Basic Concepts Of Chemistry

Measurement in Chemistry

1. **Significant figures**. The total number of **digits** in a number including the last digit whose value is uncertain is called the number of significant figure. e.g., 12.4267 ± 0.0001 has six significant figures.

(a) Rules for assigning significant figures

(i) All non–zero digits as well as the zeros between the non – zero digits are significant e.g. 6003 has four significant figures.

(ii) Zeros to the left of the first non–zero digit in a number are not significant e.g. 0.0045 has two significant figures.

(iii) If a number ends in zeros but these zeros are to the right of the decimal point, then these zeros are significant e.g. 2.600 has four significant figures.

(iv) If a number ends in zeros but these zeros are not to the right of a decimal point, these zeros may or may not be significant e.g.10600 written as 1.06×10^4 has three significant figures but when written as 1.060×10^4 has four significant figures and so on.

Calculation involving significant figure

(i) The result on an addition or subtraction should be reported to the same number of decimal places as that of the term with least number of decimal places e.g. sum 4.528 + 6.24 = 10.768 will be reported as 10.77.

(ii) The result of multiplication or division should be reported to the same number of significant figures as is possessed by the least precise term e.g. 4.324 × 2.8 = 12.1072 will be reported as 12.

(iii) If calculation involves a number of steps, the result should contain the same number of significant figures as that of the least precise number involved, other than the exact numbers. e.g. $\frac{4.28 \times 0.146 \times 3}{0.0418} = 44.84784$ should be reported as 44.8.

(b) Rounding off. It implies as follows :

(i) If the digit just next to the last digit to be retained is less than 5, the last digit is taken as such and all other digits of its right are dropped e.g. 1.234 = 1.23.

(ii) If the digit just next to the last digit to be retained is greater than 5, the digit to be retained is increased by 1and all other digits on its right are dropped e.g. 1.2366 = 1.24

(iii) If the digit just next to the last digit to be retained is equal to 5, the last significant figure is left unchanged if it is even and is increased by 1 if it is odd e.g. 1.235 = 1.24, 1.225 = 1.22.

(c) Scientific Notation. It is convenient to represent very small and very large numbers in exponential forms, which is called scientific notation. It is written as $N \times 10^x$, where N is a number with a single non zero digit to the left of the decimal point and x is an integer. N represents the number of significant figures.
For instance, a very small number (say 0.00000461) as per scientific notation can be written as 4.61 $\times 10^{-6}$ and it has three significant figures. Similarly, a very big number (say 760000000000) can be written as 7.6×10^{11} with only two significant figures.
Some solved examples : 2.500 has four significant figures; 94.5 has three ; 0.52 has two; 0.006 has one 2.07 has three 0.0034 has two ; 2.0 has two; 2.50 has three ; 0.0200 has three ; 6.02×10^{23} has three ; 0.005 has one significant figure and 0.001864 can be rounded of to 0.00186. Similarly 0.000650 can be written as 6.50×10^{-4}.

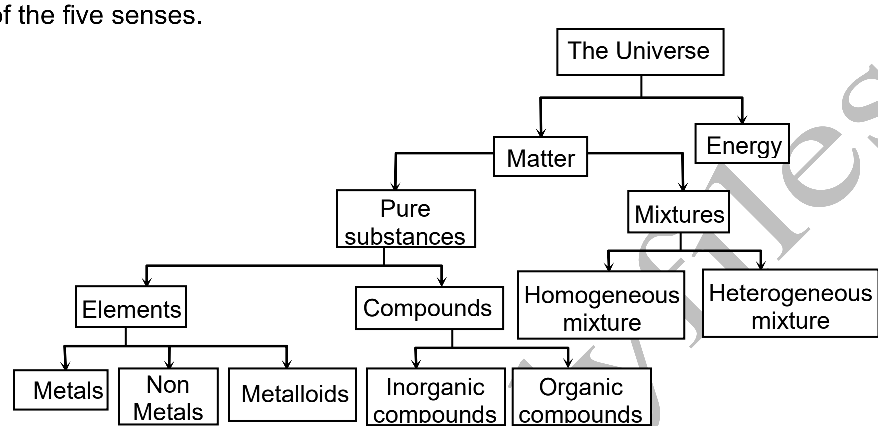
2. **SI units.** The SI unit of some common physical quantities are :
length = meter (m). Mas = kilogram (kg), Time = second (s), Temperature = Kelvin (K), Electric current = ampere (A), Amount of substance = moles (mol), Volume = m^3, Density = kgm^{-3}, Velocity = ms^{-1}, Force = Newton (N), Pressure = Pascal (Pa) or Nm^{-2}, Work/Energy = Joule(J).

3. **Some useful conversion factors.** 1Å = 10^{-10} m, 1nm = 10^{-9}m, 1pm = 10^{-12}m,
1 litre = $10^{-3}m^3$ = 1 dm^3, 1 atm = 760 mm or torr = 101.325 pa or Nm^{-2},
1 bar = 10^5 Nm^{-2} = 10^5 pa, 1 calorie = 4.184 J, 1 electron volt (eV) = 1.6022×10^{-19} J.

4. **Units with special names**

Quantity		S.I. unit	Symbol and definition
1.	Force	newton	N = kg ms^{-2} = Jm^{-1}. One newton is equal to the force required to hold 1 kg mass against gravity.
2.	Pressure	pascal	pa = newton per square metre Nm^{-2}. One pascal is exerted when one newton force is applied on an area of 1 m^2. pa = Nm^{-2} = kg m^{-1} s^{-2}
3.	Energy	joule	J = Nm = kg m^2s^{-2}. One joule energy is required to exert against a force of one newton through one metre
4.	Power	Watt	W (kg m^2 s^{-3} or Nm)
5.	Electric charge	coulomb	C
6.	Electric potential difference	Volt	V ($JA^{-1}s^{-1}$ or kg m^2s^{-3} A^{-1})
7.	Frequency	Hertz	Hz(s^{-1})

5. **Matter**. Matter is defined as anything, which occupies space, possesses mass and can be judged by one or more of the five senses.



Matter can be classified on the basis of chemical composition as under:
(a) Elements (b) Compounds (c) Mixtures

6. **Elements.** An elements is pure substance which can neither be broken nor built from simple substances by any known physical or chemical methods. So an element is the simplest form of matter which is made up of one type of atoms.

7. **Compounds.** A product obtained by the combinations of two or more elements in a fixed ratio by weight having different properties from the combining elements is known as a **compound**. The compounds can be decomposed back into the same elements from which it is formed by suitable chemical methods. The properties of the compound do not change e.g. water, carbon dioxide.

8. **Mixture.** A product obtained by mixing two or more substances (elements or compounds) so that the properties of the constituents remain unchanged is known as a mixture.

Types of Mixture : Two types of mixture are there
(i) Homogeneous mixture
(ii) Heterogeneous mixture

Homogeneous mixture: This type of mixture has similar composition throughout. Such homogeneous mixtures are called solutions. Air, petrol, an alloy are examples of homogeneous mixture.

Heterogeneous mixture : These type of mixtures does not have similar composition through out e.g., smoke.

Solved Problems

Illustration 1 **Match the following prefixes with their multiples:**

Prefixes	Multiples
(a) micro	10^{6}
(b) deca	10^{9}
(c) mega	10^{-6}
(d) giga	10^{-15}
(e) femto	10

micro = 10^{-6} , deca = 10 , mega = 10^{6} , giga = 10^{9} , femto = 10^{-15}

Illustration 2 **Convert the following into basic units:**

(a) 28.7 pm (b) 15.15 μs (c) 25365 mg

(a) 28.7 pm = 28.7 pm x $\frac{10^{-12}\text{m}}{1\,\text{pm}} = 2.87 \times 10^{-11}\text{m}$

(b) 15.15 μ s = 15.15 μ s x $\frac{10^{-6}\text{s}}{1\,\mu\,\text{s}} = 1.515 \times 10^{-5}\text{s}$

(c) 25365 mg = 25365 mg x $\frac{1\,\text{g}}{1000\text{ mg}} \times \frac{1\,\text{kg}}{1000\text{ g}} = 2.5365 \times 10^{-2}\text{ kg}$

Illustration 3 **Express the following in the scientific notation:**

(a) 0.0048 (b) 234,000 (c) 8008 (d) 500.0 (e) 6.0012

(a) 4.8 x 10^{-3} (b) 2.34 x 10^{5} (c) 8.008 x 10^{3} (d) 5.000 x 10^{2} (e) 6.0012 x 10^{0}

Illustration 4 **How many significant figures are present in the following?**

(a) **0.0025 (b) 208 (c) 5005 (d) 126,000 (e) 500.0 (f) 2.0034 (g) π**

(a) 2, (b) 3, (c) 4, (d) 6, (e) 4, (f) 5 (g) infinite

Illustration 5 **Round up the following upto three significant figures:**

(a) 34.216 (b) 10.4107 (c) 0.04597 (d) 280.8

(a) 34.2 (b) 10.4 (c) 0.0460 (d) 281

Illustration 6 **How many significant figures should be present in the answer of the following calculations?**

(a) $\frac{0.02856 \times 298.15 \times 0.112}{0.5785}$ **(b) 5 x 5.354 (c) 0.0125 + 0.7864 + 0.0215**

(a) The least precise term 0.112 has 3 significant figures. Hence, answer is 3

(b) 5 is an exact number. The other term 5.364 has 4 significant figures. Hence, answer is 4.

(c) Each term has 3 significant figures. Answer is 3.

Laws of Chemical combination

1. Laws of conservation of mass

The law was first studied by Lavosier in 1774. According to this law, "in all chemical changes the total mass of the system remains constant or in a chemical change mass can neither be created nor destroyed". This means when a chemical reaction takes place, the total mass of all the reaction products is exactly equal to the total mass of the reactants from which the products were obtained.

Limitation : This law is not valid in nuclear phenomenon. In nuclear phenomenon total mass of the reactant is not equal to total mass of the product and mass defect is observed. So according to Einstein this law is better known as law of conservation of mass and energy.

2. Law of definite or constant proportion

This law was given by Proust in 1979. According to this law "Whatever may be the source of a compound, the constituent elements of the compound are always in a same constant proportion.

Example : CO_2 is obtained –

(a) by burning carbon, $C + O_2 \longrightarrow CO_2$

(b) by heating $NaHCO_3$, $2NaHCO_3 \xrightarrow{\text{heat}} Na_2CO_3 + CO_2 + H_2O$

(c) by heating $CaCO_3$, $CaCO_3 \xrightarrow{\text{heat}} CaO + CO_2$

(d) by action of acids on lime stone, $CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$ etc.

but in CO_2, C : O (by weight) is always 12 : 32

Limitation of Law of definite proportion

(i) The law is not applicable if an element exists in different isotopes which maybe 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.

(ii) The elements may combine in the same ratio but the compounds formed may be different. For example, in the compounds, 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.

3. Law of multiple proportions

This law was put forward by Dalton in 1803. According to this law if two elements combine to form more than one compound, then the weights of one which combine with a fixed weight of the other bear a simple ratio to one another. For example, hydrogen and oxygen combine to form two compounds H_2O (Water) and H_2O_2 (Hydrogen peroxide)

In water, hydrogen 2 parts and oxygen 16 parts. In hydrogen peroxide, hydrogen 2 parts and oxygen 32 parts combine. The weight of oxygen which combine with same weight of hydrogen in these two compounds bear a **simple ratio 1 : 2.**

Nitrogen forms five stable oxides.

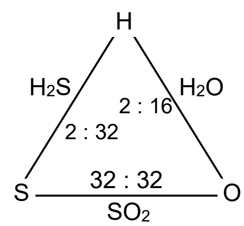
N_2O	Nitrogen 28 parts	oxygen 16 parts
NO	Nitrogen 14 parts	oxygen 16 parts
N_2O_3	Nitrogen 28 parts	oxygen 48 parts
N_2O_4	Nitrogen 28 parts	oxygen 64 parts
N_2O_5	Nitrogen 28 parts	oxygen 80 parts

The weights of oxygen which combine with same weight of nitrogen in the five compounds bear a ratio 16 : 32 : 48 : 64 : 80 or 1 : 2 : 3 : 4 : 5.

Limitations. If two elements form several compounds of varying composition the law of multiple proportion does not holds. For example titanium forms several oxides of varying composition such as $TiO_{1.46-1.56}$ and $TiO_{1.9-2.0}$.

4. Law of reciprocal proportion

This law was given by Richter in 1792. According to this law, "If two elements separately combine with a fixed weight of a third element, the ratio in which two elements will combine together, will be same or its simple multiple in which they combine with fixed weight of third element."



Example: Sulphur and Oxygen combine with Hydrogen to form H_2S and H_2O separately.

In H_2S H : S = 2 : 32

In H_2O H : O = 2 : 16

Also Sulphur and Oxygen combines to form SO_2 in which S : O = 32 : 32, which is a simple multiple of 32 : 16.

5. Law of combining volumes (Gay-Lussac's chemical law)

This law was enunciated by **Gay-Lussac** in 1808. According to this law, gases react with each other in the simple ratio of their volumes and if the product is also in gaseous state, the volume of the product also bears a simple ratio with the volumes of gaseous reactants when all volumes are measured under similar conditions of temperature and pressure.

$$\underset{1\,vol}{H_2} + \underset{1\,vol}{Cl_2} \rightarrow \underset{2\,vol}{2HCl} \quad \text{ratio} \quad 1:1:2$$

$$\underset{1\,vol}{2H_2} + \underset{1\,vol}{O_2} \rightarrow \underset{2\,vol}{2H_2O} \quad \text{ratio} \quad 2:1:2$$

$$\underset{2\,vol}{2CO} + \underset{1\,vol}{O_2} \rightarrow \underset{2\,vol}{2CO_2} \quad \text{ratio} \quad 2:1:2$$

6. **Avogadro's Law.** Equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules or equal number of moles.

Solved Problems

Illustration 7 **Two aqueous solutions X and Y containing 0.585 g of sodium chloride and 1.70 g of silver nitrate respectively were mixed. The weight of silver chloride formed was found to be 1.432 g and 0.853g $NaNO_3$ were obtained. Show that these observation confirms the law of chemical conservation of mass.**

The chemical reaction, taking place between NaCl and $AgNO_3$ in aqueous solution is as

$NaCl + AgNO_3 \rightarrow AgCl + NaNO_3$

Wt. of sodium chloride + wt. of silver nitrate before experiment = 0.585 + 1.70 = 2.285 g

Wt. of silver chloride + wt. of sodium nitrate after experiment = 1.432 + 0.853 = 2.285 g

since the masses of reactants before the reaction are equal to the masses of products after the reaction, the above data confirm the law of conservation of mass.

Illustration 8 **Hydrogen sulphide (H_2S) contains 94.11% sulphur, water (H_2O) contains 11.11% hydrogen and sulphur dioxide (SO_2) contains 50% oxygen. Show that the results are in agreement with the law of reciprocal proportions.**

In water, the weight of hydrogen = 11.11 g

The weight of oxygen = 100 – 11.11 = 88.89

In sulphur dioxide, the weight of sulphur = 50 g

The weight of oxygen = 100 –50 = 50 g

$\therefore$ The weight of sulphur that combines with 88.89 g of oxygen = $\frac{50}{50} \times 88.89 = 88.89$ g

$\therefore$ The ratio between the weight of sulphur and hydrogen which combine with a fixed weight of oxygen (88.89 g) is 88.99 : 11.11 or 8 : 1 …(i)

In hydrogen sulphide, the weight of sulphur = 94.11 g

The weight of hydrogen =100 – 94.11 = 5.89 g

$\therefore$ The ratio between weights of sulphur and hydrogen is 94.11:5.89 or 16: 1 …(ii)

The two ratios (i) and (ii) are related as $\frac{8}{1} : \frac{16}{1}$ or 1 : 2

Which are simple multiples of each other. Therefore, the law of reciprocal proportions holds good.

Illustration 9 **6.488 g of lead combine directly with 1.002 g of oxygen to form lead peroxide (PbO_2). Lead peroxide is also produced by heating lead nitrate and it was found that the percentage of oxygen present in lead peroxide is 13.38 percent. Use these data to illustrate the law of constant composition.**

Step 1. To calculate the percentage of oxygen in first experiment.

Weight of peroxide formed = 6.488+ 1.002 = 7.490 g.

7.490 g lead peroxide contain 1.002 g of oxygen 100 g of lead peroxide will contain oxygen

$= \frac{1.002}{7.490} \times 100 = 13.38$ g i.e. oxygen present = 13.38%

Step 2. To compare the percentage of oxygen in both the experiments.

Percentage of oxygen in PbO_2 in the first experiment = 13.38

Percentage of oxygen in PbO_2 in the second experiment = 13.38

Since the percentage composition of oxygen in both the samples of PbO_2 is identical, the above data illustrate the law of constant composition.

Illustration 10 **Carbon is found to form two oxides, which contain 42.9% and 27.3% of carbon respectively. Show that these figures illustrate the law of multiple proportions.**

Step 1. To calculate the percentage composition of carbon and oxygen in each of the two oxides

	First oxide	Second oxide	
Carbon	42.9%	27.3%	(Given)

Oxygen(by difference) 57.1% 72.7%

Step 2. To calculate the weights of carbon which combine with a fixed weight i.e., one part by weight of oxygen in each of the two oxides.

In the first oxide, 57.1 parts by weight of oxygen combine with carbon = 42.9 parts.

∴ 1 part by weight of oxygen will combine with carbon

$$= \frac{42.9}{57.1} = 0.751 \cdot$$

In the second oxide. 72.7 parts by weight of oxygen combine with carbon = 27.3 parts.

∴ 1 parts by weight of oxygen will combine with carbon

$$= \frac{27.3}{72.7} = 0.376$$

Step 3. To compare the weights of carbon which combine with the same weight of oxygen in both the oxides.

The ratio of the weights of carbon that combine with the same weight of oxygen (1 part) is

0.751 : 0.376 or 2 : 1

Since this is a simple whole number ratio, so the above data illustrate the law of multiple proportions.

Unsolved Problem

1. What mass of silver nitrate will react with 5.85 g of sodium chloride to produce 14.35 g of silver chloride and 8.5 g of sodium nitrate, if the law of conservation of mass is true ?

2. 2.16 g of copper metal when treated with nitric acid followed by ignition of the nitrate gave 2.70 g of copper oxide. In another experiment 1.15 g of copper oxide. In another experiment 1.15 g of copper oxide upon reduction with hydrogen gave 0.92 g of copper. Show that the above data illustrate the Law of Definite Proportions.

3. Two oxides of nitrogen contain the following percentage compositions:
(i) Oxide A contains 63.64% nitrogen and 36.36% oxygen
(ii) Oxide B contains 46.67% nitrogen and 53.33% oxygen
Establish the law of multiple proportion.

4. 61.8 g of A combine with 80 g of B. 30.9 g of A combine with 106.5 g of C, B and C combine to form compound CB_2. Atomic weights of C and B are respectively 35.5 and 6.6. Show that the law of reciprocal proportions is obeyed.

5. (i) 10g of lead on heating gave 10.75 g of Litharage, PbO.
(ii) 9.775g of red lead (Pb_3O_4) yielded on strong heating 9.545 g of Lithrage
(iii) 4.87 g of lead peroxide [PbO_2] gave on heating 4.545 g of Litharge.
Show that there results illustrate law of multiple proportion.

DALTON'S ATOMIC THEORY

John Dalton in the year 1808 gave a theory known as Dalton's atomic theory, which is a land mark in the history of chemistry. The main features of Dalton's atomic theory are :

(i) Elements consist of minute, indivisible, indestructible particles called atoms.

(ii) Atoms of an element are similar to each other as regards their mass and size.

(iii) Atoms of different elements differ in properties and have different masses and sizes.

(iv) Compounds are formed when atoms of different elements combine with each other in simple numerical ratio.

(v) Atoms can neither be created nor be destroyed or transformed into atoms of other elements.

(vi) The relative numbers and kind of atoms are always the same in a given compound.

This theory convincingly explained the various laws of chemical combination, but the theory has undergone a complete shake up with the modern concept of structure of atom. The following are the **modified views** regarding Dalton's atomic theory.

(i) The atom is no longer supposed to be indivisible. The atom is not a simple particle but a complex one.

(ii) Atoms of the same element may not necessarily possess the same mass but possess the same atomic number and show similar chemical properties, this led to the discovery of isotopes

(iii) Atoms of one element can be converted into atoms of other element hence artificial radioactivity was discovered.

(iv) Atoms of different elements may possess the same mass but they always have different atomic numbers and differ in chemical properties and these type of elements are called as isobars.

(v) In certain organic compounds, like proteins, starch, cellulose etc., the ratio in which atoms of different elements combine cannot be regarded as simple one.

Atom :

It is the smallest particle of an element, which takes part in a chemical reaction. It may or may not have independent existence e.g., atom of hydrogen (H) and atom of oxygen(O).

Molecule :

The smallest particle of a substance (element or compound) which is capable of independent existence is called a molecule e.g., H_2 , O_2, N_2, H_2O.

Atomic mass :

The atomic mass of an element may be defined as the average relative mass of an atom of the element as compared to the mass of an atom of carbon ${}_6C^{12}$ taken as 12 a.m.u.

The atomic mass has no units. It simply tells how many times an atom of the element is heavier than $\frac{1}{12}$th of an atom of carbon. 1 a.m.u. (atomic mass unit) = 1.66×10^{-24} g. It is called 1 Avogram or 1 aston. Lightest element is hydrogen and heaviest naturally occurring atom is U – 238.

Gram atomic mass :

It is the atomic mass of an element expressed in grams. This amount of the element is called one gram atom e.g., 1 gm atom of carbon means 12 gm of carbon and 1 gm atom of Helium means 4 gm.

Calculation of average atomic mass :

If an element exists in two isotopes having atomic masses 'a' and 'b' in the ratio of m : n, then average atomic mass = $\frac{m \times a + n \times b}{m + n}$.

For example, natural boron is a mixture of 20% (B –10) and 80% (B – 11).

Hence its atomic mass = $\frac{20 \times 10 + 80 \times 11}{100} = 10.8$ a.m.u.

Atomic mass of carbon is 12.011 a.m.u.

Methods of determining atomic masses.

(i) Atomic weight = eq. wt. $\times$ valency

(ii) Dulong and Petit's method. For solid elements (except Be, B, C and Si).

At. wt. $\times$ specific heat = 6.4

$\therefore$ Approx. atomic wt. = $\frac{6.4}{\text{Sp. heat}}$

Valency = $\frac{\text{Approx.atomic wt.}}{\text{eq. wt}}$

Now, Exact atomic wt. = Eq. wt. $\times$ Valency

(iii) From ratio of heat capacities, for gases. We have

Ratio $\gamma = C_p/C_v$	1.66	1.40	1.33	1.26
Atomicity	1	2	3	4

$\therefore$ Atomic mass of the gaseous element = $\frac{\text{Mol mass}}{\text{Atomicity}}$

Where, Molecular mass can be found from vapour density (M. W. = 2 $\times$ V.D.)

(iv) Volatile chloride method (or vapour density method for elements forming volatile chlorides)

Molecular weight of the chloride = 2 $\times$ V.D.

If x is the valency of the element (M), then formula of its chloride will be MCl_x. Hence Mol. wt. of the chloride MCl_x = At. wt. of M + 35.5x

$2 \times$ V.D. = Eq. wt. of M $\times$ valency + 35.5x

$2 \times$ V.D. = E $\times$ x + 35.5x = x(E + 35.5)

$\therefore \quad x = \dfrac{2 \times V.D}{E + 35.5}$

Knowing the equivalent wt. E of the element, the value of x can be calculated.
Then Atomic wt. = Eq. wt. × Valency

(v) From percentage of an element in a given compound. For example, potassium chromate (K_2CrO_4) is found to contain 26.78% of Cr. Its atomic wt. can be calculated as follows :
Suppose atomic wt. of Cr = a

$\therefore \quad$ Mol. wt. of $K_2CrO_4 = 2 \times 39 + a + 4 \times 16 = 142 + a$

$\therefore \quad$ % of Cr in $K_2CrO_4 = \dfrac{a \times 100}{142 + a} = 26.78$

On solving $\quad a = 52.0$

Molecular Mass or Weight

It is relative weight of one molecule of a compound to the weight of C-12. It can be calculated either.

(i) By adding up the atomic weights of all the constituent atoms present in one molecule

(ii) By vapour density of a gas which is related to its molecular weight as under :
Molecular weight = 2 × Vapour density.

Gram Molecular Mass

When the molecular mass of a substance on a.m.u. scale is expressed in grams, this is referred to as gram molecular mass (G.M.M.) or one gram molecule.

Equivalent weight

Equivalent weight of an element or compound is the number of parts by weight of it, which combines with or displaces from directly or indirectly, 1.008 parts by weight of hydrogen or 8 parts by weight of oxygen or 35.5 parts by weight of chlorine or 108 parts by weight of Ag.

Some important relations for calculating equivalent mass are given below :

(i) Equivalent mass of an element = $\dfrac{\text{Atomic mass}}{\text{Valency}}$.

(ii) Equivalent mass of a acid = $\dfrac{\text{Molecular mass}}{\text{Basicity}}$

[Basicity is number of replacable H^+ ion in an acid]

(iii) Equivalent mass of base = $\dfrac{\text{Molecular mass}}{\text{acidity}}$

[acidity is number of replacable OH^- ion in a base]

(iv) Equivalent mass of simple salts = $\dfrac{\text{Molecular mass}}{\text{Total + ve charge on cation}}$

(v) Equivalent mass of oxidising or reducing agents = $\dfrac{\text{Molecular mass}}{\text{Number of electron change}}$

(vi) Equivalent weight of an ion = $\dfrac{\text{Formula weight}}{\text{charge on ion}}$

Gram Equivalent Weight

Gram equivalent weight of an element or compound is its equivalent weight expressed in grams.

Determination of equivalent weight

(i) **Hydrogen displacement method.** Calculate the weight of the metal, which displaces 1.008 parts by weight of hydrogen. It is called eq. wt of metal.

(ii) **Oxide formation or reduction of the oxide method.** Calculate the weight of the metal which combines with or displaces 8 parts by weight of oxygen.

(iii) **Chloride formation method.** Calculate the weight of the element which combines with or displaces 35.5 parts by weight of chlorine.

(iv) **Metal displacement method.** When a more electropositive metal is added to displace the less electropositive metal from its salt.

Then $\frac{\text{Weight of metal added}}{\text{Weight of metal displaced}} = \frac{\text{Eq.wt. of metal added}}{\text{Eq. wt. of metal dispaced}}$

(v) **Double decomposition method.** For a reaction of the type

$AB + CD \longrightarrow AD\downarrow + BC$

So we can calculate equivalent weight of metal added or displaced if others are given.

$$\frac{\text{Mass of AB taken}}{\text{Mass of AD formed}} = \frac{\text{Eq.wt.of AB}}{\text{Eq.wt.of AD}} = \frac{\text{Eq. wt. of A} + \text{Eq. wt.of B}}{\text{Eq. wt. of A} + \text{Eq. wt. of D}}$$

(vi) **Electrolytic method (Faraday' second law).** When the same quantity of electricity flows through solutions of different electrolytes, then,

$$\frac{\text{Mass of X deposited}}{\text{Mass of Y deposited}} = \frac{\text{Eq. wt. of X}}{\text{Eq. wt. of Y}}$$

(vii) **Eq. wt. of an element is the wt.** which is deposited by 96500 coulombs of charge.

(viii) **Silver salt method (for organic acids only).** A known weight of silver salt of the organic acid is ignited to give a residue as Ag, then

$$\frac{\text{Eq. wt. of RCOOAg}}{\text{Eq.wt.of Ag(108)}} = \frac{\text{Weight of silver salt}}{\text{Weight of silver}}$$

Eq. wt. of acid (RCOOH) = Eq. wt. of RCOOAg -107

(ix) **Conversion method.** When one compound of a metal is converted into another compound of the same metal(e.g., metal carbonate $\longrightarrow$ metal oxide), then

$$\frac{\text{Weight of compound I}}{\text{Weight of compound II}} = \frac{\text{Eq. wt. of compound I}}{\text{Eq.wt,of compund II}}$$

$$= \frac{\text{Eq. wt. of metal} + \text{Eq,.wt.of anion of compound I}}{\text{Eq. wt,of metal} + \text{Eq..wt of anion of compound II}}$$

An N solution is termed as a **Normal solution** and consists of one-gram equivalent of the solute dissolved per litre of the solution. Similarly, a **decinormal solution** ($\frac{N}{10}$ solution) represents a solution containing 1/10-gram equivalents of the solute per litre of the solution.

Normality of a solution is dependent on the temperature.

Solved Problems

Illustration 11 **The chloride of a metal contains 71% chlorine by weight and the vapour density of it is 50. What is the atomic weight of the metal ?**

Mol. wt. of metal chloride = 50 × 2 =100;

Let metal chloride be MCl_n then

Eq. of metal = Eq. of chloride, or $\frac{29}{E} = \frac{71}{35.5}$

$\therefore \quad E = \frac{29}{2}$;

Now a + 35.5n = 100

or n.E + 35.5n = 100; $\quad\therefore\quad$ n = 2

Therefore, a = 2 × E = 2 × (29/2) = 29.

Illustration 12 **2g of a metal in H_2SO_4 gives 4.51 g of the metal sulphate. The specific heat of metal is 0.057 cal/g. Calculate the valency and atomic weight of metal.**

At. Wt. $\times$ specific heat = 6.4

a $\times$ 0.057 = 6.4

a = 112.28

Now equivalent of metal = Equivalent of metal sulphate

$\frac{2}{E} = \frac{4.51}{E+48}$ ($\because$ Eq. wt. of SO_4^{-2} = 48)

$\therefore$ E = 38.24

$\therefore$ Valency metal = $\frac{\text{At. wt.}}{\text{Eq. wt.}} = \frac{112.28}{38.24}$ = 2.93

≈ 3 ($\because$ Valency is integer)

$\therefore$ Exact at. wt. of metal = Eq. wt. $\times$ valency

= 38.23 $\times$3 = 114.72

Illustration 13 **1.020 g of metallic oxide contains 0.540 g of the metal. Calculate the equivalent mass of the metal and hence its atomic mass with the help of Dulong and Petit's law. Taking the symbol for the metal as M, find the molecular formula of the oxide. The specific heat of the metal is 0.216 cal deg^{-1} g^{-1}.**

Mass of oxygen in the oxide = (1.020 – 0.540) = 0.480 g

Equivalent mass of the metal = $\frac{0.540}{0.480} \times 8 = 9.0$

According to Dulong and Petit's law.

Approx. atomic mass = $\frac{6.4}{\text{Sp.heat}} = \frac{6.4}{0.216} = 29.63$

Valency of the metal = $\frac{\text{At.mass}}{\text{Eq.mass}} = \frac{29.63}{9.0} = 3$.

Hence, the formula of the oxide is M_2O_3.

Illustration 14 **P and Q are two elements, which forms P_2Q_3 and PQ_2. If 0.15 mole of P_2Q_3 weighs 15.9 g and 0.15 mole of PQ_2 weighs 9.3 g, what are atomic weights of P and Q?**

Let at. wt. of P and Q are a and b respectively.

Mol. Wt. of P_2Q_3 = 2a + 3b

Mol. Wt. of PQ_2 = a + 2b

Now given, that 0.15 mole of P_2Q_3 weigh 15.9 g

(2a + 3b) = $\frac{15.9}{0.15}$ $\left(\because \frac{\text{wt.}}{\text{mol.wt.}} = \text{mole}\right)$

Similarly, (a + 2b) = $\frac{9.3}{0.15}$

Solving these two equations

b = 18

a = 26

Unsolved Problems

6. Element X reacts with oxygens to produce a pure sample of X_2O_3, In an experiment it is found that 1.00 g of X produces 1.16 g of X_2O_3. Calculate the atomic weight of X. Given: atomic weight of oxygen, 16.0 g mol^{-1}.

7. 1.60 g of a metal were dissolved in HNO_3 to prepare its nitrate. The nitrate on strong heating give 2 g oxide. What is the equivalent weight of metal ?

8. The oxide of a metal contains 60% of the metal. What will be the percentage of bromine in the bromide of the metal, if the valency of the metal is the same in both the oxide and the bromide.

9. 3.6 g of a metal oxide on reduction with hydrogen left 3.2 g of the metal. The atomic weights of the metal and oxygen are 64 and 16 respectively. What will be the simplest formula of the oxide ?

10. The carbonate of Metal contains 50% Metal. What is equivalent weight of metal ?

Avogadro's Number

It is the number of atoms present in one-gram atom of an element or the number of molecules present in one-gram molecule of the substance. Its value is 6.02×10^{23}. 22.4 litres of a gas at NTP also contains Avogadro's number of molecules.

Mole

A collection of 6.02×10^{23} particles is called mole.

S.I definition of a mol is the amount of the substance which contains as many elementary entities as there are atoms in exactly 12g (0.012 Kg) of carbon ^{12}C isotope.

1 mole of a substance = 6.02×10^{23} particles

1 mole

- Monoatomic Elements eg. Cu, Ag, K, Na etc. — **6.02×10^{23} atoms**
- Covalent compounds & Polyatomic elements eg. CO_2, H_2SO_4, SO_3, NH_3, N_2, S_8, P_4 etc. — **6.02×10^{23} molecules**
- Ionic Compounds eg. NaCl, KCl, $Ca_3(PO_4)_2$ — **6.02×10^{23} formula units**
- Ions (Gaseous state or in aqueous solutions) eq. Ca^{+2}, Mg^{+2}, Al^{+3} — **6.02×10^{23} ions**

Weight of 6.02×10^{23}

- Atoms → Gram atomic weight or 1 gm atom
- molecules → Gram molecular weight or 1 gm molecule
- Formula Units → Gram formula weight or 1 gm formula
- ions → Gram ionic weight or 1 gm ion

One Mole of a substance or 6.02×10^{23} molecules of a substance or 1 gm molecule of a substance	In gaseous → State at NTP occupy	22.4 litres or 22.4 dm^3 or 22400 ml. or 0.0224 m^3

The no. of molecules present in 1 cm^3 of an ideal gas at STP is called **Loschmidt number**. Its value is 2.69×10^{19}

Solved Problems

Illustration 15 **In three moles of ethane (C_2H_6). Calculate the following :**

(a) Number of moles of carbon atoms.

(b) Number of moles of, hydrogen atoms.

(c) Number of molecules of ethane.

One mole of C_2H_6 contains : 2 moles of carbon and 6 moles of H atom,

$\therefore$ 3 moles of C_2H_6 contain: 6 moles of carbon and 18 moles of H atoms

Molecules of C_2H_6 = 3 x 6.02^{23} = 18.06 x 10^{23}

Illustration 16 **Calculate the number of atoms in each of the following :**
(a) 52 moles of He (b) 52 u of He (c) 52 g of He

(a) One mole of He = 6.02 x 10^{23} atoms

$\therefore$ 52 mole of He = 52 x 6.02 x 10^{23} = 3.13 x 10^{25} atoms.

(b) $\because$ 4 u of He = one atom of He

$\therefore$ 52 u of He = $\frac{1\times 52}{4} = 13$ atoms

(c) One moles of He = 4 g = 6.02 x 10^{23} atoms

$\therefore$ 52 g He = $\frac{6.02\times 10^{23}\times 52}{4} = 7.83\times 10^{24}$ atoms.

Illustration 17 **Which one of the following will have largest number of atoms?**
(a) 1 g Au (s) (b) 1 g Na (s) (c) 1 g Li (s) (d) 1 g of Cl_2 (s)

(a) 1 g Au = $\frac{1}{197}$ mol $= \frac{1}{197}\times 6.02\times 10^{23}$ atoms

(b) 1 g Na = $\frac{1}{23}$ mol $= \frac{1}{23}\times 6.02\times 10^{23}$ atoms

(c) 1 g Li = $\frac{1}{7}$ mol $= \frac{1}{7}\times 6.02\times 10^{23}$ atoms we

(d) 1 g Cl_2 = $\frac{1}{71}$ mol $= \frac{1}{71}\times 6.02\times 10^{23}$ molecules $= \frac{2}{71}\times 6.02\times 10^{23}$ atoms

thus, Li has largest number of atoms.

Illustration 18 **What will be the mass of one ^{12}C atom in g?**

^{12}C shows that mass of 6.02 x 10^{23} carbon atoms = 12 gram

Mass of one C – atom = $\frac{12}{6.02\times 10^{23}} = 199\times 10^{-23}$ g

Illustration 19 **How many g atom and no. of atoms are there in**
(a) 60 g carbon (b) 224.4 g Cu ?
Given At. Wt. of C and Cu are 12 and 63.6 respectively.
Avogadro's no. = 6.02 × 10^{23}.

$\because$ g atom $= \frac{\text{wt.}}{\text{at. wt.}}$ and No. of atoms = $\frac{\text{wt.}\times \text{Av. No.}}{\text{at. wt.}}$

(a) $\therefore$ For 60 g C :

g atom = $\frac{60}{12} = 5$ No. of atoms $= \frac{60\times 6.02\times 10^{23}}{12} = 30.1\times 10^{23}$

(b) For 224.4 g Cu :

g atom = $\frac{224.4}{63.6} = 3.53$

No. of atoms = $\frac{224.4\times 6.02\times 10^{23}}{63.6} = 21.24\times 10^{23}$

Illustration 20 **Find the no. of g atoms and weight of an element having 2 × 10^{23} atoms. At. weight of element is 32.**

$\because$ N atoms have 1 g atom

2 × 10^{23} atoms have $= \frac{2\times 10^{23}}{6.023\times 10^{23}} = 0.33$ g atom

$\because$ N atoms of element weight 32 g

$\therefore$ 2 × 10^{23} atoms of element weight $= \frac{32\times 2\times 10^{23}}{6.023\times 10^{23}} = 10.63$ g

Illustration 21 **Calculate weight of 1 atom of hydrogen**

$\because$ N atom of H weigh 1 g

$\therefore$ 1 atom of H weigh $\frac{1}{6.023\times10^{23}} = 1.66\times10^{-24}$ g

Illustration 22 **How many gram atom of S are present in 49 g H_2SO_4 ?**

$\because$ 98 g H_2SO_4 has 32 g S or 1 g atom S

$\therefore$ 49 g H_2SO_4 has 32 g S $= \frac{1\times49}{98} = 0.5$ g atom of S

Illustration 23 **Calculate molecules of methane, C and H atoms in 25 g methane.**

Mol. wt. of CH_4 = 16

$\because$ 16 g CH_4 has molecules = N

$\therefore$ 25 g CH_4 has molecules = $\frac{6.023\times10^{23}\times25}{16} = 9.41\times10^{23}$ molecules

Further 16 g CH_4 has C atoms = N

$\therefore$ 24 g CH_4 has C atoms = $\frac{6.023\times10^{23}\times25}{16} = 9.41\times10^{23}$ atoms of C

Also 16 g CH_4 has H atoms = 4 N

$\therefore$ 25 g CH_4 has H atoms = $\frac{4\times25\times6.023\times10^{23}}{16} = 3.764\times10^{24}$ atoms of H

Illustration 24 **Weight of one atom of an element is 6.644×10^{-23} g. Calculate g atom of element in 40 kg.**

Wt. of 1 atom of element = 6.644×10^{-23} g

$\therefore$ Wt. of N atom of element = $6.644\times10^{-23}\times6.023\times10^{23}$ = 40

$\because$ 40 g weight of element has a 1 g atom

$\therefore$ 40×10^3 g of element has a 1 g atom = $\frac{40\times10^3}{40} = 10^3$ g atom

Illustration 25 **Calculate the no. of mole of water in 488 g $BaCl_2.2H_2O$.**

Mol. Wt. of $BaCl_2.2H_2O$ = 244 g

244 g $BaCl_2.2H_2O \equiv$ 36 g H_2O = 2 mole of water

488 $BaCl_2.2H_2O \equiv \frac{2\times488}{244}$ = 4 mole H_2O

Illustration 26 **Calculate the no. of atoms and volume of 1 g He gas at NTP.**

$\because$ 4 g He has 6.023×10^{23} atoms

$\therefore$ 1 g He has $\frac{6.023\times10^{23}}{4}$ atoms = 1.506×10^{23} atoms

Also,

$\because$ 4 g He has volume at NTP = 22.4 litre

$\therefore$ 1 g He has volume at NTP = $\frac{22.4}{4} = 5.6$ litre

Illustration 27 **Calculate the volume of STP occupied by (i) 14 g of nitrogen, (ii) 1.5 moles of carbon dioxide and (iii) 10^{21} molecules of oxygen.**

(i) Molecular mass of nitrogen = 28 u
1 mole of nitrogen = 28 g = 22.4 litres at STP
i.e. 28 g of nitrogen occupy 22.4 litres at STP
$\therefore$ 14 g of nitrogen will occupy $= \frac{22.4}{28} \times 14 = 11.2$ litres at STP

(ii) 1 mole of carbon dioxide = 22.4 litres at STP
$\therefore$ 1.5 moles of carbon dioxide will occupy

$\frac{22.4}{1} \times 1.5 = 33.6$ litres at STP

(iii) 1 mole of O_2 molecules
$= 6.022 \times 10^{23}$ molecules = 22.4 litres at STP
i.e. 6.022×10^{23} molecules of oxygen occupy 22.4 litres at STP
10^{21} molecules of oxygen will occupy
$= \frac{22.4}{6.022 \times 10^{23}} \times 10^{21}$ litres at STP
$= 3.72 \times 10^{-2}$ litres at sTP
$= 3.72 \times 10^{-2} \times 10^{3}$ cm^3 at STP = 37.2 cm^3 at STP

Illustration 28 **Calculate the mass of (i) 0.1 mole of KNO_3 (ii) 1×10^{23} molecules of methane and (iii) 112 cm^3 of hydrogen at STP.**

(i) 1 mole of KNO_3 = 101 g
($\because$ Formula mass of KNO_3 = $1 \times 39 + 1 \times 14 + 3 \times 16$ = 101 u)
$\therefore$ 0.1 mole of KNO_3 = 101×0.1 = 10.1 g of KNO_3

(ii) 1 mole of CH_4 = 16 g = 6.022×10^{23} molecules
i.e. 6.022×10^{23} molecules of methane have mass = 16 g
$\therefore$ 1×10^{23} molecules of methane will have mass
$= \frac{16}{6.022 \times 10^{23}} \times 10^{23} = 2.657$ g

(iii) 1 mole of H_2 = 2 g = 22400 cm^3 at STP
i.e. 22400 cm^3 of H_2 at STP have mass = 2 g
$\therefore$ 112 cm^3 of H_2 at STP will have mass
$= \frac{2}{22400} \times 112 = 0.01$ g

Illustration 29 **Calculate no. of oxalic acid molecules in 100 mL of 0.02N oxalic acid.**

Normality = 0.02

$\therefore$ Molarity $= \frac{0.02}{2}$ ($\because$ valency factor = 2)

$\therefore$ Mole of oxalic acid $= \frac{0.02}{2} \times \frac{100}{1000}$ ($\because$ mole = M $\times$ $V_{(\ell)}$)

$\therefore$ No. of molecules of oxalic acid = $0.001 \times 6.023 \times 10^{23} = 6.023 \times 10^{20}$

Illustration 30 **Insulin contains 3.4% sulphur. Then what will the minimum molecular mass of the insulin ?**

$\because$ 3.4 g sulphur is present in 100 g insulin

$\therefore$ 32 g sulphur will be present in $\frac{100}{3.4} \times 32$ g insulin = 940

$\therefore$ Molar mass of insulin is about 940 amu

Unsolved Problems

11. Calculate the following :
(i) Mass of 1×10^{23} molecules of CH_4
(ii) Number of moles and molecules in 18 g of glucose.
(iii) Mass of 0.2 mole of Ethane (C_2H_6).
(iv) Mass of 6.02×10^{22} molecules of cane sugar.
(v) Volume of 6.023×10^{25} molecules of carbon dioxide (CO_2).

12. Which of the following has largest number of oxygen atoms
1.0 g of O atoms, 1.0 g of O_2, 1.0 g of ozone (O_3) Justify your answer.

13. Find the total no. of protons in 10 g of $CaCO_3$?

14. How many g of S are required to produce 100 moles and 100 g H_2SO_4 separately ?

15. 10 g each of CO_2, NH_3 and O_2 are kept in three flasks. What will be the correct order of arrangements of flasks containing decreasing no. of atoms.

16. Calculate no. of moles present in
(i) 15 litre of H_2 gas at STP
(ii) 5 litre of N_2 gas at NTP
(iii) 0.5 g of H_2 gas
(iv) 10 g of O_2 gas

17. (a) Assuming the density of water to be 1 g/ml calculate the volume occupied by one molecule of water.
(b) Assuming water to be spherical, calculate the diameter of water molecule.

18. Calculate the no. of moles is following
a – 392 gm of sulphuric acid
b – 44.8 litres of CO_2 of STP
c – 6.022×10^{23} molecules of oxygen
d- 9 gm of aluminium
e – 1 metric ton of iron (1 metric ton = 10^3 kg)
f – 7.9 mg of C_9
g – 65 μ g of C

Composition of a solution : Composition or concentration of a solution refers to the amount of solute dissolved in a particular amount of solvent. Composition of a solution can be expressed in any of the following ways.

(i) **Weight percent (w/w) :** The amount (in grams) of a solute present in 100 g of the solution is called the weight percent of the solution.

(ii) **Volume percent :**
In case of liquid in liquid(v/v): This is generally used to express the composition of a solution of two liquids. It is defined as the volume of a solute (in ml) present in 100 ml of the volume of the solution.
In case of solid in liquid (w/v): It is defined as weight of solute (in grams) present in 100 ml of solution. **Grams per litre (g/L) :** It refers to the amount of solute in grams present in one litre of the solution.

(iii) **Molarity (M):** It is defined as the number of moles (or gram moles) of a solute dissolved per litre of the solution at a particular temperature. It is expressed by M. 1 M solution is called a **molar solution** and contains 1-gram mole of solute dissolved in 1 litre of the solution.
The unit of molarity is mol L^{-1} or mol dm^{-3}.
If w grams of a solute are dissolved in V cm^3 of a solution, its molarity M is given by

$$M = \frac{n \times 1000}{V(mL)} = \frac{n}{V(L)} = \frac{w \times 1000}{M' \times V(mL)}$$ [where M′ is molecular mass of solute]

If V_1 mL of solution of M_1 react exactly with V_2 mL of solution M_2 then – $M_1 \times V_1 = M_2 \times V_2$
However if reaction between 2 solution is studied in terms of molarity

$$\Rightarrow \quad \frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2}$$

where n_1 = no. of moles of reactant 1
n_2 = no. of moles of reactant 2
If V_1 ml of a solution of M_1 is mixed with V_2 ml of solution of M_2 then M_3 can be calculated as

$$M_3 = \frac{M_1V_1 + M_2V_2}{V_1 + V_2}$$

Molarity of a solution is dependent on the temperature because it involves the volume which is a function of temperature.

(iv) **Formality (F) :** It is defined as the number of gram formula mass dissolved in 1 litre of the solution at a particular temperature. It is also dependent on temperature.

(v) **Normality (N) :** It is defined as the number of gram equivalents of the solute dissolved per litre of the solution at a particular temperature.

If w grams of a solute are present in V cm^3 of a solution, then its normality N is given by

$$N = \frac{w \times 1000}{eq.wt. \times V(ml)}$$

Also equivalent = N × V in litre = $\frac{\text{wt. of solute}}{\text{equivalent weight of solute}}$

Mili equivalent = N × V in ml = $\frac{\text{wt. of solute}}{\text{equivalent weight of solute}} \times 1000$

Law of Equivalence

Equivalent and mili equivalent of reactant reacts in equal number to give same number of equivalent or mili equivalent of product separately. So according to Law of Equivalence

If V_1 mL of solution N_1 react exactly with V_2mL of solution N_2 then

$$N_1 \times V_1 = N_2 \times V_2$$

If V_1 mL of a solution of N_1 is mixed with V_2 mL of solution of N_2 then N_3 can be calculated as

$$N_3 = \frac{N_1V_1 + N_2V_2}{V_1 + V_2}$$

Relationship between Normality and Molarity

Molarity = $\frac{\text{Equivalent mass}}{\text{Molecular mass}}$ x Normality

(vi) **Molality (m) :** It is defined as the number of gram moles of a solute dissolved in 1000 g (1 kg) of the solvent.

If w grams of a solute are dissolved in w grams of a solvent, the molarity m of the solution is given by

$$m = \frac{n \times 1000}{W(gm)} = \frac{w \times 1000}{M' \times W(gm)}$$

where M′ is the molecular mass of the solute, W is the weight of solvent. Molality of a solution is independent of the temperature.

Solved Problems

Illustration 32 **A solution is prepared by adding 2 g of a substance A to 18 g of water. Calculate the mass percent to the solute.**

Solution Mass percent of A = $\frac{\text{Mass of A}}{\text{Mass of solution}} \times 100$

$$= \frac{2\,g}{2\,g \text{ of } A + 18\,g \text{ of water}} \times 100$$

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

Illustration 33 **Calculate the molarity of NaOH in the solution prepared by dissolving its 4 g in enough water to form 250 mL of the solution.**

Solution Since molarity (M)

$$= \frac{\text{No. of moles of solute}}{\text{Voluime of solution in litres}}$$

$$= \frac{\text{Mass of NaOH / Molar mass of NaOH}}{0.250\ \text{L}}$$

$$= \frac{4\ \text{g} / 40\ \text{g}}{0.250\ \text{L}} = \frac{0.1\ \text{mol}}{0.250\ \text{L}}$$

$= 0.4\ \text{mol L}^{-1} \quad = 0.4\ \text{M}$

Note that molarity of a solution depends upon temperature because volume of a solution is temperature dependent.

Illustration 34 **Calculate the concentration of nitric acid in moles per litre in a sample which has a density, 1.41g mL⁻ and the mass per cent of nitric acid in it being 69%.**

Solution Molar mass of nitric acid (HNO_3) = 63

Concentration in moles per litre = $\frac{\% \times 10 \times d}{M.M} = \frac{69 \times 10 \times 141}{63} = 15.44$

Illustration 35 **What is the concentration of sugar ($C_{12}H_{22}O_{11}$) in mol L⁻ if its 20 g are dissolved enough in water to make a final volume up to 2 L?**

Solution Molar mass of $C_{12}H_{22}O_{11}$ = 342

Molar conc. = $\frac{w}{mx\ V(lit)} = \frac{20}{342 \times 2} = 0.0292$

Illustration 36 **If the density of methanol is 0.793 kg L⁻ , what is its volume needed for making 2.5 L of its 0.25 M solution?**

Solution Molar mass of methanol (CH_3OH) = 32, d = 0.793 g/mL

Molarity = $\frac{W}{mx\ V(lit)}$ = or $0.25 = \frac{w}{32 \times 2.5}$ or $w = 20$, Volume $= \frac{\text{mass}}{\text{density}} = \frac{20}{0.793} = 25.22\text{mL}$

Illustration 37 **The concentration of cholesterol ($C_{27}H_{46}O$) in normal blood is approximately 0.005 M. How many grams of cholesterol are in 750 mL of blood?**

Solution Let amount of cholesterol in normal blood

$= x\ g = \frac{x}{386}\ \text{mol} = \frac{x}{386 \times 0.750} = 0.005\ \text{M}$

$\Rightarrow \quad x = 1.4475\ g$

Illustration 38 **How are 0.50 mol Na_2CO_3 and 0.50 M Na_2CO_3 different?**

Solution Molar mass of Na_2CO_3 = 2 x 23 + 1 x 12 + 3 x 16 = 106

0.5 mol Na_2CO_3 = 0.5 x 106 = 53 g Na_2CO_3

0.50 M Na_2CO_3 shows that 0.50 mole or 0.5 x 126 g, 53 g Na_2CO_3 is present in one litre solution.

Illustration 39 **What mass (in grams) of a 0.500 molal solution of sodium acetate in water would you use to obtain 0.150 mole sodium acetate (82 g mol^{-1})**

Solution Molality = $\frac{\text{mole of solute}}{\text{kg of solvent}}$

$\Rightarrow \quad 0.5 = \frac{w \times 1000}{82 \times 1000} \quad \Rightarrow \quad w = 41\ g$

If solution is 0.500 molal, it means 1000 g H_2O has 0.500 mol (= 41 g) sodium acetate

$\therefore$ Total mass of solution = 1041 g

Thus, 0.500 mol of sodium acetate is in =1041 g solution

Hence, 0.150 mol of sodium acetate is in = $\frac{1041}{0.500} \times 0.150$ = 312.3 g

Illustration 40 **The density of 3 M solution of NaCl is 1.25 g mL^{-1}. Calculate molality of the solution.**

Solution M = 3 mol L^{-1}

Mass of NaCl

In 1L solution = 3×58.5 = 175.5 g

Mass of

1 L solution = 1000×1.25 = 12 50 g

(Since density = 1.25 g mL^{-1})

Mass

Water in solution = 1250 – 175.5

= 1074.5 g

Molality = $\frac{\text{No. of moles of solute}}{\text{Mass of solvent in kg}}$

$= \frac{3 \text{ mol}}{1.0745 \text{ kg}}$

= 2.79 m

Often in a chemistry laboratory, a solution of a desired concentration is prepared by diluting a solution of known higher concentration. The solution of higher concentration is also known as stock does not change with temperature since mass remains unaffected with temperature.

Illustration 41 **Calculate the normality of the resulting solution made by adding 2 drops (0.1 mL) of 0.1N H_2SO_4 in 1 litre of distilled water.**

∵ Meq. of solute does not change on dilution

Meq. of H_2SO_4 (conc.) = Meq. of H_2SO_4(dil.)

0.1 × 0.1 = N × 1000 (∵ Meq. = N × V is mL)

N = 10^{-5}

Illustration 42 **Calculate the amount of KOH required to neutralize 15 Meq. of the following:**

(a) HCl **(b) $KHSO_4$** **(c) N_2O_5** **(d) CO_2**

∵ 15 Meq. of each separately react with KOH and therefore only 15 Meq. of KOH are required every time.

∴ Meq. of KOH required = 15 $\frac{w}{56} \times 1000 = 15$ ∴ w = 0.84 g

Illustration 43 **What volume of water is required to make 0.20 N solution from 1600 mL of 0.2050 N solution?**

Meq. of conc. solution = 1600 × 0.2050 = 328

Let after dilution volume becomes V mL

Meq. of dil. solution = 0.20 × V

∴ 328 = 0.20 × V ∴ V = 1640 mL

Thus volume of water used to prepare 1640 mL of 0.20 N solution

= 1640 – 1600 = 40 mL

Unsolved Problems

19. Calculate normality and molarity of the following:

(a) 0.74 g of $Ca(OH)_2$ in 5 ml of solution

(b) 3.65 g of HCl in 200 ml of solution

(c) 1/10 mole of H_2SO_4 in 500 ml of solution

20. Calculate normality of mixture obtained by mixing:

(a) 100 ml of 0.1 N HCl + 50 ml of 0.25 NaOH

(b) 100 ml of 0.2 M H_2SO_4 + 200 ml of 0.2 M HCl

(c) 100 ml of 0.2 M H_2SO_4 + 100 ml of 0.2 M NaOH.

(d) 1 g equivalent of NaOH + 100 ml of 0.1 N HCl.

21. What weight of Na_2CO_3 of 95% purity would be required to neutralize 45.6 ml of 0.235 N acid?

22. How much AgCl will be formed by adding 200 mL of 5N HCl to a solution containing 1.7 g $AgNO_3$.

23. Calculate molality of 1 litre solution of 93% H_2SO_4 (w/v). The density of solution is 1.84 g ml^{-1}

STOICHIOMETRY

It refers to the calculations based on chemical equations. The overall chemical change can be conveniently represented by the balanced chemical equation. The chemical provides us information regarding:

(a) Molar ratio between reactants and products

(b) Mass ratio between reactants and products and

(c) Volume ratio between gaseous reactants and products.

The calculations based on the chemical equations are of three types namely:

(i) mass – mass relationship

(ii) mass – volume relationship

(iii) volume – volume relationship

Limiting reagent

In general when a chemical reaction is carried out, one of the reagents will be used in excess and it is known as excess reagent. The reagent that is not present in excess is the one that will determine how much product can be obtained and is thus referred to as the limiting reagent. The reagent that gives the least number of moles of the product is the limiting reagent. With the help of limiting reagent, the amount of the excess reagent, which has been utilised and the amount remaining can be calculated.

Percent yield

Generally when a reaction is carried out in the laboratory we do not obtain actually the theoretical amount of the product. The amount of the product that is actually obtained is called the actual yield. Knowing the actual yield and theoretical yield, the percent yield can be calculated by the following formula.

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoritical yield}} \times 100$$

Solved Problems

Illustration 44 **Calculate the amount of water (g) produced by the combustion of 16 g of methane.**

Solution The balanced equation for combustion of methane is:

$CH_4\,(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2H_2O\,(g)$

(i) 16 g of CH_4 corresponds one mole.

(ii) From the above equation, 1 mole of CH_4 (g) gives 2 mol of H_2O (g). 2 mol of water (H_2O) = 2 × (2 + 16) = 2 × 18 = 36 g

1 mol H_2O = 18 H_2O $\Rightarrow \frac{18\,g\,H_2O}{1\,mol\,H_2O} = 1$

Hence 2 mol H_2O × $\frac{18\,g\,H_2O}{1\text{ mol }H_2O}$

=2 × 18 g H_2O = 36 g H_2O

Illustration 45 **How many moles of methane are required to produce 22 g CO_2 (g) after combustion?**

Solution According to the chemical equation,

$CH_4\,(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2H_2O\,(g)$

44g CO_2 (g) is obtained from 16 g CH_4 (g).

[∵ 1 mol CO_2 (g) is obtained from 1 mol or CH_4 (g)]

Mole of CO_2 (g)

$= 22\,g\ CO_2\,(g) \times \frac{1\,mol\,CO_2\,(g)}{44\,g\,sCO_2\,(g)}$

=0.5 mol CO_2 (g)

Hence, 0.5 mol CO_2 (g) would be obtained from 0.5 mol CH_4 (g) or 0.5 mol of CH_4 (g) would be required to produce 22 g CO_2 (g).

Illustration 46 **Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction,**

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

What mass of $CaCO_3$ is required to react completely with 25 mL of 0.75 M HCl ?

Solution Moles of HCl = molarity x V(lit) = 0.75 x 25 x 10^{-3} = 0.01875

Mass of HCl = moles x mol. wt. of HCl (36.5) = 0.01875 x 36.5 = 0.684 g

$CaCO_3 + 2\,HCl \rightarrow CaCl_2 + H_2O + CO_2$

100 g 2 x 36.5 =73 g

∵ 73 g HCl requires $CaCO_3$ = 100 g

∴ 0.684 g HCl requires $CaCO_3$ = $\frac{100 \times 0.684}{73} = 0.937\,g$

Illustration 47 **Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid according to the reaction**

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

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

Solution $4\,HCl + MnO_2 \rightarrow MnCl_2 + Cl_2 + 2H_2O$

146 g 87 g

∵ 87 g MnO_2 reacts with HCl = 146 g

$\therefore$ 5 g MnO_2 reacts with HCl = $\dfrac{146 \times 5}{87} = 8.39$ g

Illustration 48 **3.68 g of a mixture of calcium carbonate and magnesium carbonate when heated strongly leaves 1.92 g of a white residue. Find the percentage composition of the mixture.**

Solution Let x g of $CaCO_3$ be present in the mixture.

The mass of $MgCO_3$ in the mixture = (3.68 – x) g.

$$\underset{100\,g}{CaCO_3} \xrightarrow{Heat} \underset{56\,g}{CaO} + CO_2\ ;\ \underset{84\,g}{MgCO_3} \xrightarrow{Heat} \underset{40\,g}{MgO} + CO_2$$

$$\therefore\ \frac{56 \times x}{100} + \frac{40 \times (3.68 - x)}{84} = 1.92$$

On solving, x = 2

Percentage of $CaCO_3$ = $\dfrac{2}{3.68} \times 100 = 54.35\%$

Percentage of $MgCO_3$ = 100 – 54.35 = 45.65%

Illustration 49 **Calculate the mass of (CaO) that can be prepared by heating 200 kg of limestone ($CaCO_3$) which is 95% pure.**

Solution Amount of pure $CaCO_3$ = $\dfrac{95}{100} \times 200$ = 190 kg = 190000 g

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

1 mole $CaCO_3 \equiv$ 1 mole CaO

100 g $CaCO_3 \equiv$ 56 g CaO

$\because$ 100 g $CaCO_3$ give 56 g CaO

$\therefore$ 190000 g $CaCO_3$ will give $\dfrac{56}{100} \times$ 190000 g CaO =106400 g = 106.4 kg

Illustration 50 **Calculate the weight of iron which will be converted into its oxide by the action of 18 g of steam**

Solution The reaction occurs as

$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$

Mole ratio of reaction suggests :

$$\frac{\text{Mole of Fe}}{\text{Mole of } H_2O} = \frac{3}{4}$$

$\therefore$ Mole of Fe = $\dfrac{18}{18} \times \dfrac{3}{4} = \dfrac{3}{4}$

$\therefore$ Weight of Fe = $\dfrac{3}{4} \times 56 = 42$ g

Illustration 51 **If ten volumes of dihydrogen gas reacts with five volumes of dioxygen gas, how many volumes of water vapour would be produced?**

Solution $2H_2 + O_2 \rightarrow 2H_2O$

2 vol 1 vol 2 vol

10 vol of H_2 will produce 10 vol of H_2O and 5 vol of H_2 will produce 10 vol of H_2O

No reactant will remain left behind. H_2O vapours produced = 10 vol.

Illustration 52 **In a reaction $A + B_2 \rightarrow AB_2$**

Identify the limiting reagent. If any, in the following reaction mixtures.
(a) 300 atoms of A + 200 molecules of B_2
(b) 2 mol A + 3 mol B_2
(c) 100 atoms of A + 100 molecules of B
(d) 5 mot A + 2.5 mol B_2
(e) 2.5 mol A + 5 mol B_2

Solution A + B_2 $\rightarrow AB_2$
1 atom 1 molecule
2 atom

(a) 300 atoms of A require 300 molecules of B_2. But we have 200 molecules of B_2. Therefore, A is in excess. Hence, limiting reagent is B_2.
(b) 2 moles of A require 2 moles of B_2. We have 3 moles of B_2. Therefore, B_2 is in excess. Hence, A is the limiting reagent.
(c) 100 atoms of A require 100 molecules of B_2. No reactant is in excess. There is no limiting reagent.
(d) 2.5 mole of A will react with 2.5 mole of B_2. Thus A is an excess. Limiting reagent is B.
(e) 2.5 moles of A require 2.5 moles of B_2. The reagent B_2 (5mol) is in excess. Hence, A is the limiting reagent.

Illustration 53 **Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation :** $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
(a) Calculate the mass of 'ammonia produced if 2.00 x 10^3 g dinitrogen reacts with 1.00 x 10^3 g of dihydrogen.
(b) Will any of the two reactants remain unreacted?
(c) If yes, which one and what would be its mass?

Solution $N_2 + 3H_2 \rightarrow 2NH_3$ 28 g 6 g 2 x 17 = 34 g
$\because$ 28 g N_2 produces ammonia = 34 g
$\therefore$ 2 x 10^3 g N_2 produces ammonia = $\frac{34\times2\times10^3}{28} = 2.43\times10^3$g
Again $\because$ 6 g H_2 produces ammonia = 34 g
$\therefore$ 1 x 10^3 g H_2 produces ammonia = $\frac{34\times1\times10^3}{6} = 5.67\times10^3$g
ammonia obtained from N_2 is less in quantity. Therefore ammonia produced = 2.43 x 10^3 g.
(b) Ammonia obtained from N_2 is less, hence N_2 is the limiting reagent. Some H_2 will remain unreacted.
(c) $\because$ 28 g N_2 reacts with H_2 = 6 g
$\therefore$ 2 x 10^3 g N_2 with $H_2 = \frac{6\times2\times10^3}{28} = 4.286\times10^2$g
H_2 left = 1 x 10^3 – 4.286 x 10^2 = 571. 4 g

Illustration 54 **1 g of Mg is burnt in a closed vessel which contains 0.5 g of O_2:**
(i) Which is limiting reagent ?
(ii) Which reactant is left in excess?
(iii) Find the mass of the excess reactant.

Solution The balanced equation is:

$$\underset{\substack{2\times24\\=48\,g}}{2Mg} + \underset{\substack{2\times16\\=32\,g}}{O_2} \longrightarrow \underset{\substack{2(24+16)\\=80g}}{2MgO}$$

48 g of Mg require oxygen = 32 g
1 g of Mg requires oxygen = $\frac{32}{48}$ = 0.66 g
But only 0.5g oxygen is available. Hence, O_2 is a limiting agent and a part of magnesium will not burn.
$\therefore$ Magnesium will be left in excess.
32 g of O_2 react with magnesium = 48 g
0.5 g of O_2 will react with magnesium = $\frac{48}{32}$ × 0.5 = 0.75 g
Hence the mass of excess magnesium = (1.0–0.75) = 0.25 g

Illustration 55 **50.0 kg of N_2 (g) are mixed to produce NH_3 (g). Calculate the NH_3 (g) formed. Identify the limiting reagent in the production of NH_3 in this situation.**

Solution A balanced equation for the above reaction is written as follows:

Calculate of moles:

N_2 (g) + $3H_2$ (g) $\rightleftharpoons$ NH_3 (g)

Moles of N_2

$$= 50.0\,kg\,N_2 \times \frac{1000\text{ g } N_2}{1\text{ kg } N_2} \times \frac{1\text{ mol } N_2}{28.0\text{ g } N_2} \quad = 17.86 \times 1\,10^2 \text{ mol}$$

Moles of H_2

$$= 10.00\text{ kg}\,H_2 \times \frac{1000\text{ g}\,H_2}{1\text{ kg } N_2} \times \frac{1\text{ mol } H_2}{2.016\text{ g } H_2} = 4.96 \times 10^3\text{ mol}$$

According to the above equation, 1 mol N_2 (g) requires 3 mol H_2 (g), for the reaction. Hence, for 17.86 $\times 10^2$ mol of N_2, the moles of H_2 (g) required would be

$$17.86 \times 10^2\text{ mol}\,N_2 \times \frac{3\text{ mol}\,H_2\text{ (g)}}{1\text{mol}\,N_2\text{ (g)}}$$

= 5.36×10^3 mol H_2

But we have only 4.96×10^3 mol H_2. Hence, dihydrogen is the limiting reagent in this case. So NH_3 (g) would be formed only from that amount of available dihydrogen i.e., 4.96×10^3 mol

Since 3 mol H_2 (g) give 2 mol NH_3 (g)

$$4.96 \times 10^3\text{ mol } H_2\text{ (g)} \times \frac{2\text{ mol } NH_3\text{ (g)}}{3\text{ mol } H_2\text{ (g)}} \quad = 3.30 \times 10^3\text{ mol } NH_3\text{ (g)}$$

3.30×10^3 mol NH_3 (g) is obtained.

If they are to be converted to grams, it is done as follows:

1 mol NH_3 (g) 17.0 g NH_3 (g)

$$3.30 \times 10^3\,1\text{ mol } NH_3\text{ (g)} \times \frac{17.0\text{ g } NH_3\text{ (g)}}{1\text{mol } NH_3\text{ (g)}} = 3.30 \times 10^3 \times 17\text{ g } NH_3\text{ (g)}$$

= 56.1×10^3 g NH_3 56.1Kg NH_3

Unsolved Problems

24. Calculate the mass of 60% H_2SO_4 required to decompose 50 g of chalk (calcium carbonate).

25. How much marble of 96.5% purity would be required to prepare 10 litres of carbon dioxide at STP when the marble is acted upon by dilute hydrochloric acid ?

26. Find out the volume of Cl_2 at STP produced by the action of 100 cm^3 of 0.2 N HCl on excess of MnO_2.

27. The reaction, $2C(s) + O_2(g) \longrightarrow 2CO(g)$
is carried out by taking 24 g of carbon and 96 g O_2. Find out:
(a) Which reactant is left in excess?
(b) How much of it is left?
(c) How many moles of CO are formed?
(d) How many grams of other reactant should be taken so that nothing is left at the end of the reaction?

28. If 20 g of $CaCO_3$ is treated with 20 g of HCl, how many grams of CO_2 can be generated according to the following equation?
$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

Percentage composition of compounds

For the determination of empirical and molecular formula the percentage composition of each element of the compound should be known. It represents the weight of each of the constituent elements in 100 parts of it. The percentage of the compound is readily calculated from the formula of the compound.

Molecular weight of the substance is obtained from its formula by adding up the weights of all the atoms of the constituent elements present in the molecule.

Let the molecular weight of a substance be M and x be the weight of an element present in the molecule. Then

$$\% \text{ of element} = \frac{\text{weight of element}}{M} \times 100 = \frac{x}{M} \times 100$$

Empirical formula. The formula of a compound, which gives the simple whole number ratio of the atoms of the various elements present in one molecule of the compounds is called its Empirical formula. The empirical formula of glucose ($C_6H_{12}O_6$) has been found to be CH_2O which indicates that the ratio of C, H and O atoms in molecules of glucose is 1 : 2 : 1.

Molecular formula. It is the actual formula of a compound, which gives the actual number of atoms of various elements present in one molecule of the compound e.g. molecular formula of glucose is $C_6H_{12}O_6$.

Calculation of Empirical formula mass. It is obtained by adding the atomic masses of the various atoms present in the empirical formula.
Relationship between empirical formula and molecular formula
Molecular formula = n × Empirical formula
where n is any integer such as 1, 2, 3 …. etc.

$$n = \frac{\text{Molecules mass}}{\text{Empirical formula mass}}$$

Calculation of molecular mass

(i) Molecular mass = 2 × Vapour density (V.D.)

$$\text{V.D.} = \frac{\text{Wt. of a certain volume of a gas(vapours)}}{\text{Wt. of same volume of } H_2 \text{ under the same conditions of temperature and pressure}}$$

(ii) Molecular mass = mass of 22.4 litres of gas at S.T.P.

Calculation of empirical formula. First calculate % age of oxygen = 100 – Sum of % ages of all other elements. Then E.F. is calculated through the following steps.

Element	% age	Atomic mass	Relative no. of atom $= \frac{\%\text{ age}}{\text{At. mass}}$	Simplest Atomic ratio	Simplest whole No. atomic ratio.

Solved Problems

Illustration 56 **An element A (at.wt. = 75) and B (at. wt. = 25) combine to form a compound. The compound contains 75% A by weight. What is the formula of the compound ?**

g atom of A = $\frac{75}{75} = 1$;

g atom of B = $\frac{25}{25} = 1$;

$\therefore$ Ratio of g atoms of A and B = 1:1.

Illustration 57 **A welding fuel gas contains carbon and hydrogen only. Buring 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 :**
(a) empirical formula, (b) molar mass of the gas and (c) molecular formula.

Molar mass of CO_2 = 44 g, Mas of carbon in it = 12 g

Mass of carbon in 3.38 g of CO_2 = $\frac{12 \times 3.38}{44} = 0.9218$

Molar mass of H_2O = 18 g, Mass of hydrogen in it = 2 g

$\therefore$ Mass of hydrogen in 0.690 g H_2O = $\frac{2 \times 0.690}{18} = 0.0767$

since fuel gas contains only C and H.
therefore total mass = 0.9218 + 0.0767 = 0.9985 g

$\% C = \frac{0.9218 \times 100}{0.9985} = 92.32$, $\quad \% H = \frac{0.0767 \times 100}{0.9985} = 7.68$

Let us calculate the empirical formula :

% by mass	Divide by atomic mass	Simple molar ratio	Simplest whole number ratio
C = 92.32	$\frac{92.32}{12}=7.69$	$\frac{7.69}{7.68}=1$	1
H = 7.68	$\frac{7.68}{1}=7.68$	$\frac{7.68}{7.68}=1$	1

Empirical formula of fuel gas = CH , Let us calculate the molar mass of fuel gas

$\because$ 10 L of the gas weighs = 11.6 g

$\therefore$ 22.4 L of the gas weighs = $\frac{11.6\times 22.4}{10}=25.986=26$

Empirical formula mass of CH = 12 + 1 = 13

n = $\frac{\text{Molecular mass}}{\text{empirical mass}}=\frac{26}{13}=2$

$\therefore$ Molecular formula = $(CH)_2 = C_2H_2$

Illustration 58 **A compound containing sodium, sulphur, hydrogen and oxygen gave the following results on analysis :-**

Na = 14.28%, S = 9.92%, H = 6.20%

Calculate the molecular formula of the anhydrous compound. If all the atoms of hydrogen in the compound are present in combination with oxygen as water of crystallization, what is the structure of the crystalline salt ? The molecular mass of the crystalline salt is 322.

Element	Symbol	Percentage of element	At. mass of element	Relative no. of atoms = $\frac{\text{Percentage}}{\text{At. mass}}$	Simplest atomic ratio	Simples wholeno. atomic ratio
Sodium	Na	14.28	23	$\frac{14.28}{23}=0.62$	$\frac{0.62}{0.31}=2$	2
Sulphur	S	9.92	32	$\frac{9.92}{32}=0.31$	$\frac{0.31}{0.31}=1$	1
Hydrogen	H	6.20	1	$\frac{6.2}{1}=6.20$	$\frac{6.20}{0.31}=20$	20
Oxygen	O	69.60	16	$\frac{69.60}{16}=4.35$	$\frac{4.35}{0.31}=14$	14

$\therefore$ Empirical Formula is $Na_2SH_{20}O_{14}$

Empirical formula mass= $2\times 23.0 + 1\times 32.0 + 20\times 1.0 + 14\times 16.0 = 322$

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

Molecular formula = n × Empirical formula = 1 × $Na_2SH_{20}O_{14}$. ∴ $Na_2SO_4 . 10H_2O$

Illustration 59 **The percentage composition of ferrous ammonia sulphate is 14.32% Fe^{2+} ; 9.20% NH_4^+ ; 49.0% SO_4^{2-} and 27.5% H_2O. What is the empirical formula of the compound ?**

Calculation of empirical formula.

Formula of the constituent	**Percentage**	**Mol. mass / At. mass of constituents**	**Relative no. of constituents** $= \frac{\text{Percentage}}{\text{Mol. mass}}$	**Simplest ratio of constituents**
Fe^{2+}	14.32	56	$\frac{14.32}{56} = 0.2557$	$\frac{0.2557}{0.2557} = 1$
NH_4^+	0.20	18	$\frac{9.20}{18} = 0.5111$	$\frac{0.5111}{0.2557} = 2$
SO_4^{2-}	49.0	96	$\frac{49.0}{96} = 0.5104$	$\frac{0.5104}{0.2557} = 2$
H_2O	27.57	18	$\frac{27.57}{18} = 1.532$	$\frac{1.532}{0.2557} = 6$

Hence the empirical formula of the compound is $(Fe^{2+})\ (NH_4^+)_2\ (SO_4^{2-})_2\ (H_2O)_6$

Illustration 60 **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 = 15.5% ; Sulphur = 24.96% ; Oxygen = 49.2%.
Find the simplest formula of the anhydrous and crystalline salt.

Element	**Symbol**	**Percentage of elements**	**At. mass of elements**	**Relative no. of atoms** $= \frac{\text{Percentage}}{\text{At.mass}}$	**Simplest atomic ratio**	**Simplest whole no. atomic ratio**
Potassium	K	15.1	39	$\frac{15.10}{39} = 0.39$	$\frac{0.39}{0.39} - 1$	1
Aluminium	Al	10.50	27	$\frac{10.50}{27} = 0.39$	$\frac{0.39}{0.39} = 1$	1
Sulphur	S	24.96	32	$\frac{24.96}{32} = 0.78$	$\frac{0.78}{6.39} = 2$	2
Oxygen	O	49.92	16	$\frac{49.92}{16} = 3.12$	$\frac{3.12}{0.39} = 8$	8

Thus the empirical formula of the anhydrous salt is $KAlS_2O_8$

Empirical formula mass of anhydrated salt

= 1 × 39.0 + 1 × 27.0 + 2 × 32.0 + 8 × 16.0 = 258.0 u

To calculate the empirical formula mass of Hydrated salt

Let the Empirical formula mass of Hydrated salt = 100.0 u

Loss of weight due to dehydration = 45.6%
Weight of anhydrous salt is 100 – 45.6 = 54.4 u
If empirical formula mass of the anhydrous salt is 258, that of hydrated is

$$\frac{100}{54.4} \times 258 = 473.4 \text{ u}$$

Total loss in mass due to dehydration

$$= 474.3 - 258.0 = 216.3 \text{ u}$$

No. of molecules of water

$$= \frac{216.3}{18} = 12$$

So formula is $KAlS_2O_8 \cdot 12\ H_2O$

Unsolved Problems

29. Calculate percentage composition of each in Na_2SO_4 ?

30. Find the percentage composition of potassium chlorate ($KClO_3$).

31. An organic liquid having carbon, hydrogen, nitrogen and oxygen was found to contain C= 41.37%, H = 5.75%, N = 16.09% and the rest oxygen. Calculate the molecular formula of the liquid if its V.D. is 43.3.

Solved Conceptual Problems

1. ***What the is the equivalent weight of $KH(IO_3)_2$ as an oxidant in pressure of 4.0 (N) HCl when ICl becomes the reduced form ?(K = 39.0, I = 127.0).***

In $KH(IO_3)_2$, IO_3 is present as IO_3^-. Oxidation state of I will be

$$x - 6 = -1 \quad \text{or} \quad x = +5$$

$$KH(\overset{+5}{I}O_3)_2 \rightarrow 2\ \overset{+1}{I}Cl$$

Decrease in oxidation state = +10 –2 = 8

$\therefore$ Equivalent wt. of $KH(IO_3)_2 = \frac{\text{Mol. wt.}}{8} = \frac{39 + 1 + 2(127 + 48)}{8} = 48.75$

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

3. ***Is the law of constant composition true for all types of compounds? Explain why or why not.***
No, law of constant composition is not true for all types of compounds. It is true for the compounds obtained from one isotope. For example, carbon exists in two common isotopes, ^{12}C and ^{14}C. When it forms CO_2 from ^{12}C, the ratio of masses is 12 : 32 = 3 : 8 but from ^{14}C, the ratio will be 14 : 32 = 7 :16 which is not same as in the first case.

4. ***Why Law of conservation of mass should better be called as Law of conservation of mass and energy?***
In nuclear reactions, it is observed that the mass of the products is less than mass of the reactants. The difference of mass, called the mass defect, is converted into energy according to Einstein equation, $E = \Delta m c^2$. Hence, we better call it as a law of conservation of mass and energy.

5. ***Calculate equivalent mass of iron in $FeCl_3$?***

Equivalent mass of iron = $\frac{\text{Molecular mass}}{\text{valency}} = \frac{56}{3} = 18.66$

Solved Objective Problems

1. The number of gram molecules of oxygen in 6.02×10^{24} CO molecules is
(a) 10 g molecules **(b)** 5 g molecules
(c) 1 g molecule **(d)** 0.5 g molecule.

Sol. **(b)** 1 mol of CO $\equiv$ 1 g atom of O or 6.02×10^{23} molecules of CO $\equiv \frac{1}{2}$ g molecules of O_2

$\therefore$ 6.02×10^{24} molecules of CO

$= \frac{1}{2} \times \frac{6.02 \times 10^{24}}{6.02 \times 10^{23}}$ = 5 g molecules

2. Two containers P and Q of equal volume (1 litre each) contain 6g of O_2 and SO_2 respectively at 300 K and 1 atmosphere. Then
(a) No. of molecules in P is less than in Q
(b) No. of the molecules in Q is less than that in P
(c) No. of molecules in P and Q are same
(d) Either **(a)** or **(b)**

Sol. **(b)** 6 g O_2 = 6/32 mol, 6 g SO_2 = 6/44 mol. Thus vessel Q has less no. of moles.

3. Haemoglobin contains 0.33% of iron by weight. The molecular weight of haemoglobin is approximately 67200. The number of iron atoms (at. wt. of Fe = 56) present in one molecule of haemoglobin is
(a) 6 **(b)** 1
(c) 4 **(d)** 2

Sol. **(c)** Fe presen in 67200 u = $\frac{0.33}{100} \times 67200$

= 222 u = $\frac{222}{56}$ = 4 atoms

4. In the reaction 4 NH_3(g) + 5 O_2 (g) $\longrightarrow$ 4 NO (g) + 6 H_2O (ℓ), when 1 mole of ammonia and 1 mole of O_2 are made to react to completion
(a) 1.0 mole of NO is produced **(b)** 1.0 mole of NO will be produced
(c) all the oxygen will be consumed **(d)** all the ammonia will be consumed

Sol. **(c)** O will be limiting reactant. Hence all oxygen will be consumed.

5. 50 ml 10 NH_2SO_4, 25 m 2N HCl and 40 ml 5 N HNO_3 were mixed together and the volume of the mixture was made 1000 ml by adding water. The normality of the resultant solution will be:
(a) 1 N **(b)** 2 N
(c) 3 N **(d)** 4 N

Sol. **(a)** $N_1V_1 + N_2V_2 + N_3V_3 = N_4V_4$ $10 \times 50 + 12 \times 25 + 5 \times 40 = N_4 \times 1000$ of N_4 = 1 N

6. 10 dm^3 of N_2 gas and 10 dm^3 of gas X at the same temperature contain the same number of molecules. The gas X is
(a) CO **(b)** CO_2
(c) H_2 **(d)** NO

Sol. **(a)** No. of moles of N_2 and X should be equal. This can be so if X has same molecular weight as N_2.

7. How many moles of electron weight one kilogram?
(a) 6.023×10^{23} **(b)** $\frac{1}{9.108} \times 10^{31}$
(c) $\frac{6.023}{9.108} \times 10^{54}$ **(d)** $\frac{1}{9.108 \times 6.023} \times 10^{8}$

Sol. **(d)** 1 mol fo electrons weight
= $(9.108 \times 10^{-31}) \times (6.023 \times 10^{23})$kg

or $9.108 \times 10^{-31} \times 6.023 \times 10^{23}$ kg of electrons = 1 mol of electrons
$\therefore$ 1 kg of electrons
= $\frac{1}{9.108 \times 6.023} \times 10^{8}$ mol

8. Number of atoms 558.6 g Fe (molar mass Fe = 55.86 g mol^{-1}) is
(a) twice that in 60 g carbon **(b)** 6.023×10^{22}
(c) half that 8 g He **(d)** $558.6 \times 6.023 \times 10^{23}$

Sol. **(a)** 558.5 g Fe = 10 moles
= $10 \times 6.023 \times 10^{23}$ atoms
60 g C = 5 moles = $5 \times 6.023 \times 10^{23}$ atoms

9. The maximum number of molecules if present in
(a) 15 L of H gas S.T.P. **(b)** 5 L of N_2 gas at S.T.P.
(c) 0.5 g of H_2 gas **(d)** 10 g of O_2 gas

Sol. **(a)** At STP, 22.4 L of any gas = 6.02×10^{23} molecules

$\therefore$ 15 L H_2 = $\frac{6.02\times10^{23}}{22.4}\times15$ = 4.03 × 10^{23}

5 L N_2 = $\frac{6.02\times10^{23}}{22.4}\times5$ = 1.344 × 10^{23}

2g H_2 = 6.02 × 10^{23} molecules

$\therefore$ 0.5 g H_2 = $\frac{6.02\times10^{23}}{22.4}\times0.5$ molecules

32 g O_2 = 6.02 × 10^{23} molecules

$\therefore$ 10 g O_2 = $\frac{6.02\times10^{23}}{32}$ × 10 molecules

= 1.88 × 10^{23} molecules

10. Number of water molecules in the drop of water, if 1 ml of water has 20 drops and A is Avogadro's number, is

(a) 0.5 A / 18
(b) 0.05 A
(c) 0.5 A
(d) 0.05 A / 18

Sol. **(d)** 1 drop of water = $\frac{1}{20}$ ml = $\frac{1}{20}$ g $\left(d_{H_2O}=1gml^{-1}\right)$

18 g of water = A molecules

$\therefore$ $\frac{1}{20}$g of water = $\frac{A}{18}\times\frac{1}{20}=\frac{0.05A}{18}$ molecules

Unsolved Conceptual Problems

1. The temperature difference between the two bodies is 5°C. What will be the difference in Kelvin ?
2. Can an element have equivalent mass more than its atomic mass?
3. Write down the relation that will show the relationship between normality and molarity.
4. A and B are ground together in a pestle and mortar. A large amount of heat is evolved and a substance C having properties different from A and B is formed. Is C an element, compound or mixture?
5. State Avogadro's law?
6. What is the relation between vapour density and molecular weight of a gas?
7. Why element have fractional atomic masses?
8. Between one gram of water and one gram of methanol, which has more number of molecules?
9. Give empirical formula of glucose and hydrogen peroxide.
10. What is meant by a limiting reagent?
11. Define the terms : (i) element (ii) compound (iii) mixture, and give two examples of each.
12. State the law of multiple proportions. How the oxides of nitrogen illustrate the law of multiple proportions?
13. Define the terms:

(i) Atomic mass
(ii) Gram atomic mass
(iii) Molecular mass
(iv) Gram molecular mass

Practice Problems

1. How many molecules and atoms of sulphur are present in 0.1 mole of S_8 molecules.
2. Calculate the number of moles of iodine in a sample containing 1 × 10^{22} molecules.
3. Calculate the number of molecules and atoms in 11.2 ℓ of O_2 at NTP.
4. A 0.005 cm thick coating of copper is deposited on a plate of 0.5 m^2 total area. Calculate the number of copper atoms deposited on the plate (density of copper = 7.2 g cm^{-3}, at. mass = 63.5)
5. Calculate the number of molecules present in a spherical drop of water having a radius 1 mm if density of water is 1g/cm^3.
6. How many moles of methane are required to produce 22g of CO_2 (g) after combustions.
7. Calculate the volume of oxygen at NTP that would be required to convert 5.2 ℓ of carbon monoxide to carbon dioxide.
8. If 20 g of $CaCO_3$ is treated with 20 g of HCl, how many grams of CO_2 can be produced according to reaction :
$CaCO_{3\,(s)} + 2HCl_{(aq)} \longrightarrow CaCl_{2\,(aq)} + H_2O_{(L)} + CO_{2\,(g)}$

9. How many moles and grams of sodium chloride are present in 250 mL of 0.50 m NaCl solution.

10. The density of 3M solution of NaCl is 1.25 g/ml. Calculate the molality of the solution.

Unsolved Objective Problems

1. In compound A, 1.0 g nitrogen combines with 0.57 g oxygen. In compound B, 2.0 g nitrogen unite with 2.24 g oxygen and in compound C, 3.0 g nitrogen combine with 5.11 g oxygen. These results obey the law of :
(a) multiple proportions **(b)** constant proportions
(c) reciprocal proportions **(d)** none of these

2. The molecular mass of chloride, MCl, is 74.5. The equivalent mass of the metal M will be :
(a) 39.0 **(b)** 74.5
(c) 110.0 **(d)** 35.5

3. The oxide of an element possesses the molecular formula, M_2O_3. If the equivalent mass of the metal is 9, the atomic mass of the metal will be :
(a) 27 **(b)** 18
(c) 9 **(d)** 4.5

4. The percentage of available chlorine in a sample of bleaching powder, $CaOCl_2.2H_2O$, is :
(a) 30 **(b)** 50
(c) 43.5 **(d)** 59.9

5. A sample of PCl_3 contains 1.4 moles of the substance. How many atoms are there in the sample?
(a) 4 **(b)** 5.6
(c) 8.431×10^{23} **(d)** 3.372×10^{24}

6. 5.85 g NaCl is dissolved in 1 L water. The number of ions of Na^+ and Cl^- in 1 mL of this solution will be:
(a) 6.02×10^{19} **(b)** 1.2×10^{22}
(c) 1.2×10^{20} **(d)** 6.02×10^{20}

7. Number of electrons present in 3.6 mg of NH_4^+ are :
(a) 1.2×10^{21} **(b)** 1.2×10^{20}
(c) 1.2×10^{22} **(d)** 2×10^{-3}

8. Calculate the correct number of significant figures in $4.26 - \dfrac{15.635}{5.0}$
(a) 1.12 **(b)** 1.103
(c) 1.1 **(d)** 1.13

9. Which has the highest mass?
(a) 50 g of iron **(b)** 5 moles of N_2
(c) 0.1 g atom of Ag **(d)** 10^{23} atoms of carbon

10. The equivalent weight of $Zn(OH)_2$ in the following reaction is equal to

$$Zn(OH)_2 + HNO_3 \longrightarrow Zn(OH)(NO_3) + H_2O$$

(a) $\dfrac{\text{Molecular weight}}{1}$ **(b)** $\dfrac{\text{Molecular weight}}{2}$
(c) $2 \times$ Molecular weight **(d)** $3 \times$ Molecular weight

11. A compound contains atoms of three elements A, B and C. If the oxidation number of A is +2, B is + 5 and that of C is – 2, the possible formula of the compound is :
(a) $A_3(BC_4)_2$ **(b)** $A_3(B_4C)_2$
(c) ABC_2 **(d)** $A_2(BC_3)_2$

12. Among the species given below which have same mass ?
(i) 0.1 g-molecule of sulphur dioxide (ii) 0.1 g-molecule of N_2O
(iii) 0.1 g-molecule of dry ice (iv) Avogadro number of CO molecules
(a) (i), (ii) **(b)** (ii), (iii)
(c) (i), (iii), (iv) **(d)** All have same mass

Answers

Unsolved Problems

1. 17.0 g
2. % of O_2 in both case=20%
3. Do yourself
4. Do yourself
5. Do yourself

6. 150
7. 32
8. 87%
9. M_2O
10. 30
11. (i) 2.657 g (ii) 0.1 moles and 6.023×10^{22} molecules (iii) 6g
(iv) 34.2 g (v) 2240 L
12. All contain equal number of atoms
13. 3×10^{24} protons
14. 3200 g S and 32.65 g S
15. $NH_3 > CO_2 > O_2$
16. (i) 0.669 (ii) 0.223 (iii) 0.25 (iv) 0.312
17. a. 2.989×10^{-23} ml
b. 3.85×10^{-8} cm
18. a – 4 moles
b – 2 moles
c – 1 moles
d – 0.33 moles
e –1.786×10^{4} moles
g – 5.458×10^{-6} moles.
19. (a) 4N, 2M (b) 0.5 N, 0.5 M (c) 0.2 M, 0.4 N
20. (a) 0.0167 (b) 0.267 (c) 0.1 (d) 9.9
21. 0.5978 g
22. 1.435 g
23. 10.42 m
24. 81.67 g
25. 46.26 g
26. 112 cm^3
27. (a) O_2 (b) 64 g (c) 2 moles (d) 72 g of carbon should be taken
28. 8.8 g
29. 32.39% Na, 22.54% S, 45.07% O
30. K = 31.84%, Cl = 28.98%, O = 39.18%
31. $C_3H_5NO_2$

Practice Problems

1. 6.02×10^{22} molecule, 4.81×10^{23} atoms
2. 0.0166 mol
3. 3.01×10^{23} molecule, 6.02×10^{23} atom
4. 1.71×10^{24} atoms
5. 1.40×10^{20}
6. 0.5 mol
7. 2.6 ℓ
8. 8..80 g
9. 125 mol, 7.312 g
10. 2.79

Unsolved Objective Problems

1	A	5	D	9	B
2	A	6	C	10	A
3	A	7	A	11	A
4	C	8	C	12	B