SOME BASIC CONCEPTS OF CHEMISTRY

1.1 Importance of Chemistry :

Q. What is chemistry ?

Solution : Chemistry is defined as that branch of science which deals with the study of composition, structure and properties of matter and the changes which the matter undergoes under different conditions and the laws which govern these changes.

Q. Briefly discuss the importance of chemistry.

Solution : (1) Chemical industries manufacturing fertilizers, alkalis, acids, salts, dyes, polymers, drugs, soaps, detergents, metals, alloys and other inorganic and organic chemicals, including new materials, contribute in a big way to the national economy.

(2) Chemistry plays an important role in meeting human needs for food, health care products and other materials aimed at improving the quality of life. This is exemplified by the large scale production of a variety of fertilizers, improved varieties of pesticides and insecticides.

(3) Similarly many life saving drugs such as cisplatin and taxol, are effective in cancer therapy and **AZT** (Azidothymidine) used for helping AIDS victims, have been isolated from plant and animal sources or prepared by synthetic methods.

(4) In recent years chemistry has tackled with a fair degree of success some of the pressing aspects of environment degradation. Safer alternatives to environmentally hazardous refrigerants like **CFCs** (chlorofluorocarbons), responsible for ozone depletion in the stratosphere, have been successfully synthesised.

1.2 Nature of Matter :

Q. Define matter. Briefly describe the physical as well as chemical classification of matter.

Solution : Anything which has mass and occupies space is called matter. Everything around us, for example, book, pencil, water, air, all living being etc. are composed of matter.

There are two ways of classifying matter : (a) Physical classification (b) Chemical classification

(a) Physical classification : Based on physical state under ordinary conditions of temperature and pressure, matter is classified into the following three types : (i) Solids (ii) Liquids (iii) Gases.

In a **solid**, the constituent particles are very closely packed in an orderly fashion and there is not much freedom of movement (except vibratory motion). Hence, they possess a definite shape and a definite volume.

In a **liquid**, the particles are comparatively less close to each other and, therefore, they can move around. Hence, they have a definite volume but no definite shape.

In a **gas**, the particles are far apart and therefore have a much greater freedom of movement. Hence, they possess neither definite volume nor definite shape.

(b) Chemical classification : Many of the substances present around you are mixtures. For example, sugar solution in water, air, tea etc., are all mixtures. A mixture contains two or more substances present in it (in any ratio) which are called its components. A mixture may be **homogeneous** or **heterogeneous**.

In a **homogeneous** mixture, the components completely mix with each other and its composition is uniform throughout. Sugar solution, and air are thus, the examples of **homogeneous** mixtures.

Heterogeneous mixtures, the composition is not uniform throughout and sometimes the different components can be observed. For example, the mixtures of salt and sugar, grains and pulses along with some dirt (often stone) pieces, are **heterogeneous** mixtures.

Q. What are pure substances. What is difference between pure substance and mixture.

Solution : Pure substances have characteristics different from the mixtures. They have fixed composition, whereas mixtures may contain the components in any ratio and their composition is variable. Copper, silver, gold, water, glucose are some examples of pure substances. Glucose contains carbon, hydrogen and oxygen in a fixed ratio and thus, like all other pure substances have a fixed composition. Also, the constituents of pure substances cannot be separated by simple physical methods, whereas mixture can be seperated by physical method e.g., distillation, crystallisation etc.

Q. Discuss briefly the classification of pure substances.

Solution : Pure substances can be further classified into **elements** and **compounds**.

Element : consists of only one type of particles. These particles may be atoms or molecules.

Sodium, copper, silver, hydrogen, oxygen etc. are some examples of elements. They all contain atoms of one type. However, the atoms of different elements are different in nature. Some elements such as sodium or copper, contain single atoms held together as their constituent particles whereas in some others, two or more atoms combine to give molecules of the element. Thus, hydrogen, nitrogen and oxygen gases consist of molecules in which two atoms combine to give their respective molecules.

Compound : When two or more atoms of different elements combine, the molecule of a compound is obtained. The examples of some compound are water, ammonia, carbon dioxide, sugar etc.

Thus, the atoms of different elements are present in a compound in a fixed and definite ratio and this ratio is characteristic of a particular compound.

The properties of a compound are different from those of its constituent elements.

1.3 Properties of Matter and their Measurements :

Q. What are the difference types of properties of substances ?

Solution : These properties can be classified into two categories – **physical properties** and **chemical properties**.

(1) **Physical properties** are those properties which can be measured or observed without changing the identity or the composition of the substance. Some examples of physical properties are colour, odour, melting point, boiling point, density etc.

(2) **Chemical properties** require a chemical change to occur. The examples of chemical properties are characterisitic reactions of different substances; these include acidity or basicity, combustibility etc.

Q. What are the different system of units ?

Solution : In earlier time three such systems, the CGS, the FPS (or British) system and the MKS system were in use extensively till recently.

The base units of length, mass and time in these systems were as follows :

- In CGS system they were centimetre, gram and second respectively.
- In FPS system they were foot, pound and second respectively.
- In MKS system they were metre, kilogram and second respectively.

The system of units which is at present internationally accepted for measurement is the Systeme Internationale d' United (French for International System of Units), abbreviated as SI. The SI, with standard scheme of symbols, units and abbreviations, was developed and recommended by General Conference on Weights and Measures in 1971 for international usage in scientific, technical, industrial and commercial work.

Q. In SI, what are the seven fundamental (base) quantities ? Write down their units and define them.

Solution : In SI, there are seven base units which are :

(a) Length : The unit is metre (m). The metre is the length of the path travelled by light in vacuum during a time interval of 1/299, 792, 458 of a second. (1983)

(b) Mass : The unit is kilogram (kg). The kilogram is equal to the mass of the international prototype of the kilogram (a platinum-iridium alloy cylinder) kept at international Bureau of Weight and Measures, at Serves, near Paris, France. (1889)

(c) Time : The unit is second (s). The second is the duration of 9, 192, 631, 770 periods of the radiation corresponding to the transition between the two hyperfine levels of the ground state of cesium-133 atom. (1967)

(d) Electric current : The unit is ampere (A). The ampere is that constant current which, if maintained in two straight parallel conductors of infinite length, of negligible circular cross-section, and placed 1 metre apart in vacuum, would produce between these conductors a force equal to 2×10^{-7} newton per metre of length. (1948)

(e) Thermo dynamic Temperature : The unit is kelvin (K). The kelvin, is the fraction 1/273.16 of the thermodynamic temperature of the triple point of water. (1967)

(f) Amount of substance : The unit is mole (mol). The mole is the amount of substance of a system, which contains as many elementary entitles as there as atoms in 0.012 kilogram of carbon - 12. (1971)

(g) Luminous intensity : The unit is candela (cd). The candela is the luminous intensity, in a given direction, of a source that emits monochromatic radiation of frequency 540×10^{12} hertz and that has a radiant intensity in that direction of 1/683 watt per steradian. (1979)

Q. What is the advantage of SI system over other system of units ?

Solution : (i) SI is a coherent system of units i.e., a system based on a certain set of fundamental units, from which all derived units are obtained. (ii) SI is a rational system of units, as it assigns only one unit to a particular physical quantity. For example, joule is the unit for all types of energy. (iii) SI unit used decimal system, conversions within the system are quite simple and convenient.

Q. What is the different between mass and weight ?

Solution : Mass of a substance is the amount of matter present in it while **weight** is the force exerted by gravity on an object. The mass of a substance is constant whereas its weight may vary from one place to another due to change in gravity.

Q. Write the relation between ^{0}F and ^{0}C.

Solution : $^0\mathbf{F} = \frac{9}{5}(^0\mathbf{C}) + 32$.

Q. At what temperature will both the Celsius and Fahrenheit scale read the same value ?

Solution : Suppose both read the same value as 'x'.

Then as $^0\mathbf{C} = \frac{5}{9}(^0\mathbf{F} - 32)$

$\therefore$ $\mathbf{x} = \frac{5}{9}(\mathbf{x} - 32)$

$x = -40^0$

1.4 Uncertainity in Measurement :

Q. What is scientific notation ? Explain it with example.

Solution : Scientific notation for numbers i.e., exponential notation in which any number can be represented in the form $N \times 10^n$ where n is an exponent having positive or negative values and N can vary between 1 to 10.

Thus, we can write 232.508 as 2.32508×10^2 in scientific notation.

Similarly, 0.00016 can be written as 1.6×10^{-4}.

Q. What is the difference between accuracy and precision ? Give example.

Solution : The accuracy of a measurement is a measure of how close the measured value is to the true value of the quantity. Precision tells us to what resolution or limit the quantity is measured.

The accuracy in measurement may depend on several factors, including the limit or the resolution of the measuring instrument. For example, suppose the true value of a certain length is near 3.678 cm. In one experiment, using a measuring instrument of resolution 0.1 cm, the measured value is found to be 3.5 cm, while in another experiment using a measuring device of greater resolution, say 0.01 cm, the length is determined to be 3.38 cm. The first measurement has more accuracy (because it is closer to the true value) but less precision (its resolution is only 0.1 cm), while the second measurement is less accurate but more precise.

Q. What is significant figure ?

Solution : Every measurement involves errors. Thus, the result of measurement should be reported in a way that indicates the precision of measurement. Normally, the reported result of measurement is a number that includes all digits in the number that are known reliably plus the first digit that is uncertain. The reliable digits plus the first uncertain digit are known as significant digits or significant figures. Take the example that the period of a simple pendulum is 1.62 s, the digits 1 and 6 are reliable and certain while the digit 2 is uncertain.

Q. Write down the rules for significan figure.

Solution : Rules for determining the number of siginificant figures :

Rule 1 : All the non-zero digits are significant.

Rule 2 : All the zeros between two non-zero digits are significant, no matter where the decimal point is, if at all.

Rule 3 : If the number is less than 1, the zeros on the right of decimal point but to the left of the first non-zero digit are not significant. For example $\underline{0}.\underline{00}2308$, the underlined zeros are not significant.

Rule 4 : The terminal or trailing zeros in a number without a decimal point are not significant. e.g. 123 m = 12300 cm = 123000 mm has three significant figures.

Rule 5 : The trailing zeros in a number with a decimal point are significant. e.g. 3.500 has four significant figure.

Rule 6 : A choice of change of different units does not change the number of significant figures in a measurement. For example A length is reported to be 4.700 m. Now suppose we change units, then 4.700 m = 470.0 cm = 4700 mm = 0.004700 km numbers of significant figures in this remains 4. To avoid above confusion, use scientific notation. 4.700m = 4.700×10^2 cm = 4.700×10^3 mm. The power of 10 is irrelevent to the determination of significant figures.

Rule 7 : The digit 0 conventionally put on the left of a decimal for a number less than 1 (like 0.1250) is never significant. However, the zeroes at the end of such number are significant in a measurement.

Rule 8 : The multiplying or dividing factors which are neither rounded numbers nor numbers representing measured values, are exact and have infinite number of significant digits. For example in $\mathbf{r = \frac{d}{2}}$ or $s = 2\pi r$, the factor 2 is an exact number and it can be written as 2.0, 2.00 or 2.0000 as required.

Q. Write down the number of significant figures in the following : (i) 5729 N (ii) 5.729 N (iii) 5729.00 N (iv) 5729 × 10^5 N (v) 5700 N (vi) 57.000 N (vii) 0.02370 N (viii) 0.02307 N (ix) 5.700 × 10^3 N

Solution : (i) 4 (ii) 4 (iii) 6 (iv) 4 (v) 2 (vi) 5 (vii) 4 (viii) 4 (ix) 4

Q. Write down the rules for rounding off the uncertain digits.

Solution : Rule 1 : The preceding digit is raised by 1 if the insignificant digit to be dropped (the underlined digit in the case) is more than 5, and is left unchanged if the latter is less than 5. For e.g., A number 2.746 rounded off to three significant figures is 2.75, while the number 2.743 would be 2.74.

Rule 2 : If the proceding digit is even, the insignificant digit is simply dropped and, if it is odd, the proceding digit is raised by 1. Then, the number 2.745 rounded off to three significant figures becomes 2.74. On the other hand, the number 2.735 rounded off to three significant figures becomes 2.74 since the preceding digit is odd.

Q. Write down the different rules for Arithmetic Operations with significant figues.

Solution : (1) In multiplication or division, the final result should retain as many significant figures as are there in the original number with the least significant figures. For example, if mass of an object is measured to be, say, 4.237 g (four significant figures) and its volume is measured to be 2.51 cm^3 (3 significant figures) then if its density will contain 3 significant figure i.e, $\mathbf{Density = \frac{4.237\,g}{2.51\,cm^3} = 1.69\,g\,cm^{-3}}$.

(2) In addition or subtraction, the final result should retain as many decimal place as are there in the number with the least decimal places. For e.g., the sum of the numbers 436.32 g, 227.2 g and 0.301 g by mere arithmetic addition, is 663.821 g. The final result should, therefore, be rounded off to 663.8 g.

Q. How many significant figures are there in each of the following numbers ?

(i) 6.005 (ii) 6.022 × 10^{23} (iii) 8000 (iv) 0.0025 (v) π (vi) the sum 18.5 + 0.4235 (vii) the product 14 × 6.345.

Solution : (i) Four because the zeros between the non-zero digits are significant figures.

(ii) Four because only the first term gives the significant figures and exponential term is not considered.

(iii) Four. However, if expressed in scientific notation as 8×10^3, it will have only one significant figure, as 8.0×10^3, 8.00×10^3 or 8.000×10^3, it will have 2, 3 and 4 significant figures.

(iv) Two because the zeros on the left of the first non-zero digit are not significant.

(v) As $\pi = \frac{22}{7}$ = 3.1428571.........hence it has infinite number of significant figures.

(vi) Three because the reported sum will be only upto one decimal place, i.e., 18.9.

(vii) Two because the number with least number of significant figures involved in the calculation (i.e., 14) has two significant figures.

Q. A sample of nickel weighs 6.5425 g and has a density of 8.8 g/cm³. What is the volume ? Report the answer to correct decimal place.

Solution : $\mathbf{Volume = \frac{Mass}{Density} = \frac{6.5425}{8.8} = 0.74\,cm^3}$

The result should have two significant figures because the least precise term (8.8) has two significant figures.

1.5 Laws of Chemical Combinations :

Q. Define the law of conservation of mass with example.

Solution : According to this law "the total mass of reactants are conserved in the products if they are not allowed to escape" i.e. during the chemical changes mass is neither created nor destroyed. e.g.,

TEST : Consider the following reaction in a closed container

$$Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$$

Total mass of reactant (from balanced equation)

= Mass of 1 mole of Cu + Mass of 2 moles of H_2SO_4

= 63.5 + 196 = 259.5 g

Total mass of products = $\mathbf{W_{CuSO_4} + W_{SO_2} + 2 \times W_{H_2O}}$

= 159.5 + 64 + 36 = 259.5 g

Q. 90 g of $KClO_3$ when heated produced 1.92 g of oxygen and the residue (KCl) left behind weighs 2.96 g. Show that these results illustrate the law of conservation of mass.

Solution : Mass of $KClO_3$ taken = 4.90 g.

Total mass of the products ($KCl + O_2$) = 2.96 + 1.92 = 4.88 g

Difference between the mass of the reactant and the total mass of the products = 4.90 – 4.88 = 0.02 g.

This small difference may be due to experimental error.

Thus, law of conservation of mass holds good within experimental errors.

Q. Define law of definite proportions with example ?

Solution : The law states that "a chemical compound always contains the same elements combined together in the same proportion by mass". In another words, in a chemical compound the elements are present in a fixed and not in arbitrary ratio by mass. For example :

Pure water collected from any source always contains hydrogen and oxygen combined together in the same proportion of 1 : 8 by mass.

$H_2O \rightarrow 2H : O$

i.e., 2 : 16

i.e., 1 : 8

Q. Define law of multiple proportions.

Solution : It states that if two elements can combine to form more than one compoun, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers. For examples :

(A) (i) Carbon + Oxygen = Carbon mono-oxide

(ii) Carbon + Oxygen = Carbon dioxide

Consider the equations (i) and (ii)

(i) $$C+\frac{1}{2}O_2 = CO$$

12g 16g = 12 : 16

(ii) $C + O_2 = CO_2$

12g 32g = 12 : 32

$$\therefore \quad \frac{O\,in\,CO}{O\,in\,CO_2} = \frac{16}{32} = \frac{1}{2}$$

Q. What are the limitations of law of definite proportions ?

Solution : (1) The law is not applicable if an element exists in different isotopes which may be involved in the formation of the compound. For example, in the formation of the compound CO_2, if C-12 isotope combines, the ratio of C : O is 12 : 32, but if C-14 isotope combines, the ratio of C : O is 14 : 32.

(2) The elements may combine in the same ratio but the compounds formed may be different. For example, in the compound, C_2H_5OH and CH_3OCH_3 (both having same molecular formula viz. C_2H_6O) the ratio of C : H : O = 24 : 6 : 16 = 12 : 3 : 8 by mass.

Q. Define Gay Lussac's law of gaseous volumes with example.

Solution : He observed that when gases combine or are produced in a chemical reaction they do so in a simple ratio by volume provided all gases are at same temperature and pressure.

Thus, 100 mL of hydrogen combine with 50 mL of oxygen to give 100 mL of water vapour.

Hydrogen + Oxygen $\rightarrow$ Water

100 mL 50 mL 100 mL

Thus, the volumes of hydrogen and oxygen which combine together (i.e., 100 mL and 50 mL) bear a simple ratio of 2 : 1.

Q. Two oxides of metal contain 27.6% by mass and 30.0% by mass of oxygen respectively. If the formula of the first oxide is M_3O_4, find that of the second.

Solution : M_2O_3

Q. Define Avogadro's Law.

Solution : Avogadro proposed that equal volumes of gases at the same temperature and pressure should contain equal number of molecules.

1.6 Dalton's Atomic Theory :

Q. What are the main points of Dalton's atomic theory ?

Solution : (1) Matter consists of indivisible atoms.

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

(3) Compounds are formed when atoms of different elements combine in a fixed ratio.

(4) Chemical reactions involve reorganisation of atoms. These are neither created nor destroyed in a chemical reaction.

Dalton's theory cound explain the laws of chemical combination.

1.7 Atomic and Molecular Mass :

Q. Using which method, mass of atom/sub-atomic particles is measured.

Solution : For measurement of small masses of atomic/sub-atomic particles etc., we make use of mass spectrograph in which radius of the trajectory is proportional to the mass of a charged particle moving in uniform electric and magnetic field.

Q. Define unified atomic mass unit.

Solution : unifined atomic mass unit (u), which has be established for expressing the mass of atoms as

1 unified atomic mass unit = 1u

= (1/12) of the mass of an atom of carbon-12 isotope $\left({}^{12}_{6}\mathrm{C}\right)$ including the mass of electrons

= 1.66×10^{-27} kg.

Q. Why there is a need of average atomic mass.

Solution : Many naturally occurring elements exist as more than one isotope. When we take into account the existence of these isotopes and their relative abundance (per cent occurrence) the average atomic mass of that element can be computed.

Q. Define atomic mass and how it is determined.

Solution : The atomic mass of an element is the average relative mass of its atoms as compared with an atom of carbon-12 taken as 12.

The atomic masses of the elements have been determined accurately during the recent years using an instrument called **'mass spectrometer'**.

Q. Chlorine has two isotopes of atomic mass units 34.97 and 36.97. The relative abundances of these two isotopes are 0.755 and 0.245 respectively. Find the average atomic mass of chlorine.

Solution : Average atomic mass = $34.97 \times 0.755 + 36.97 \times 0.245 = 35.46$

Q. Define molecular mass.

Solution : Molecular mass is the sum of atomic masses of the elements present in a molecule. It is obtained by multiplying the atomic mass of each element by the number of its atoms and adding them together.

Q. Calculate molecular mass of glucose ($C_6H_{12}O_6$) molecule.

Solution : Molecular mass of glucose ($C_6H_{12}O_6$)

= 6(12.011 u) + 12(1.008 u) + 6(16.00 u)

= 180.162 u

Q. For what type of compounds formula mass is defined. Explain with example ?

Solution : Some substances such as sodium chloride do not contain discrete molecules as their constituents units. In such compounds, positive (sodium) and negative (chloride) entities are arranged in a three-dimensional structure.

It may be noted that in sodium chloride, one Na^+ is surrounded by six Cl^- and vice-versa.

The formula such as NaCl is used to calculate the formula mass instead of molecular mass as in the solid state sodium chloride does not exist as a single entity.

1.8 Mole Concept and Molar Masses :

Q. Define mole and explain it with example.

Solution : One mole is the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 g (or 0.012 kg) of the ^{12}C isotope.

(1) We can, therefore, say that 1 mol of hydrogen atoms = 6.022×10^{23} atoms

(2) 1 mol of water molecules = 6.022×10^{23} water molecules

(3) 1 mol of sodium chloride = 6.022×10^{23} formula units of sodium chloride

Q. Define molar mass.

Solution : The mass of one mole of a substance in grams is called its molar mass.

Q. How many atoms and molecules of sulphur are present in 64.0 g of sulphur (S_8) ? Given : At. mass of sulphur is 32 g/mol.

Solution : 1.20×10^{24} atoms

Q. Calculate the number of molecules present (i) in 34.20 grams of can sugar ($C_{12}H_{22}O_{11}$) (ii) in one litre of water assuming that density of water is 1 g/cm^3. (iii) in one drop of water having mass 0.05 g. Given (At. mass C = 12 u, H = 1 u, O = 16 u)

Solution : (i) 6.022×10^{22} molecules (ii) 3.345×10^{25} (iii) 1.67×10^{21}

1.9 Percentage Composition :

Q. How do you calculate the mass percentage of an element in the compound ?

Solution : Mass % of an element = $\frac{\textbf{mass of that element in the compound} \times 100}{\textbf{molar mass of the compound}}$

Q. What is the percentage of carbon, hydrogen and oxygen in ethanol ? [NCERT solved example]

Solution : 52.14%, 13.13%, 34.73%

Q. What is emperical and molecular formula ?

Solution : An **empirical formula** represents the simplest whole number ratio of various atoms present in a compound whereas the **molecular formula** shows the exact number of different types of atoms present in a molecule of a compound.

Q. A compound contains 4.07 % hydrogen, 24.27 % carbon and 71.65 % chlorine. Its molar mass is 98.96 g. What are its empirical and molecular formulas ? [Given, At. Mass : C = 12, H = 1, Cl = 35.453]

Solution :

Element	Percentage of element	moles of the element	simplest molar ratio
C	24.27	$\frac{24.27}{12} = 2.021$	$\frac{2.02}{2.02} = 1$
H	4.07	$\frac{4.07}{1.008} = 4.04$	$\frac{4.04}{2.02} = 2$
Cl	71.65	$\frac{71.65}{35.453} = 2.02$	$\frac{2.02}{2.02} = 1$

Thus emperical formula is CH_2Cl

Emperical mass = 12 + 2 + 35.453 = 49.453

(common molar ratio) n = $\frac{\textbf{molar mass}}{\textbf{empirical mass}} = \frac{98.96}{49.453} = 2$

Molecular formula = $(CH_2Cl)_2 = C_2H_4Cl_2$

Q. A crystalline salt on being rendered anhydrous loses 45.6% of its weight. The percentage composition of the anhydrous salt is :

Aluminium = 10.50 %;

Potassium = 15.1 %

Sulphur = 24.96 %

Oxygen = 49.92 %.

Find the simplest formula of the anhydrous and crystalline salt.

[Given At. mass : Al = 27 u, K = 39 u, S = 32 u, O = 16 u, H = 1 u]

Solution : To calculate the empirical formula of anhydrous salt

Element	Percentage of element	moles of the element	simplest molar ratio	simplest whole no. molar ratio
K	15.1	$\frac{15.10}{39} = 0.39$	$\frac{0.39}{0.39} = 1$	1
Al	10.50	$\frac{10.50}{27} = 0.39$	$\frac{0.39}{0.39} = 1$	1

S	24.96	$\frac{24.96}{32} = 0.78$	$\frac{0.78}{6.39} = 2$	2
O	49.92	$\frac{49.92}{16} = 3.12$	$\frac{3.12}{0.39} = 8$	8

Thus, the empirical formula of the anhydrous salt is $K\,AlS_2O_8$

Empirical formula mass of anhydrous salt = 258.0 u.

Loss of weight due to dehydration = 45.6 %

Therefore mass of anhydrous salt is = 54.4 u

Now, if the empirical formula mass of the anhydrous salt is 54.4, then that of hydrated salt = 100

Therefore, if the empirical formula mass of anhydrous salt is 258, that of hydrated salt = $\frac{100}{54.4} \times 258$ = 474.3 u.

Total loss in mass due to dehydration = 474.3 – 258.0 = 216.3 u

Number of molecules of water in hydrated sample = 216.3/18 = 12

Empirical formula of the hydrated salt = $KAlS_2O_8.12\,H_2O$.

1.10 Stoichiometry and Stoichiometric Calculations :

Q. What is the purpose of stoichiometry ?

Solution : Stoichiometry, thus deals with the calculation of masses (sometimes volumes also) of the reactants and the products involved in a chemical reaction.

Q. Calculate the amount of water (g) produced by the combustion of 16 g of methane. [NCERT solved example]

[Given At. Wt. : C = 12, H = 1, O = 16]

Solution : 36 g

Q. How many moles of methane are required to produce 22 g CO_2 (g) after combusion ? [NCERT solved example]

Solution : 0.5 mol.

Q. What is limiting reagent ?

Solution : The reactant which is present in the lesser amount gets consumed after sometime and after that no further reaction takes place whatever be the amount of the other reactant present. Hence, the reactant which gets consumed, limits the amount of product formed and is, therefore, called the **limiting reagent.**

Q. 50.0 kg of N_2(g) and 10.0 kg of H_2(g) are mixed to produce NH_3(g). Calculate the amount of NH_3(g) formed. Identify the limiting reagent in the production of NH_3 in this situation. [NCERT solved example]

Solution : Thus H_2 is a limiting reagent, 5.661×10^4 g.

Q. A solution is prepared by adding 2 g of a substance A to 18 g of water. Calculate the mass per cent of the solute. [NCERT solved example]

Solution : 10 %.

Q. If a substance 'A' dissolves in substance 'B' and their number of moles are n_A and n_B respectively; then find the molar fraction of A and B.

Solution : Mole fraction of A

$$= \frac{\text{No. of moles of A}}{\text{No. of moles of solution}} = \frac{n_A}{n_A + n_B}$$

Mole fraction of B

$$= \frac{\text{No. of moles of B}}{\text{No. of moles of solution}} = \frac{n_B}{n_A + n_B}$$

Q. Define molarity.

Solution : It is the most widely used unit and is denoted by M. It is defined as the number of the solute in 1 litre of the solution. Thus,

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

Q. Calculate the molarity of NaOH in the solution prepared by dissolving its 4 g in enough water to form 250 mL of the solution. [NCERT solved example]

Solution : 0.4 M

Q. The density of 3 M solution of NaCl is 1.25 g mL^{-1}. Calculate molality of the solution. [NCERT solved example]

Solution : 2.79 m

Q. How much volume of concentrated (1M) NaOH solution be taken which contains 0.2 moles of NaOH.

Solution : 200 mL

Q. If V_1 litre of M_1 molar solution is taken in a flask, if solvent is added to this solution so that its molarity decreases to V_2 litre then findout new molarity M_2 in terms of M1, V_1 and V_2, also find out volume of solvent needed to decrease the molarity of solution.

Solution : $M_2 = \dfrac{M_1V_1}{V_2}$

volume of solvent needed = $V_2 - V_1$

NCERT EXERCISE

1.1 Calculate the molecular mass of the following : [At. wt. : H = 1, C = 12, O = 16]

(i) H_2O (ii) CO_2 (iii) CH_4

1.2 Calculate the mass per cent of different elements present in sodium sulphate Na_2SO_4.

[At. wt. : Na = 23, S = 32, O = 16]

1.3 Determine the empirial formula of an oxide of iron which has 69.9% iron and 30.1% oxygen by mass. [At. wt. of Fe = 56, O = 16]

1.4 Calculate the amount of carbon dioxide that could be produced when

(i) 1 mole of carbon is burnt in air

(ii) 1 mole of carbon is burnt in 16 g of dioxygen

(iii) 2 moles of carbon are burnt in 16 g of dioxygen.

1.5 Calculate the mass of sodium acetate (CH_3COONa) required to make 500 mL of 0.375 molar aqueous solution. Molar mass of sodium acetate is 82.0248 g mol^{-1}.

1.6 Calculate the concentration of nitric acid in moles per litre in a sample which has a density, 1.41 g mL^{-1} and the mass per cent of nitric acid in it being 69%.

1.7 How much copper can be obtained from 100 g of copper sulphate ($CuSO_4$) ?

[At. wt. : Cu = 63.5, S = 32, O = 16]

1.8 Determine the molecular formula of an oxide of iron in which the mass per cent of iron and oxygen are 69.9 and 30.1 respectively. [Give At. mass of Fe = 56 u, O = 16 u]

1.9 Calculate the atomic mass (average) of chlorine using the following data :

	% Natural Abundance	Molar Mass
^{35}Cl	75.77	34.9689
^{37}Cl	24.23	36.9659

1.10 **In three moles of ethane (C_2H_6), calculate the following :**

(i) **Number of moles of carbon atoms**

(ii) **Number of moles of hydrogen atoms**

(iii) **Number of molecules of ethane**

1.11 **What is the concentration of sugar ($C_{12}H_{22}O_{11}$) in mol L^{-1} if its 20 g are dissolved in enough water to make a final volume up to 2 L ?**

1.12 **If the density of methanol is 0.793 kg L^{-1}, what is its volume needed for making 2.5 L of its 0.25 M solution ?**

1.13 **Pressure is determined as force per unit area of the surface. The SI unit of pressure, pascal is as shown below :**

$1Pa = 1N\ m^{-2}$

If mass of air at sea level is 1034 g cm^{-2}, calculate the pressure in pascal.

1.14 **What is the SI unit of masses ? How is it defined ?**

1.15 **Match the following prefixes with their multiples :**

	Prefixes	Multiples
(i)	micro	10^6
(ii)	deca	10^9
(iii)	mega	10^{-6}
(iv)	giga	10^{-15}
(v)	femto	10

1.16 **What do you mean by significant figures ?**

1.17 **A sample of drinking water was found to be severely contaminated with chloroform, $CHCl_3$ supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass).**

(i) **Express this in per cent by mass.**

(ii) **Determine the molality of chloroform in the water sample.**

1.18 **Express the following in the scientific notation :**

(i) 0.0048 **(ii)** 234,000 **(iii)** 8008 **(iv)** 500.0

(v) 6.0012

1.19 **How many significant figures are present in the following ?**

(i) 0.0025 **(ii)** 208 **(iii)** 5005 **(iv)** 126,000

(v) 500.0 **(vi)** 2.0034

1.20 **Round up the following upto three significant figures :**

(i) 34.216 **(ii)** 10.4107 **(iii)** 0.04597 **(iv)** 2808

1.21 **The following data are obtained when dinitrogen and dioxygen react together to form different compounds :**

	Mass of dinitrogen	Mass of dioxygen
(i)	14 g	16 g
(ii)	14 g	32 g
(iii)	28 g	32 g
(iv)	28 g	80 g

(a) Which law of chemical combinations is obeyed by the above experimental data ? Give its statement.

(b) Fill in the blanks in the following conversions.

(i) 1 km = ... mm = ... pm

(ii) 1 mg = ... kg = ... ng

(iii) 1 mL = ... L = ... dm^3

1.22 If the speed of light is 3.0×10^8 m s^{-1}, calculate the distance covered by light in 2.00 ns.

1.23 In a reaction $A + B_2 \rightarrow AB_2$. Identify the limiting reagent, if any, in the following reaction mixtures.

(i) 300 atoms of A + 200 molecules of B

(ii) 2 mol A + 3 mol B

(iii) 100 atoms of A + 100 molecules of B

(iv) 5 mol A + 2.5 mol B

(v) 2.5 mol A + 5 mol B

1.24 Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation : $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$. [At. wt. : N = 14, H = 1]

(i) Calculate the mass of ammonia produced if 2.00×10^3 g dinitrogen reacts with 1.00×10^3 g of dihydrogen.

(ii) Will any of the two reactants remain unreacted ?

(iii) If yes, which one and what would be the mass ?

1.25 How are 0.50 mol Na_2CO_3 and 0.50 M Na_2CO_3 different ?

1.26 If ten volumes of dihydrogen gas reacts with five volumes of dioxygen gas, how many volumes of water vapour would be produced ?

1.27 Convert the following into basic units :

(i) 28.7 pm (ii) 15.15 pm (iii) 25365 mg

1.28 Which one of the following will have largest number of atoms ?

(i) 1 g Au(s) (ii) 1 g Na(s)

(iii) 1 g Li(s) (iv) 1 g of $Cl_2(g)$

[At. Wt. Au = 197, Na = 23, Li = 7, Cl = 35.5]

1.29 Calculate the molarity of a solution of ethanol in water in which the mole fraction of ethanol is 0.040 (assume the density of water to be one).

1.30 What will be the mass of one ^{12}C atom in g ?

1.31 How many significant figures should be present in the answer of the following calculations ?

(i) $\dfrac{0.02856 \times 298.15 \times 0.112}{0.5785}$ (ii) 5×5.364

(iii) 0.0125 + 0.7864 + 0.0215

1.32 Use the data given in the following table to calculate the molar mass of naturally occurring argon isotopes :

Isotope	Isotopic molar mass	Abundance
^{36}Ar	35.96755 g mol^{-1}	0.337%
^{38}Ar	37.96272 g mol^{-1}	0.063%
^{40}Ar	39.9624 g mol^{-1}	99.600%

1.33 Calculate the number of atoms in each of the following :

(i) 52 moles of Ar (ii) 52 u of He (iii) 52 g of He

1.34 A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at STP) of this welding gas is found to weigh 11.6 g. Calculate (i) empirical formula, (ii) molar mass of the gas, and (iii) molecular formula.

1.35 If 20.0 g of $CaCO_3$ is treated with 20.0 g of HCl, how many grams of CO_2 can be generated according to the following equation :

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

Given : Molar mass of ($CaCO_3$ = 100 g mol^{-1}, HCl = 36.5 g mol^{-1}, CO_2 = 44.0 g mol^{-1})

1.36 **Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid according to the reaction :**

$4HCl(aq) + MnO_2(s) \rightarrow 2H_2O(l) + MnCl_2(aq) + Cl_2(g)$

How many grams of HCl react with 5.0 g of manganese dioxide ?

[At. wt. : Cl = 35.5, Mn = 54.94]

ANSWERS

1.1 (i) 18.02 u (ii) 44.011 u (iii) 16.043 u

1.2 mass % of sodium = 32.38, mass % of sulphur = 22.56, mass % of oxygen = 45.333

1.3 Fe_2O_3

1.4 (i) 44 g of CO_2 (ii) 22 g (iii) 22 g

1.5 15.379 g

1.6 15.425 M

1.7 39.834 g

1.8 Fe_2O_3

1.9 35.4527 g mol^{-1}

1.10 (i) 6 moles (ii) 18 moles (iii) 18.066×10^{22} ethane molecules

1.11 0.0292 mol L^{-1}

1.12 25.22 mL

1.13 101332 pascal

1.14 kg

1.15 (i) 10^{-6} (ii) 10 (iii) 10^6 (iv) 10^9 (v) 10^{-15}

1.17 $\sim 15 \times 10^{-4}$ g, 1.25×10^{-4} m

1.18 (i) 4.8×10^{-3} (ii) 2.34×10^5 (iii) 8.008×10^3 (iv) 5.000×10^2 (v) 6.0012

1.19 (i) 2 (ii) 3 (iii) 4 (iv) 3 (v) 4 (vi) 5

1.20 (i) 34.2 (ii) 10.4 (iii) 0.0460 (iv) 2810

1.21 (a) law of multiple proportion (b) (i) (10^6 mm, 10^{15} pm) (ii) (10^{-6} kg, 10^6 ng) (iii) (10^{-3}L, 10^{-3} dm^3)

1.22 6.00×10^{-1} m = 0.600 m

1.23 (i) B is limiting (ii) A is limiting (iii) Stoichiometric mixture -No (iv) B is limiting (v) A is limiting

1.24 (i) 2.43×10^3 g (ii) Yes (iii) Hydrogen will remain unreacted; 5.72×10^2 g

1.25 They are different in terms of unit

1.26 Ten volumes

1.27 (i) 2.87×10^{-11}m (ii) 1.515×10^{-11} m (iii) 2.5365×10^{-2} kg

1.28 1 g of Li has the largest number of atoms

1.29 2.222 M

1.30 1.99265×10^{-23} g

1.31 (i) 3 (ii) 4 (iii) 4

1.32 39.948 g mol^{-1}

1.33 (i) 3.131×10^{25} atoms (ii) 13 atoms (iii) 7.8286×10^{24} atoms

1.34 Empirical formula CH, molar mass 26.0 g mol^{-1}, molecular formula C_2H_2

1.35 0.94 g $CaCO_3$

1.36 8.40 g HCl

ADDITIONAL QUESTIONS AND PROBLEMS

Q. Volume of a solution changes with change in temperature, then, will the molality of the solution be affected by temperature ? Give reason for your answer.

A. No, molality does not change with temperature because it involves masses of the solute and solvent which do not change with temperature.

Q. Why molality is preferred over molarity in expressing the concentration of a solution ?

A. Molality is the number of moles of the solute present in 1 kg of the solvent whereas molarity is the number of moles of the solute present per litre of the solution. Thus, molality involves only masses which do not change with temperature whereas molarity involves volume which changes with temperature. Hence, molality is preferred over molarity.

Q. Find out the molarity of solution containing 20% NaOH.

A. 5

Q. If the solution contains 4.9% H_2SO_4 by weight (d = 1.02 g/ml), find out the molarity of the solution.

A. 0.51

Q. How many grams of $KMnO_4$ must be added to 35 ml of 0.1 (M) $KMnO_4$ solution to make it 0.6(M).

A. 2.765 g

Q. The density of an aqueous solution of NaCl containing 8.0 g of salt per 100 g of solution is 1.05 g/ml at 25^0C. Calculate the molarity of the solution.

A. 1.43[M]

Q. Find out the strength of the resulting solution when 250 cc. N/10 HCl, 100 c.c. N/5 H_2SO_4 and 650 c.c. N/10 HNO_3 are mixed together.

A. 0.11 (N)

Q. 1.08 g of Cu wire was reacted with HNO_3 and resulting solution was dried and ignited when 1.35 g of copper oxide was obtained. In another experienct, 2.3 g of copper oxide when heated in the flow of H_2 gives 1.84 g of Cu. Show that these data are in accordance with the law of constant composition.

A. % of Cu = 80% and % of oxygen = 20%

Q. The volume of a drop of rain water is 0.448 ml. How many molecules of H_2O and number of H-atoms in the drop ?

A. 1.5×10^{22} and 3.0×10^{22}

Q. Chlorophyll, the green colouring matter of plants responsible for photosynthesis, contains 2.68% of magnesium by mass. Calculate the number of magnesium atoms in 2.00 g of chlorophyll.

Q. If the elemental composition of butyric acid is found to be 54.2% C, 9.2% H and 36.6 % O, determine the empirical formula.

Q. How much 1.00 M HCl should be mixed with what volume of 0.250 M HCl in order to prepare 1.00 L of 0.500 M HCl ?

A. 333 mL of 1m HCl

Q. Calculate the final concentration of HNO_3 if 0.20 mol HNO_3 is added to a beaker containing 2.0 L of 1.1 M HNO_3 and enough pure water is added to give a final volume of 3.0 L.

A. 0.80 m

Q. If 40.00 mL of 1.600 M HCl and 60.00 mL of 2.000 M NaOH are mixed, what are the molar concentrations of Na^+, Cl^-, and OH^- in the resulting solution ? Assume a total volume of 100.00 mL.

A. 0.56 M OH^-, 0.640 Cl^-, 1.2MNa^+

Q. Calculate the molarity of the original H_3PO_4 solution if 20.0 mL of H_3PO_4 solution is required to completely neutralize 40.0 mL of 0.0500 M $Ba(OH)_2$ solution.

A. 0.0667 M

Q. **The acidic substance in vinegar is acetic acid CH_3COOH. When 6.00 g of a certain vinegar was titrated with 0.100 M NaOH, 40.11 mL of base had to be added to reach the equivalence point. What percent by mass of this sample of vinegar is acetic acid ?**

A. 4.01 %

Q. **Calculate the percent of BaO in 29.0 g of a mixture of BaO and CaO which just reacts with 100.8 mL of 6.00 M HCl. $BaO + 2HCl \rightarrow BaCl_2 + H_2O$; $CaO + 2HCl \rightarrow CaCl_2 + H_2O$**

A. 65.5 % BaO

Q. **What volumes of 12.0 N and 3.00 N HCl must be mixed to give 1.00 L of 6.00 N HCl ?**

A. 0.33 L, 0.667 L

Q. **One gram of a mixture of $CaCO_3$ and $MgCO_3$ gives 240 ml of CO_2 at N.T.P. Calculate the percentage composition of mixture (Ca = 40, Mg = 24, C = 12, O = 16).**

A. $CaCO_3$ = 62.5 %, $MgCO_3$ = 37.5%

Q. **The density of a 2.0 M solution of acetic acid (MW = 60) in water is 1.02g/mL. Calculate the mole fraction of acetic acid.**

A. 0.038

Q. **The density of a 2.03 M solution of acetic acid in water is 1.017 g/mL. Calculate the molality of the solution.**

A. 2.27 m

Q. **In an Industrial process for producing acetic acid, oxygen gas is bubbled into acetaldehyde CH_3CHO, containing manganese (II) acetate (catalyst) under pressure at 60^0C.**

$2CH_3CHO(l) + O_2(g) \rightarrow 2CH_3COOH(l)$

In a laboratory test of this reaction, 20.0 g CH_3CHO and 10.0 g O_2 were put into a reaction vessel. (a) How many grams of acetic acid can be produced by this reaction from these amounts of reactants ? (b) How many grams of the excess reactant remaining after the reaction is complete ?

A. (a) 27.24 gm (b) 2.73 gm

Q. **A plant virus is found to consist of uniform cylindrical particles of 150 Å in diameter and 5000 Å long. The specific volume of the virus is 0.75 cm^3/g. If the virus is considered to be a single particle, find its molar mass.**

A. 70.96×10^6 g/mol