Unit 1

BASIC CONCEPTS OF CHEMISTRY

1.1 Introduction

Chemistry is the science of molecules, variety of compounds and transfomations.

Chemistry has played a central role in preventing and finding solution of problems of science. Chemistry is associated with the substances formed in environment and the changes taking place in it.

According to Roald Hoffman chemistry is not only the science of 100 elements but diversity of infinite varieties of molecules which are prepared from them. This purpose can be understood and described by the fundamental particles that are atoms and molecules of matter. Is it possible to observe these particles or can they be compared ? Can we count the number of atoms or molecules in a definite quantity of matter ? Can a quantitative relationship be established by the number of atoms or molecules in a definite quantity of matter ? We shall get the answers of many such questions with the help of basic concepts of chemistry discussed in this unit.

Moreover we will also study how to present physical quantity in its numeric form.

1.2 Importance of Study of Chemistry

Science is a bridge to understand the nature. In other words, science is a continuous human effort to understand nature using systematic knowledge.

The study of chemistry is so wide and diversified, that it is classified in various disciplines such as inorganic, organic, physical, analytical industrial, biochemistry etc.

Nowadays, study of chemistry becomes very easy due to principle and pro-principle, which are derived from different facts based on their effective interlinking.

Chemistry is the science of occurrence and synthetic composition, structure and properties of substance. The branches of science like chemistry, physics, biology, geology are interrelated to one another. Chemistry plays an important role in our everyday life. We experience various chemical phenomena from the time we get up in the morning till we go to bed at night. The principles of chemistry are useful in diverse areas, such as functioning of brain, operation of computer, weather patterns, digestion of food etc. Chemical industries manufacturing fertilizers, alkalies, acids, soaps, detergents, salts, polymers, alloys, dyes, drugs and other inorganic and organic chemicals, including new materials contribute in a big way to the national economy.

Chemistry has increased the comforts in the human life. As we know production of fertilizers on large scale and production of pesticides and insecticides are carried out through chemistry. A large group of semisynthetic and synthetic compounds are included in chemistry. Also chemistry has advanced a lot in pharmaceutical field. The discovery of medicines used in cancer and the life saving drugs like 'cisplatin' and 'texol' owe to chemistry.

Like the two sides of a coin, on one side with many advantages due to the development of chemistry, on the other hand there is a possibility of tremendous damage to the human life and environment. In the present circumstances some of the efforts made in chemistry have successfully resulted in solving (abating) the global problems like pollution.

Chlorofluorocarbon (CFC) used in refrigerator and airconditioner are hazardous to ozone layer and the environment. The safer alternatives of CFC have been invented and can be successfully prepared synthetically. Nowadays in refrigerators, instead of CFC, the less hazardous substance to environment known as HFC-134a (1,1,1,2,-tetra fluoroethane) is used.

The chemistry involved in production of chemicals without harming the environment is known as Green chemistry.

However, certain challenges like understanding of biochemical reaction, use of enzymes for large scale productions of chemicals, combination of specific types of materials, sources of energy are there before the future generation of chemists. The country needs reputed and creative chemists for the counterattack of such challanges.

1.3 Nature of Matter

You have studied about matter in the lower standards. "Anything that has mass and occupies space is matter". Matter is made up of particles.

Based on their physical state and chemical characteristics matter is classified as under :

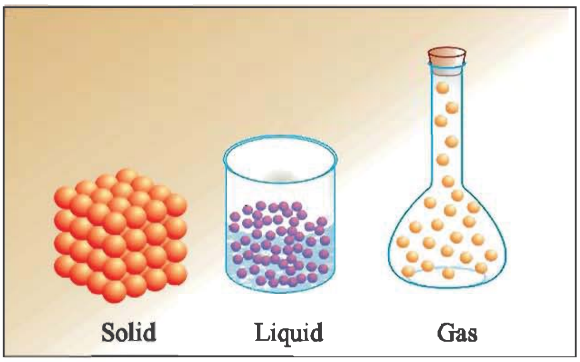
Matter

- Based on physical state of matter
  - Solid
  - Liquid
  - Gas
- Based on chemical characteristics
  - Element
  - Compound
  - Mixture
    - Homogeneous
    - Heterogeneous

1.3.1 Classification based on physical states :

Natural and man-made, everything around us is made up of matter i.e., pens, clocks, spectacles, air, water, animals are made of matter. They have mass and they occupy space (volume). Matter can exit in three physical states : (1) solid (2) liquid and (3) gas.

These three states can be represented as shown in fig. 1.1.



Arrangement of particles in solid, liquid and gaseous state

Figure 1.1

(1) Solid State : Solids have definite volume and definite shape. In solid, the particles are held very close to each other and the intermolecular attractive forces in their component particles keep them together tightly. They retain their specific shape. e.g. ice, iron etc.

(2) Liquid State : Liquids have definite volume but no definite shape. They take shape of the container in which they are poured. In liquid state distances between component particles is more than that in a solid. In liquid state the intermolecular attractive forces between component particles is sufficient enough to keep them together but not enough to keep their places fixed. So liquids are fluids. e.g. water, bromine, benzene etc.

(3) Gaseous State : Gases have neither definite volume nor definite shape. They completely occupy the whole space of the container in which they are filled and spread into the entire container. e.g. air, argon etc.

These three states of matter are interconvertible by changing the conditions of temperature and pressure, i.e. they change their state.

$$\text{Solid (Ice)} \underset{\text{Decreasing temperature}}{\overset{\text{Increasing temperature}}{\rightleftarrows}} \text{Liquid (Water)} \underset{\text{Decreasing temperature}}{\overset{\text{Increasing temperature}}{\rightleftarrows}} \text{Gas (Vapour)}$$

1.3.2 Classification based on chemical structure :

The matter is classified into three categories, element, compound and mixture, on the basis of their chemical structure.

(1) Element : Lavoisier (1743-94) gave the explantion for element. According to him, element consists of only one type of atoms. In different elements, atoms are different. The atoms of one element are different from atoms of the other element. Thus, every element has its own independent property which is not the same in other elements, e.g carbon, sodium, oxygen etc.

(2) Compound : When two or more elements combine chemically with one another, a compound is formed. When the compounds are formed, the elements present in them show new type of properties by losing their own individual properties. For example, hydrogen (H) and oxygen (O) are gases, water (H_2O) formed by their combination is a liquid. Here, hydrogen burns explosively in air and oxygen is a supporter of combustion; but water is used as a fire extinguisher. Thus, hydrogen and oxygen change their properties in water.

(3) Mixture : A mixture is a material consisting of two or more kinds of matter, each retaining its own characteristic properties. Mixture can be separated by physical methods. Mixture can be classified in two different categories :

(i) Homogeneous mixture (ii) Heterogeneous mixture

(i) Homogeneous mixture : In the homogeneous mixture the components are in same physical state. Solution is a homogeneous mixture. In a homogeneous mixture, composition is uniform throughout and does not separate from definite borderline, e.g. mixture of sugar and water (sugar solution), salt and water (salt solution), oxygen and nitrogen (air), zinc and copper (alloy brass). The components of sugar solution can be separated by evaporation. Oxygen and nitrogen gases are obtained by liquification of the air-mixture.

(ii) Heterogeneous mixture : In the heterogeneous mixture the components are different in their physical state. In the heterogeneous mixture the componentes do not mix with each other. But can be separated by definite boundary of two different physical states. In heterogeneous mixture the composition is not uniform throughout, e.g sodium chloride (NaCl) mixed with iron (Fe) makes heterogeneous mixture. Here, both are in similar solid state but as different entities in mixture and separate from borderline surface.

The component of both the types of mixture can be separated by using physical methods such as filtration, crystallization, distillation etc. For example, to separate the mixture of salt (NaCl) and iron, it is dissolved in water and filtered iron remains on filter paper and solution of NaCl is obtained. The component salt (NaCl) can be separated by using evaporation and iron by filtration.

1.4 Physical Quantities

The value of physical quantity is always equal to a definite numeric value and a definite unit. It is necessary to describe the physical quantity with the least possible units. For example, mass of a substance is 5.0 kg. Here, just writing 5.0 carries no meaning, but adding the term kg after 5.0 indicates its mass. In order to express measured or calculated quantity, it is essential to have a suitable system of units. Thus, for the system of units some least number of these quantities are selected in such a way that other quantities can be derived from them. The least value of quantities of units selected are known as fundamental units or basic units. Other units of measurements which can be derived from the fundamental units are known as derived units. For example, mass of piece of iron is 5.0 kg. Here kg is a basic unit of mass. By using this fundamental unit, derived units like density, volume etc. can be obtained.

There are various systems of units for physical measurement which were developed at different times. Some of them are as follows :

(i) FPS system (Foot, Pound, Second) in the year 1588.

(ii) CGS system (Centimeter, Gram, Second) in the year (1791-1795).

(iii) MKS system (Meter, Kilogram, Second) in the year (1791-1795). India accepted (ii) and (iii) in 1956.

(iv) SI (Le System, International d' Units) in the year 1971.

To solve the difficulties created by a number of methods, IUPAC (International Union of Pure and Applied Chemistry) and IUPAP (International Union of Pure and Applied Physics) have recommended SI method to be used uniformly.

The scientific world requires a method for international dealings. This method should be uniform, reliable, standardised and universally acceptable and should have uniform units.

The International System of units was defined, in 1960 during the 11th General Conference of International Bureau of Weights and Measures which was held at Sevres near Paris. The organization created a diplomatic treaty for uniformity in units. International system of units was accepted in 1971. In French, it is called "Le Systeme International d' Units". In short, it is called SI system also. In this system seven basic (fundamental) units are defined. These are as under :

Table 1.1 SI units (Basic units)

Physical quantity	Symbol for quantity	Symbol of SI units	Name of SI units
Length	l	m	meter
Mass	m	kg	kilogram
Time	t	s	second
Electric current	I	A	ampere
Thermodynamic temperature	T	K	Kelvin
Amount of substance	n	mole	mole
Luminous intensity	I_v	cd	candela

Table 1.2 Prefixes used in the SI system

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

Definitions of some of the SI units and derived units are mentioned below :

(1) Mass : Amount of matter present in a substance is called Mass. Its SI unit is kilogram.

Kilogram : The mass of cylinder prepared from platinum-irridium (Pt-Ir) alloy kept in International Bureau of Weight and Measures is called 1 kilogram.

The mass of a substance can be determined in the laboratory by using an analytical balance. However, its fractions–gram, milligram, microgram are used in laboratories due to the smaller amounts of chemicals used in chemical reactions.

(2) Volume (Derived unit) : Derived unit of volume by using SI units is $(\text{meter})^3$.

$$\begin{aligned} \text{volume} &= \text{length} \times \text{breadth} \times \text{height} \\ &= \text{meter} \times \text{meter} \times \text{meter} \\ &= (\text{meter})^3 \end{aligned}$$

In chemistry laboratories, smaller volumes are often denoted in cm^3 or dm^3 units. A common unit liter (L), which is not an SI unit, is used for measurement of volume of liquids. When smaller volume is used, the unit is milliliter (mL). In laboratory, volume of liquids can be measured by cylinder, burette, pipette and volumetric flask. These measuring devices are shown in fig. 1.2.

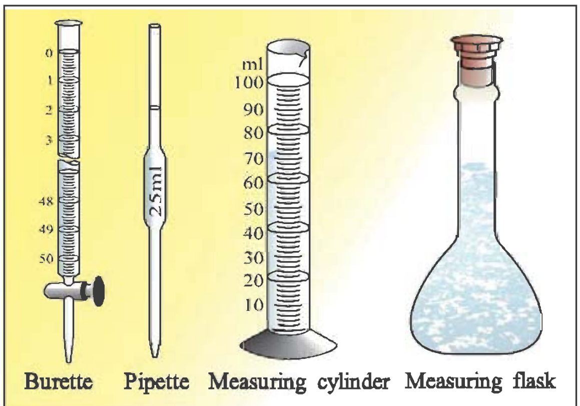


Figure 1.2

1 Liter (L) = 1000 Milliliter (mL)

$= 1000 \text{ cm}^3$

$= 1 \text{ (decimeter)}^3$ OR dm^3

$(\because\ 10 \text{ cm} = 1\text{dm})$

(3) Density : (Derived unit) : Density of a substsnce is the amount of mass per unit volume. Derived unit of density by using SI units is as under :

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{\text{SI unit of mass}}{\text{SI unit of volume}}$$

$$= \frac{\text{kg}}{\text{m}^3}$$

$$= \text{kg m}^{-3}$$

Smaller unit of density is g/cm^3 OR g cm^{-3} where mass is expressed in gram and volume in cm^3.

(4) Temperature (Basic unit) : Measurement of temperature is done in the two units-degree Celsius (^{0}C) and degree Fahrenheit (^{0}F). SI unit of temperature is Kelvin (K). Both the scales can be related to each other by following relationship :

$$^0\text{F} = \frac{9}{5}(^0\text{C}) + 32$$

$$\text{K} = {}^0\text{C} + 273.15$$

but normally expressed as

$$\text{K} = {}^0\text{C} + 273$$

$$0^0\text{C} = 32^0\text{F} = 273 \text{ K}$$

It should be noted that temperatures below 0^0C are possible in Celsius scale but in Kelvin scale negative temperature is not possible.

(5) Length : (Basic unit) : SI unit of length is meter. The meter was redefined in 1983 by CGPM (Conference Generale des Poids at Measures).

Meter : The distance travelled by light in vacuum in time interval of 1/299,792,458 second is called 1 meter.

The meter was originally defined as the length between two marks on a Pt-Ir bar kept at a temperature 0^0C as a standard reference.

Dear students, question arises in your mind that why Pt-Ir is chosen ? Because it is highly resistant to chemical attack and its mass or length does not change for a long time.

1.5 Laws of Chemical Combination

When two or more substances react chemically a compound is formed.

The following are the laws which govern the formation of chemical compounds :

1.5.1 Law of Conservation of Mass : "Matter can neither be created nor destroyed." This law was put forth by Antoine Lavoisier in 1789. He performed carefully experimental studies for combustion reactions for reaching to the above conclusion. This law formed the basis for several later developments in chemistry. In fact this was the result of planned experiments performed by Lavoisier.

1.5.2 Dalton's Atomic theory : We know that Dalton's atomic theory can explain the laws of chemical combination. As a result of several experimental studies, John Dalton (1776-1884) was led to a law "Matter is made up of small indivisible particles." These smallest indivisible particles are called atoms. Dalton's atomic theory is helpful to know chemical reaction but during his time the fundamental particles of atom like proton, electron, neutron were not discovered.

In 1808, Dalton published *A New System of Chemical Philosophy* in which he proposed law of chemical combination which provides explanation for existence of atom. Dalton's assumptions are as follows:

(1) All elements are made up of smallest particles called atoms. Atom is indivisible.

(2) All the atoms of a given element are identical but are different from those of other elements.

(3) All the atoms of a given element have identical properties including identical mass. Atoms of different elements differ in mass.

(4) Compounds are formed when atoms of different elements combine in a fixed ratio. Atoms present in a compound have a definite composition.

(5) All the atoms of a given element are identical in mass.

(6) Compounds are formed by the combination of atoms of two or more elements in the ratio of small whole numbers. The smallest particles formed by chemical combination of two or more atoms are called molecules.

(7) Chemical reactions involve rearrangement of atoms. These are neither created nor destroyed during chemical reaction.

There is no importance of Dalton's law with reference to nuclear reactions and discovery of isotopes because Dalton's view that atom is indivisible was found wrong. Atom is composed of fundamental particles like proton, electron and neutron.

1.5.3 Law of Constant Proportion : "A given compound always contains exactly the same proportion of elements by weight".

Joseph Proust (1754-1826) observed that samples of cupric carbonate obtained naturally and prepared synthetically in the laboratory had the same percentage composition of elements.

Cupric carbonate $CuCO_3$	% of Cu	% of C	% of O
Natural sample	51.35	9.74	38.91
Synthetic sample	51.35	9.74	38.91

Water (H_2O) produced during the reaction of hydrogen gas with oxygen gas and water obtained by decomposition of hydrogen peroxide (H_2O_2) has same composition. A given compound of water always contains same proportion. Water (H_2O) always contains 2.016 g of hydrogen and 16.0 g of oxygen.

The understanding of law of constant proportion is provided by the Dalton's atomic theory. According to this theory, molecules are made up of atoms and the elements for a given compound are same in each of their samples. As the mass of each atom of a given element is same, the mass of its each molecule will always be the same. i.e

$$H_{2(g)} + \frac{1}{2}O_{2(g)} = H_2O_{(l)}$$

Here simple ratio of volumes of elements H:O in water is 2:1 and ratio 1:8 of the masses for H, (2 g) and O (16 g) which gives fixed mass of H_2O (18 g)

1.5.4 Law of Multiple Proportion : This law was proposed by Dalton in 1803. If two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole number.

For example, hydrogen combines with oxygen to form two compounds,water and hydrogen peroxide. The masses of oxygen 16.0 g and 32.0g combine with fixed mass of hydrogen 2.0 g in H_2O and H_2O_2 respectively.

H_2	+	$\frac{1}{2}O_2$	=	H_2O
hydrogen		oxygen		water
2.0 g		16.0 g		18.0 g
H_2	+	O_2	=	H_2O_2
hydrogen		oxygen		hydrogen peroxide
2.0 g		32.0 g		34.0 g

Here, hydrogen and oxygen combine to form two different compounds. The amount of oxygen that combines with 2.0g of hydrogen in two compounds can be expressed as per the law of multiple proportions as 16:32 bearing simple ratio 1:2.

1.5.5 Law of Combining Weights : According to this law, masses of two elements which separately react chemically with identical masses of a third element are also the masses which react with each other and are in simple multiples of combining weight of an element which is either equal to its atomic weight or simple multiple of it. For example,

in HCl,	H	+	C1
	1 g		35.5 g
in Cl_2O,	O	+	2C1
	16 g		2 × 35.5
			= 71 g

Thus, combined masses of Cl (35.5g) in Cl_2O react with 8.0 gram of oxygen. According to the law of combining weights, 35.5 gram of chlorine reacts with 1.008 g and 8.0 g of hydrogen and oxygen respectively in both the compounds. Hydrogen and oxygen combine with each other to form water (H_2O)where the mass ratio is 2:16 = 1:8.

1.6 Atomic Mass, Molecular Mass, Molar Mass and Mole Concept

(1) Atomic mass : Dalton proposed "Atoms of every element have their own definite charateristic mass." It is called atomic mass.

Every element has its own characteristic atomic mass. It is very difficult to measure the mass of an extremely small atom.With the help of a modern instrument like mass spectrometer the definite value of atomic mass of an atom can be measured.

However, the present system of atomic masses is based on carbon-12 as the standard and has been agreed upon in 1961 by IUPAC and IUPAP. Here C–12 is one of the isotopes of carbon and can be represented as ^{12}C. In this system ^{12}C is assigned a mass of exactly 12 atomic mass unit (amu) and masses of all other atoms are given relatively to this standard. Mass of an atom of hydrogen is defined as a mass equal to one twelth of the mass of one ^{12}C atom.

Nowadays amu is expressed as 'u' as a unit for unified mass.

1 amu = 1.66056×10^{-24}g

Mass of one hydrogen atom = 1.6736×10^{-24}g

$$\text{Atomic mass of hydrogen} = \frac{1.6736 \times 10^{-24}\text{g}}{1.66056 \times 10^{-24}\text{g}}$$

$$= 1.0078 \text{ amu}$$

$$\approx 1.008 \text{ amu}$$

$$\approx 1.008 \text{ u}$$

Similarly the mass of oxygen = 15.995 amu

= 15.995 u

≈ 16.00 amu ≈ 16.00u

(2) Molecular Mass : Molecular mass can be calculated from the atomic masses of all atoms present in a compound. If we know the molecular formula of any compound, the molecular mass can be found by considering total number of atoms present and adding together their total atomic masses. e.g.,

(i) Calculate molecular mass of water (H_2O) molecule

Molecular mass of H_2O

= 2 (Atomic mass of H) + 1(Atomic mass of O)

= 2(1.008u) + 1 (16.0u)

= 18.016u

(ii) Calculate molecular mass of sucrose ($C_{12}H_{22}O_{11}$) molecule.

Molecular mass of $C_{12}H_{22}O_{11}$

= 12 (Atomic mass of C) + 22 (Atomic mass of H) + 11(Atomic mass of O)

= 12 (12u) + 22 (1u) + 11 (16u)

= 144u + 22u + 176u

= 342u

(3) Molar Mass and Mole Concept :

In SI system, mole is introduced as one of the seven basic quantities for the amount of the substance. The number of atoms or molecules even in a small amount of any substance is very large. Repeated use of such large numbers would be tedious and leads to inevitable errors in result. Thus, a definite quantity of mass unit of simillar magnitude is required in practice. Units like one dozen for 12 items, score for 20 items, gross for 144 items as we use in our daily life, compounds used in laboratory for experiments are mentioned in unit of gram. In this unit there are innumerable atoms. In order to relate the masses of different compounds, scientists have given a unit called 'Mole'. Number of composite particles in 1 mole of molecule, atom or ion = 6.022×10^{23}.

Number of atoms 6.022×10^{23} present in 12 gram of isotope ^{12}C as a standard, is accepted as Mole. This figure is called Avogadro's number (N_A). The mass of a ^{12}C is found to be equal to 1.992648×10^{-23} gram by spectrometer. To really appreciate largeness of the number let us write it with all the zeros without using any powers of ten. 602213670000000000000000.

Hence, so many entities constitute one mole of a particular substance.

We can, therefore, say that

1 mole of hydrogen atoms = 6.022×10^{23} atoms of hydrogen

1 mole of water molecules = 6.022×10^{23} moloecules of water.

1 mole of sodium chloride = 6.022×10^{23} molecules of sodium chloride

Knowing that one mole of carbon weighs 12 gram, the number of atoms in it will be as under :

$$\text{Number of atoms of carbon C} = \frac{\text{mass of 1 mole C}}{\text{mass of 1 atom C}}$$

$$= \frac{12 \text{ gram mol}^{-1}}{1.992648 \times 10^{-23} \text{ gram atom}^{-1}}$$

$$= 6.022 \times 10^{23} \text{ atom mol}^{-1}$$

$$\text{Mole} = \frac{\text{Mass of atom or molecule in gram}}{\text{Atomic mass or molecular mass in gram mole}^{-1}}$$

Molar mass of an element or compound is the mass in gram of one mole (6.022×10^{23} particles) of atoms or molecules contained in it.

Example 1.1 : How many moles of 'C' atoms will be there in 100 gram carbon dioxide (CO_2) ? Calculate the number of carbon atoms. (molecular mass of CO_2 = 44 gram-mole^{-1})

Solution :

$$\text{Moles of } CO_2 = \frac{\text{Weight of } CO_2}{\text{Molecular mass of } CO_2}$$

$$= \frac{100 \text{ gram}}{44 \text{ gram mole}^{-1}}$$

$= 2.27$ mole

1 mole molecules of CO_2 contains 1 mole atoms of C

$\therefore$ Mole of C atoms = 2.27 mole

The number of carbon atoms

= mole of carbon $\times$ Avogadro number

$= 2.27 \times 6.022 \times 10^{23}$

$= 13.669 \times 10^{23}$

1.7 Percentage Composition of Element and Molecular Formula

1.7.1 Percentage Composition of Element :

When compound is formed from two or more elements, the amount of element present in a compound is always proportional to its definite mass. If molecular formula of compound is known, one can calculate the mass percentage of element present in that compound. In contrast to this, if one knows the mass percentage of all elements present in a compound, its molecular formula can be determined.

Mass % of an element

$$= \frac{\text{Mass of the element in the compound} \times 100}{\text{Molar mass of the compound.}}$$

Example 1.2 : Calculate the percentage of both elements present in water (H_2O).

Solution :

Atomic mass of H = 1.0 gram mole^{-1}

Atomic mass of O = 16.0 gram mole^{-1}

Molecular mass of H_2O = 18.0 gram mole^{-1}

$$\text{mass \% of Hydrogen} = \frac{2(1.0\text{g mol}^{-1}) \times 100}{18.0\text{g mol}^{-1}}$$

$= 11.11\%$

$$\text{mass \% of Oxygen} = \frac{(16.0 \text{ g mol}^{-1}) \times 100}{18.0 \text{ g mol}^{-1}}$$

$= 88.89\ \%$

Example 1.3 : Calculate the percentage of each element present in ethanol (C_2H_5OH).

Solution :

Atomic mass of C = 12.0 gram mole^{-1}

Atomic mass of H = 1.0 gram mole^{-1}

Atomic mass of O = 16.0 gram mole^{-1}

Molecular mass of ethanol (C_2H_5OH)

$= 2(12) + 6(1) + 1\ (16)$

$= 24 + 6 + 16$

$= 46$ gram mole^{-1}

$$\text{Mass \% of H} = \frac{6(1.0) \times 100}{46} \approx 13.04\%$$

$$\text{Mass \% of C} = \frac{2(12) \times 100}{46} \approx 52.17\%$$

$$\text{Mass \% of O} = \frac{16 \times 100}{46} \approx 34.78\%$$

1.7.2 Empirical Formula and Molecular Formula : In order to decide the molecular formula of any compound, it is necessary to know the percentage composition of elements in it. From this information simple formula of a compound can be decided. This simple formula indicates the relative proportion of atoms of element. It is a formula showing the relative proportion of component atoms in a compound. Thus, formula which represents the composition of a molecule is an empirical formula. Empirical formula of a compound can be calculated in the following sequence :

(i) Indicate the symbol of elements present in a compound.

(ii) Calculate percentage of elements present.

(iii) Calculate the ratio of percentage of elements and atomic masses of elements which can give ratio of atoms present in element.

(iv) The ratio of each element is to be divided by the ratio of element having lowest simple integral ratio will be obtained.

(v) Calculate formula mass of empirical formula.

(vii) Calculate multiple number (n) from infomation of empirical formula mass and molecular mass.

$$n = \frac{\text{Molecular mass}}{\text{Formula mass of empirical formula.}}$$

(viii) The empirical formula is multiplied with 'n' and molecular formula is obtained.

Molecular formula = n $\times$ Empirical formula

where n = Integral number

Example 1.4 : The percentages of carbon, hydrogen and oxygen in an organic substance are 54.55, 9.06, and 36.39 respectively. Find the empirical formula and molecular formula. (Molecular mass of organic substance = 88 gram mole^{-1})

Elements	Atomic mass	Percentage	Atomic ratio	Ratio of simple whole number	Simple(whole) number
C	12	54.55	$\frac{54.55}{12} = 4.55$	$\frac{4.55}{2.27} = 2.0$	2
H	1	9.06	$\frac{9.06}{1} = 9.06$	$\frac{9.06}{2.27} = 3.99$	4
O	16	36.39	$\frac{36.39}{16} = 2.27$	$\frac{2.27}{2.27} = 1.0$	1

$\therefore$ Empirical formula = C_2H_4O

$\therefore$ The empirical formula mass of organic substance = 24 + 4 + 16

= 44 gram formula weight^{-1}

$$\text{Integral number (n)} = \frac{\text{Molecular mass}}{\text{Formula mass}}$$

$\therefore$ n = 88/44 = 2

n = 2

Molecular formula = 2 × empirical formula

= 2 × C_2H_4O

= $C_4H_8O_2$

1.8 Stoichiometry of Chemical Reactions and Calculations

Proper distributions of the element in equation means stoichiometry. Stoichiometry, thus deals with the calculation of masses of the reactants and the products involved in a chemical reaction.

A chemical reaction can be expressed in the form of chemical equation from which we obtain a large amount of qualitative and quantitative information. Quantitative information is avaliable from balanced chemical equation available from stoichiometry of given reaction. A balanced equation for combustion of methane is given below. Let us consider what information is available from stoichiometry.

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$$

	CH_4	O_2	CO_2	H_2O
Mole	1 mole	2 moles	1 mole	2 moles
Molecules	1 molecule	2molecules	1molecule	2molecules
Volume*	22.4 L	2×22.4 = 44.8L	22.4 L	2 × 22.4 = 44.8 L
Mass	16.00 g	2 × 32 = 64 g	44.0g	2 × 18= 36 g
Number of molecules	6.022×10^{23}	$2 \times 6.022 \times 10^{23}$	6.022×10^{23}	$2 \times 6.022 \times 10^{23}$

*** At STP**

The stoichiometry of the reactants in the above reaction is 1:2 and the stoichiometry of the product is also 1:2. Thus, for a general reaction aA+bB → cC+dD is the stoichiometries of reactants and products are a:b and c:d respectively.

1.8.1 Balancing of a Chemical reaction equation :

According to the law of conesrvation of mass, a balanced chemical reaction equation has the same number of atoms of each element on both the sides of the equation. If chemical equation is unbalanced, try to balance it by trial and error method with the use of proper multiple number. The method to balance the equation by same number of atoms of each element on both sides is known as method of balancing equation. You have studied in standard 9 the balancing of chemical reaction equation. For example,

$P_{4(s)} + O_{2(g)} \rightarrow P_4O_{10(s)}$ (Unbalanced equation)

In this equation, phophorus atoms are balanced but not the oxygen atoms. To balance them we must place the multiple 5 on the left of oxygen.

$P_{4(s)} + 5O_{2(s)} \rightarrow P_4O_{10(s)}$ (Balanced equation)

The reactions given below are not balanced :

(i) $PbS + O_2 \rightarrow PbO + SO_2$

(ii) $Pb_3O_4 + 4HNO_3 \rightarrow Pb(NO_3)_2 + PbO_2 + H_2O$

(iii) $Ca_3(PO_4)_2 + H_2SO_4 \rightarrow Ca(H_2PO_4)_2 + CaSO_4$

(iv) $Cu_2S + O_2 \rightarrow Cu_2O + SO_2$

(v) $MnO_2 + KOH + O_2 \rightarrow 2K_2MnO_4 + H_2O$

After balancing the above reactions they are written as

(i) $2PbS + 3O_2 \rightarrow 2PbO + 2SO_2$

(ii) $Pb_3O_4 + 4HNO_3 \rightarrow 2Pb(NO_3)_2 + PbO_2 + 2H_2O$

(iii) $Ca_3(PO_4)_2 + 2H_2SO_4 \rightarrow Ca(H_2PO_4)_2 + 2CaSO_4$

(iv) $2Cu_2S + 3O_2 \rightarrow 2Cu_2O + 2SO_2$

(v) $2MnO_2 + 4KOH + O_2 \rightarrow 2K_2MnO_4 + 2H_2O$

1.9 Stoichiometry of Reaction in Solution

A majority of reactions are carried out in solution in the laboratories. Therefore, it is important to understand as how the amount of substance is expressed when it is present in the form of solution. The amount of substance present in its given volume can be expressed in units like normality, molarity, molality, mole fraction, w/w percentage. Let us now discuss and study each one of them in detail.

(i) Normality (N) : One litre of solution prepared by dissolving one gram equivalent of a substance is called 1 normal (N) solution or the normality of the solution is 1

$$\text{Normality} = \frac{\text{gram litre}^{-1}}{\text{equivalent weight}}$$

$$= \frac{\text{weight of solute in gram}}{\text{equivalent weight} \times \text{volume of solution in litre.}}$$

$$\text{Equivalent weight of acid} = \frac{\text{Molecular weight of acid (gmol}^{-1})}{\text{Basicity of acid}}$$

$$\text{Equivalent weight of base} = \frac{\text{Molecular weight of base (gmol}^{-1})}{\text{Acidity of base}}$$

Exampale 1.5 : Find the normality of an aqueous solution in which 73 gram hydrochloric acid is dissolved in 500 ml solution.

Solution :

Molar mass of HCl = 36.5 gram mole^{-1}.

Molar mass and equivalent weight of HCl are equal.

Equivalent weight of HCl = 36.5 gram equi^{-1}

Volume of solution 500 ml = 0.5 litre.

Weight of solute HCl in solution = 73 gram

$$\text{Normality} = \frac{\text{weight of solute in gram}}{\text{equivalent weight} \times \text{volume of solution in litre}}$$

$$= \frac{73}{36.5 \times 0.5} = 4\text{N}$$

(ii) Molarity (M) : One litre of solution containing one gram mole of a substance is called 1 molar (M) solution or the molarity of soution is 1. It is the most widely used unit.

$$\text{Molarity} = \frac{\text{gram litre}^{-1}}{\text{molecular mass}}$$

$$= \frac{\text{weight of solute in gram}}{\text{molecular mass} \times \text{volume of solution in litre}}$$

Note that normalities and molarities are dependent on temperature because of the change in the volume of solution with temperature.

Example 1.6 : Find the molarity of an aqueous solution in which 4 gram NaOH is dissolved in 5 litre aqueous solution.

Solution :

Molar mass of NaOH = 40 gram mole^{-1}

Volume of solution = 5 liter

Weight of solute NaOH = 4 gram

$$\text{Molarity} = \frac{\text{weight of solute in gram}}{\text{molecular mass} \times \text{volume of solution in litre}}$$

$$= \frac{4}{40 \times 5}$$

= 0.02 M

(iii) Molality (m) : One gram mole of a solute, when dissolved in 1 kilogram of the solvent, the molality (m) of the solution obtained is 1 molal (m) or it is 1 m solution.

$$\text{Molality} = \frac{\text{weight of solute in gram}}{\text{molecular mass} \times \text{weight of solvent in kilogram}}$$

Note that with the change in temperature the molality does not change because the weight is not affected by temperature.

Example 1.7 : Find the molality of the solution when 149 gram of KOH is dissolved in 1.5 kilogram of solvent.

Solution :

Molecular mass of KOH = 56 gram mole^{-1}

Weight of solute KOH = 149 gram

Mass of solvent = 1.5 kilogram

$$\text{Molality} = \frac{\text{weight of solute in gram}}{\text{molecular mass} \times \text{weight of solvent in kilogram}}$$

$$= \frac{149}{56 \times 1.5}$$

= 1.77 m

(iv) Mole Fraction : The ratio of the moles of any component to the total moles in a solution is the mole fraction.

Mole fraction of component

$$= \frac{\text{moles of component}}{\text{number of total moles of components in a solution}}$$

The sum of the mole fraction of all the components in a solution is always equal to 1.

Example 1.8 : Caclulate the mole fraction of NaOH and water in a solution formed by dissolving 4 gram of NaOH in 180 gram of water

Solution :

Molar mass of NaOH = 40 gram $mole^{-1}$

$$\text{Moles of (solute) NaOH} = \frac{\text{weight}}{\text{molar mass}}$$

$$= \frac{4}{40} = 0.1 \text{ mole}$$

$$\text{Moles of } H_2O \text{ (solvent)} = \frac{\text{weight}}{\text{molar mass}}$$

$$= \frac{180}{18} = 10 \text{ mole}$$

Total moles = 0.1 mole + 10 mole = 10.1 mole

Mole fraction of NaOH

$$= \frac{\text{moles of NaOH}}{\text{total moles of solution}}$$

$$= \frac{0.1}{10.1}$$

= 0.0099 mole

Mole fraction of H_2O

$$= \frac{\text{moles of } H_2O}{\text{total moles of solution}}$$

$$= \frac{10 \text{ mole}}{10.1 \text{ mole}}$$

= 0.9901 mole

$\therefore$ Sum of mole fractions =

mole fraction of NaOH + mole fraction of H_2O

= 0.0099 + 0.9901

= 1

$\therefore$ The sum of the mole fractions of all the components is 1

(v) Precentage by weight : (% w/w)

The weight of a substance in gram dissolved in 100 gram solution is called percentage by weight (% w/w). Such a solution is called percentage proportion with reference to the weight of solute.

$$\% \text{ w/w} = \frac{\text{weight of solute} \times 100}{\text{weight of solution in gram}}$$

(weight of solution = weight of solute + weight of solvent)

With the change in temperature, the values of molality, mole fraction, percentage w/w do not change.

Example 1.9 : How many grams of NaOH will be required to prepare 500 gram solution containing 5% w/w NaOH ?

Solution :

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

$$5\% = \frac{\text{weight of solute} \times 100}{500}$$

$$\text{weight of solute} = \frac{5 \times 500}{100}$$

= 25 gram

SUMMARY

The study of chemistry is very imporant because its domain has encompassed to cover every field of life. Chemists have studied the composition and properties of substances and transformations they undergo. The arrangement of constituents in three states of matter is different which indicates its properties. Matter can be classified into element, mixture and compound. If the element is formed, only of one type of constitutent, it can be atom or molecule. When atoms of two or more elements combine with each other in definite proportion, compound is formed. Mixture is formed in different ways.

When the study of properties of a substance is carried out the measurement becomes equally important. At the same time universally accepted unit method is essential. Hence, it is agreed upon that there must be a universally same and general method of measurement. It is known in brief as SI unit.

Atomic analysis is useful in expressing the units of physical quantities in different ways.

The combination of different atoms can be determined from the fundamental laws of chemistry. These laws are law of mass conservation, law of definite proprtion, law of multiple proportion; law of combining weights etc. All these laws lead to atomic theory of Dalton.

The atomic weight of an element is related to ^{12}C. The atomic weight of carbon-12 is 12u. The atomic weights of remaining elements are determined by IUPAC and IUPAP considering this as standard. The molecular mass of a molecule is obtained by calculation and adding atomic masses of all atoms. The molecular formula of a compound, the masses of the diffrernt elements present in molecular mass is determined by determination of percentage of elements present and determining empirical formula.

The number of atoms, molecules or any component in any system is expressed with reference to Avogadro constant 6.0×10^{23}. This is known as 1 mole. The quantitative study of reactants and products is called stoichiometry.

Chemical reactions show the chemical changes in different elements and compounds. A balanced chemical equation provides many more informations.

Apart from this, the different methods to express the proportion of the substance in a solution are mole fraction, weight percentage, normality molarity and molality.

EXERCISE

1. Select the proper choice from the given multiple choices :

(1) Which form of classification is not correct for matter on the basis of chemical characteristics ?

(A) Mixture (B) Gas

(C) Volume (D) Solid

(2) Which of the following units is a derived unit ?

(A) Density (B) A and C

(C) Volume (D) Time

(3) What is the SI unit of mass of the matter ?

(A) kilogram (B) gram

(C) milligram (D) mole

(4) Of which alloy the cylinder standardising kilogram value of mass is made of ?

(A) Pt - Au (B) Pt - Ag

(C) Pt - Ir (D) Pt - Sn

(5) What is called the mass of the matter in unit volume ?

(A) Density (B) Volume

(C) Weight (D)Pressure

(6) What degree Farheinheit will be equivalent to 25^0C ?

(A) 298 F (B) 77 F

(C) 32 F (D) 248 F

(7) With which reference there is no importance of Dalton's law ?

(A) Isotope (B) Isobar

(C) Isoosmotic (D) Isothermic

(8) What is the value of one atom of ^{12}C obtained by mass spectrometer ?

(A) 12 gram (B) 1.992648×10^{-23} gram

(C) 1/12 gram (D) 6.022×10^{23} gram

(9) The value of sum of the mole fractions of all the components in a solution will be.........

(A) >1 (B) <1

(C) =1 (D) =0

(10) Which of the following changes with temperature ?

(A) Molality (B) Molarity

(C) % w/w (D) Mole fraction

(11) Molartiy means .... dissolved in one litre solution

(A) One mole (B) One equivalent

(C) One gram (D) 1000 gram

(12) How many decimeter^3 will be equal to 1000 ml ?

(A) 1 (B) 10

(C) 100 (D) 0.10

(13) Which of the following principle or law is not associated with chemical combination ?

(A) Law of constant proportion (B) Law of combining weight

(C) Law of multiple proportion (D) Aufbau Principle

(14) 1 amu is equal to how many grams ?

(A) 1.66056×10^{-24} gram (B) 6.022×10^{23} gram

(C) 9.191×10^{-28} gram (D) 1.992648×10^{-23} gram

(15) Concentration of solution of 2N H_2SO_4 is .... M

(A) 1 (B) 4

(C) 0.5 (D) 2

2. Write the answers of the following questions in short :

(1) Write the law of definite proportion.

(2) Mention the four importances of chemistry.

(3) Classify matter on the basis of chemical characteristics.

(4) What is a compound and a mixture ?

(5) What is heterogeneous mixture ?

(6) Mention the formula of the relation between the two units of temperature.

(7) Which are the laws of chemical combination ?

(8) Find the number of atoms in 1 mole carbon.

(9) Define : Normality, molarity, molality mole fraction.

3. Write the answers of the following questions :

(1) Mention the differences between homogeneous mixture and heterogeneous mixture.

(2) Expain in detail the derived units-volume, density and temperature.

(3) Balance the equations of following reactions :

(i) $HgS + CaO \xrightarrow{\Delta} Hg + CaSO_4 + CaS$

(ii) $Na_2CrO_4 + H^+ \rightarrow Na_2Cr_2O_7 + Na^+ + H_2O$

(iii) $MnO_2 + KOH + O_2 \rightarrow K_2MnO_4 + 2H_2O$

(iv) $Al_2O_3 + NaOH + H_2O \rightarrow Na[Al(OH)_4]$

(4) Find the percentage mass composition of each element present in $H_2C_2O_4$.

(5) In how many grams of water 36.5 gram HCl should be dissolved so that 10 % w/w solution will be obtained ?

(6) Find the molality of the solution obtained when 63 gram HNO_3 is dissoloved in 750 gram water.

4. Write the answers of the following questions in detail :

(1) Explain the stoichiometry of chemical reaction by giving example.

(2) Write a note on the mole concept.

(3) Write a note on SI unit method and explain length and temperature.

(4) Mention the hypotheses of Dalton's atomic theory.

(5) The percentage proportion of carbon, hydrogen, nitrogen and oxygen in an organic compound are 62.07 %, 10.34 %, 14.0 %, 13.59 %, respectively. Find its empirical formula. If its molecular mass is 144 gram mol^{-1}, find its molecular formula.

(6) Explain in detail atomic mass and molecular mass by giving examples.