into equation. Knowing the formulae of all reactants and products, we balance the equation by knowing the fact that in a chemical reaction atoms are neither created nor destroyed (to conform with the law of conservation of mass). Thus the number of atoms of each element must be same on both sides of the arrow in a chemical equation.

Consider, when zinc (Zn) reacts with sulphuric acid (H_2SO_4) to form zinc sulphate $ZnSO_4$ and hydrogen (H_2) gas.

The chemical equation for this reaction can be represented as:

$$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$$

Where the arrow ($\rightarrow$) is read as "gives" "produces", "yields" or "forms". The (+) sign on the left side of equation appears for "reacts with" and on the right side of an equation is read as "and". The reaction is assumed to proceed from left to right, as the arrow indicates.

2.4.4 The meaning of Chemical Equation:

The chemical equation gives the following important information about the chemical reactions.

i) The nature of reactants and products.
ii) The relative number of each i.e reactants and products.

2.4.5 Balancing of Chemical Equation:

All chemical equations must be balanced in order to comply with the law of conservation of mass. In balancing the chemical equation, to make the number of atoms of each element the same on both sides of the equation, we can change the co-efficient (the number in front the formulas) but not subscript (the number within formulas). Remember that changing a subscript in a formula e.g. from H_2O to H_2O_2 change the identity of chemical compound. The substance hydrogen peroxide (H_2O_2) is quite different from water (H_2O). In contrast, changing a co-efficient, only changes the amount but not the identity of the substance. $2H_2O$ means two water molecules and $3H_2O$ means three water molecules and so on. Reduce the co-efficient to their smallest whole number values, if necessary, dividing them by common divisor. Thus the formulas of the reactants and the products in an equation remain same i.e. fixed and cannot be altered; so the only way of balancing an equation by taking appropriate numbers of molecules of the reactants and the products concerned.

Most chemical equations can be balanced by inspection method, that is, by trial and error method, with the experience, you should be able to balance any equation quickly. In general we can balance the equation by the following steps.

1. Write the correct formulae of all reactants on the left side and the formulae of products on the right side of an equation.
2. Balance the number of atoms on each side.
3. If the number of atoms may appear more on one side than the other, balance the equation by inspection method for this, multiply the formula by co-efficient so as to make the number of atoms, same on both sides of an equation.
4. The covalent molecules of hydrogen, oxygen, nitrogen and chlorine exist as di-atomic molecules eg. H_2, O_2, N_2 and Cl_2, rather than isolated atoms, hence we must write them as such in chemical equation.
5. Finally, check the balanced equation, to be sure that the number and kind of atoms are the same on both sides of equation.

Let us consider a specific example. In the laboratory, oxygen (O_2) gas is conveniently prepared by heating potassium chlorate ($KClO_3$). The products are potassium chloride (KCl) and oxygen (O_2) gas.

To balance the equation.

Write correct formulae of reactants on left side and formulae of the products on right side of an equation.

$$KClO_3(s) \longrightarrow KCl(s) + O_2(g)$$

Balance the number of atoms on each side.

Reactants $\longrightarrow$	Products
K (1)	K (1)
Cl (1)	Cl (1)
O (3)	O (2)

We see that (K) and (Cl) elements have the same number of atoms on both sides of the equation, but there are three oxygen atoms on the left and two oxygen atoms on the right side of the equation. We can balance the oxygen atoms by placing a co-efficient 2 in front of $KClO_3$ and 3 in front of O_2.

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

Reactants $\longrightarrow$	Products
K (2)	K (1)
Cl (2)	Cl (1)
O (6)	O (6)

We balance the (K) and (Cl) atom by placing 2 in front of KCl.

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

Reactants $\longrightarrow$	Products
K (2)	K(2)
Cl (2)	Cl (2)
O (6)	O (6)

This equation is now balanced, because the number of atoms of each element are same on both sides of the equation.

Note that this equation could also be balanced with co-efficients, that are multiple of 2 of each, for example:

$$4KClO_3(s) \longrightarrow 4KCl(s) + 6O_2(g)$$

However, it is common practice to reduce the co-efficients to their smallest whole number value, here, this equation is divided by 2, so as to get smaller whole number ratio.

$$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$$

Consider, another example, when hydrogen burns in air (which contains oxygen) to form water, we can write the equation.

$$H_2(g) + O_2(g) \longrightarrow H_2O(l)$$

Balance the number of atoms on each side.

Reactants $\longrightarrow$	Products
H (2)	H(2)
O (2)	O(1)

We see that hydrogen element has the same number of atoms on both sides of the equation, but there are two oxygen atoms on left side and one oxygen atom on the right side of the equation. We can balance the oxygen atoms by placing co-efficient 2, in front of water (H_2O).

$H_2 + O_2$	$\longrightarrow$	$2H_2O$
Reactants	$\longrightarrow$	Products
H (2)		H(4)
O (2)		O(2)

Now we see that the number of oxygen atoms is same on both sides, but the number of hydrogen atoms can be balanced by placing 2 in front of (H_2) on left side of the equation, we get:

$2H_2(g) + O_2(g)$	$\longrightarrow$	$2H_2O(l)$
Reactants	$\longrightarrow$	Products
H (4)		H(4)
O (2)		O(2)

The equation is now balanced, because the number of atoms of each element are same on both sides of the equation.

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$

For further illustration, let us consider the combustion of natural gas methane (CH_4) in oxygen or air, which yields carbon dioxide (CO_2) and water (H_2O). We write from this information.

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

Balance the no. of atoms on each side

Reactants	$\longrightarrow$	Products
C (1)		C(1)
H (4)		H(2)
O (2)		O(3)

We see that the number of C-atoms is same on both sides but the number of hydrogen atoms on left side is (4), where as on right side is (2), we can balance it by placing a co-efficient 2 in front of water on right side, thus we get.

$CH_4 + O_2$	$\longrightarrow$	$CO_2 + 2H_2O$
Reactants	$\longrightarrow$	Products
C (1)		C(1)
H (4)		H(4)
O (2)		O(4)

Now, there is a difference in oxygen atoms, there are two oxygen atoms on left side but four oxygen atoms on right side, this can be balanced by placing co-efficient 2 in front of oxygen (O_2) on left side,

Thus we get

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

Reactants	$\longrightarrow$	Products
C (1)		C(1)
H (4)		H(4)
O (4)		O(4)

The equation is now balanced, because the number of atoms of each element are same on both sides of equation

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

2.4.6 Concept of Mole-Ratio Based on Balanced Chemical Equation:

As we have seen, the co-efficient in a chemical reaction represents the number of moles (molecules) and not masses of molecules. However in chemical reaction, the amount of the reactants needed can not be determined by counting molecules directly. Counting is always done by weighing. To find out the masses of reactants and products, in a balanced equation, the first step is to find out the mole-ratios in the balanced equation. Then convert the moles of reactants or products into mass.

The following steps will help in calculating the amount of reactant or products.

(1) Balance the equation for the given chemical reaction.

(2) Use the co-efficients in the balanced equation to get the mole ratio.

(3) Use the mole-ratio to calculate the number of moles of desired reactants or products.

(4) Convert the moles of reactants or products into mass, if required by the problem.

The following example illustrates the use of the above four steps, in solving the mole-ratio problems.

Calculating The Amount of Reactants.

Hydrogen (H_2) reacts with oxygen (O_2) to form water. The equation for this reaction is:

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$

i) How many moles of oxygen are needed to react with 4.5 moles of hydrogen?

ii) How many grams of hydrogen will completely react with 100g of oxygen to form water?

Answer.

Step-(1) Balance the equation for chemical reaction.

$$2H_2(g) + (O_2)(g) \longrightarrow 2H_2O(l)$$

$$2(2)g \;\; 2(16)g \longrightarrow 2(2+16g)$$

$$4g + 32g \longrightarrow 36g$$

Step-(2) Set up the mole ratio, since 2 moles of hydrogen react with 1 mole of oxygen to produce 2 moles of water.

$$2H_2 = O_2 = 2H_2O$$

$$2 \text{ moles} = 1\text{mole} = 2 \text{ moles}$$

Step-(3) (i). Calculate the moles of oxygen as follows.

The balanced eq. shows that:

2 moles of hydrogen require 1 mole of oxygen

1 mole of hydrogen requires $\frac{1}{2}$ mole of oxygen

4.5 moles of hydrogen require $\frac{1\times4.5}{2} = 2.25$ moles

$= 2.25$ moles of oxygen

ii) 32g of oxygen react with 4g of hydrogen

1g of oxygen reacts with $\frac{4g}{32}$

100g of oxygen react with $\frac{4\times100}{32} = 12.5g$

Result: i) No. of moles of oxygen = 2.25 moles.
ii) 100 g of oxygen require = 12.5g of hydrogen.

Calculating the Amount of Products:

Methane (CH_4) is a common fuel used for cooking. The chemical reaction is,

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

i) How many moles of CO_2 will be produced by complete combustion of 10 moles of CH_4.

ii) What mass of H_2O vapours will be formed?

Answer.

Step (1) Balance the equation for chemical reaction

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

Step (2) Set up the mole ratio, the balance equation shows that 1 mole of methane (CH_4) reacts with 2 moles of oxygen (O_2) to produce 1 mole of carbon dioxide (CO_2) and 2 moles of water (H_2O)

$$\underset{1\text{mole}}{CH_4} + \underset{2\text{moles}}{2O_2} \longrightarrow \underset{1\text{mole}}{CO_2} + \underset{2\text{moles}}{2H_2O}$$

Step (3)

i) Calculate the moles of (CO_2) as follows.
The balanced equation shows that:
1 mole of CH_4 on combustion produces 1 mole of CO_2
10 moles of CH_4 on combustion produces 1x10 = 10 moles of CO_2

ii) Mass of H_2O
The balanced equation indicates that:
1 mole of CH_4 on combustion produce 2 moles of H_2O
10 moles of CH_4 on combustion produces 2x10=20 moles of H_2O
Mass of H_2O = no. of moles x m. mass
Mass of H_2O = ??
No. of moles = 20 moles
M. mass of H_2O = 2+16 = 18g
∴ Mass of H_2O = 20 x 18g = 360g

Result. No. of moles of CO_2 = 10 moles
Mass of H_2O formed = 360 grams

SUMMARY

1. There are several laws that govern the composition of matter and chemical reaction. These are, the law of conservation of mass, the law of definite proportions, the law of multiple proportions and the law of reciprocal proportions.
2. Atomic mass is the average mass of naturally occurring isotopes, which is compared to the mass of one atom of carbon-12 a.m.u.
3. The molecular mass of a molecule is the sum of the atomic masses of all the atoms present in that molecule.
4. Formula mass of the substance is sum of the atomic masses of all the atoms present in the formula unit of the substance.
5. Molar mass is the mass in grams of one mole of a substance.
6. Mole is the amount of substance that contains as many elementary particles (atoms, ions, molecules or formula units) as there are atoms in 12g of carbon-12 isotope.
7. Avogadro's number, the number of particles i.e. atoms, ions, molecules or formula units in one mole of substance equal to $6.02 x 10^{23}$.
8. The empirical formula is the formula which describes the smallest or the least ratio of the combining atoms of different elements present in a molecule.
9. The Molecular formula shows the actual number of the combining atoms of all the elements present in a molecule and is always an integer multiple of empirical formula.
10. Chemical reaction is a change in which the composition of substance is changed and new substances are formed. Chemical reactions can be divided into five different types. Decomposition reaction, combination reaction, single replacement reaction, double replacement reactions and the combustion reactions.
11. A chemical equation represents the chemical change or reaction, showing reactants on the left side of an arrow and products on the right. The equation is balanced by placing co-efficients in front of the formulas, so the number of atoms of each element are same on each side of an arrow.
12. Amounts of reactants and products formed can be calculated from the balanced equation for a reaction by using the mole-ratios relating the reactants and products.

EXERCISE

1. Fill in the blanks :

(i) 18 grams of H_2O contains .... molecules.

(ii) A change which alters the composition of a substance is called.......

(iii) A reaction in which a chemical substance breaks down to form two or more simpler substances, is called................

(iv) The reaction of NaCI with $AgNO_3$ is given as

$NaCI(aq) + AgNO_3(aq) \longrightarrow AgCI(S) + NaNO_3(aq)$

is the reaction of the type....................

(v) When metals react with acids or water then produce .... gas.

(vi) .... is the reaction in which two or more substances combine together to form a single substance.

(vii) A reaction in which a substance burns in oxygen to produce heat and flame is called.........

(viii) .... is the short hand method to describing a chemical reaction.

(ix) The reaciton $Zn + 2HCI \longrightarrow ZnCl_2 + H_2$ (g) is the .... replacement reaction.

2. Tick the correct answer :

(i) Mass is neither created nor destroyed during a chemical change, is the statement of:

a) Law of conservation of mass.
b) Law of definite proportions.
c) Law of multiple proportions.
d) Law of reciprocal proportions.

(ii) A given compound always contains exactly the same proportion of elements, by mass, is the statement of:

a) Law of conservation of mass.
b) Law of definite proportions.
c) Law of multiple proportions.
d) Law of reciprocal proportions.

(iii) The average mass of natural mixture of isotopes, which is compared to the mass of one atom of C-12 a.m.u, is called:

a) Atomic number. b) Mass number.
c) Atomic mass. d) None of these.

(iv) A formula that gives only the relative number of each type of atoms in a molecule, is called:
a) Empirical formula. b) Molecular formula.
c) Molecular mass. d) Formula mass.

(v) A formula that indicates actual number and type of atoms in a molecule, is called:
a) Empirical formula. b) Molecular formula.
c) Molecular mass. d) Formula mass.

(vi) The sum of the atomic masses of all atoms in a molecule is called:
a) Empirical formula. b) Molecular formula.
c) Molecular mass. d) Formula mass.

(vii) The sum of the atomic masses of all atoms in a formula unit of substance is called:
a) Empirical formula. b) Molecular formula.
c) Molecular mass. d) Formula mass.

(viii) The mass of (1) mole of substance expressed in grams, is called:
a) Empirical formula. b) Molecular formula.
c) Molecular mass. d) Molar mass.

(ix) 44 a.m.u. of CO_2 is equal to:
a) Molar mass. b) Atomic mass.
c) Molecular mass. d) Mass number.

(x) 5 moles of H_2O are equal to:
a) 80g. b) 90g.
c) 100g. d) 90 a.m.u.

3. Write answer of the following questions :

(i) State the law of conservation of mass? Describe Landolt experiment?
(ii) State the law of definite proportions in your own words.
(iii) What is law of multiple proportions? Explain with examples.
(iv) State the law of reciprocal proportion and illustrate it with examples.
(v) What is empirical formula? Give an example?
(vi) What is molecular formula? Give an example?

(vii) Can one substance have the same empirical formula and molecular formula? Explain with examples.

(viii) What is the difference between empirical formula and molecular formula?

(ix) What is atomic mass unit?

(x) The value of atomic mass of carbon in periodic table is 12.011 a.m.u, rather than 12.00 a.m.u? Explain?

(xi) The atomic masses of ${}_{17}Cl^{35}$ (75%) and ${}_{17}Cl^{37}$ (25%). Calculate the average atomic mass of chlorine.

(xii) How many atoms are there in 5 moles of sulphur?

(xiii) What is a mole? What is the molar mass of substance? Find out the molar mass of SO_2?

(xiv) What is Avogadro's number? What number of oxygen atoms are present in 4g of oxygen?

(xv) What does Avogadro's number represent?

(xvi) What is mass in grams of a single atom of each of the following elements?

(a) Carbon(C) (b) Magnesium (Mg) (c) Calcium (Ca).

(xvii) What is mass in grams of $1x10^{20}$ atoms of Na?

(xviii) Define the following.

(a) Molecular formula mass

(b) Formula mass

(c) Molar mass

(xix) Calculate the molecular mass (in a.m.u) of each of the following substances.

(a) H_2O (b) H_2O_2 (c) C_6H_6 (d) C_2H_6O.

(xx) Calculate the formula mass (in a.m.u.) of each of the following:

(a) KNO_3 (b) Al_2O_3 (c) $CaCO_3$ (d) $MgCl_2$.

(xxi) Calculate the molar mass of the following substances.

(a) S_8 (b) CS_2 (c) $CHCl_3$ (Chloroform).

(d) CH_3 –COOH (Acetic acid).

(xxii) The formula for rust is Fe_2O_3. How many moles of (Fe) are present in 30g of rust?

(xxiii) Define the following terms:

(a) Chemical reaction.

(b) Reactants.

(c) Products.

(xxiv) What is chemical equation? What is a co-efficient? Give an example of balanced equation.

(xxv) What is combination reaction? Give an example.

(xxvi) What is decomposition reaction? Will two or more elements always be the products of this type of reaction? Explain with examples.

(xxvii) What is single replacement reaction? Give an example?

(xxviii) Explain double displacement reaction with examples?

(xxix) Balance the following equations by inspection method?

(a) $C + O_2 \longrightarrow CO$

(b) $CO + O_2 \longrightarrow CO_2$

(c) $KNO_3 \longrightarrow KNO_2 + O_2$

(d) $NaH\,CO_3 \longrightarrow Na_2CO_3 + H_2O + CO_2$

(e) $CaCO_3 + HCl \longrightarrow CaCl_2 + H_2O + CO_2$

(f) $NH_3 + O_2 \longrightarrow NO + H_2O$

(xxx) Which of the following reaction is either a decomposition reaction or combination reaction?

(a) $MgCO_3 \longrightarrow MgO + CO_2$

(b) $C_2H_4 + H_2 \longrightarrow C_2H_6$

(c) $BaCO_3 \longrightarrow BaO + CO_2$

(d) $N_2 + 3H_2 \longrightarrow 2NH_3$

(xxxi) Balance the equation and decide which one is single replacement reaction.

(a) $C_2\,H_2 + H_2 \longrightarrow C_4H_8$

(b) $Ca + H_2O \longrightarrow Ca\,(OH)_2 + H_2$

(c) $C_2H_5\text{-}OH + Na \longrightarrow C_2H_5ONa + H_2$

(xxxii) Balance the following equation, which of them is a decomposition reaction or combination reaction.

(a) $CH_4 + O_2 \longrightarrow CO_2 + H_2O$

(b) $NO_2 \longrightarrow NO + O_2$

(c) $Na + O_2 \longrightarrow Na_2O$

(xxxiii) Consider the combination of (CO) with oxygen (O_2) gas.

$$2CO + O_2 \longrightarrow 2CO_2$$

Calculate the number of moles of (CO_2) produced when 50 moles of oxygen are reacted with all of (CO)?

(xxxiv) Calcium carbonate ($CaCO_3$) on heating gives calcium oxide (CaO) and CO_2 gas.

$$Ca\,CO_3 \longrightarrow CaO + CO_2$$

Calculate how many grams of calcium oxide (CaO) can be obtained by heating 8 moles of $CaCO_3$?

(xxxv) Silicontetrachloride ($SiCl_4$) can be prepared by heating (Si) in chlorine gas (Cl_2).

$$Si + 2Cl_2 \longrightarrow SiCl_4$$

If we want to prepare 10 moles of ($SiCl_4$) how many moles of molecular chlorine (Cl_2) will be used in the reaction.

(xxxvi) Fermentation is chemical decomposition, in which glucose ($C_6H_{12}O_6$) is converted into ethyl alcohol (C_2H_5-OH) and carbon dioxide (CO_2).

$$(C_6H_{12}O_6) \xrightarrow{\text{Fermentation}} 2C_2H_5\text{-}OH + 2CO_2$$

What will be the amount of ethyl alcohol in grams and moles, which can be obtained by fermentation of 5000g of glucose.