LESSON

SOME BASIC CONCEPTS OF CHEMISTRY

Introduction

Chemistry is the scince of molecules and their transformation. It deals with the compositions, structure and properties of matter. These aspects can be best described and understood in terms of basic constituents of matter: atoms and molecules.
Their chapter also includes how physical properties of matter can be quantitatively described using numerical values with suitable units.

Importance of Chemistry

- Chemistry is the branch of science that studies the composition, properties and interaction of matter. Chemists are interested in knowing how chemical transformations occur. Chemistry plays a central role in science and is often intertwined with other branches of science like physics, biology, geology etc. Chemistry also plays an important role in daily life.
- Chemical principles are important in diverse areas, such as: weather patterns, functioning of brain and operation of a computer. 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.
- 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. 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.
- With a better understanding of chemical principles it has now become possible to design and synthesize new materials having specific magnetic, electric and optical properties. This has lead to the production of superconducting ceramics, conducting polymers, optical fibres and large scale miniaturization of solid state devices.
- Safer alternatives to environmentally hazardous refrigerants like CFCs (chlorofluorocarbons), responsible for ozone depletion in the stratosphere, have been successfully synthesised. However, many big environmental problems continue to be matters of grave concern to the chemists. One such problem is the management of the Green House gases like methane, carbon dioxide etc.
- Understanding of bio-chemical processes, use of enzymes for large-scale production of chemicals and synthesis of new exotic materials are some of the intellectual challenges for the future generation of chemists. A developing country like India needs talented and creative chemists for accepting such challenges.

Nature of Matter

Anything which has mass and occupies space is called matter.
Everything around us, for example, book, pen, pencil, water, air, all living beings etc. are composed of matter.
You are also aware that matter can exist in three physical states viz. solid, liquid and gas.
In solids, these particles are held very close to each other in an orderly fashion and there is not much freedom of movement. In liquids, the particles are close to each other but they can move around.

However, in gases, the particles are far apart as compared to those present in solid or liquid states and their movement is easy and fast.

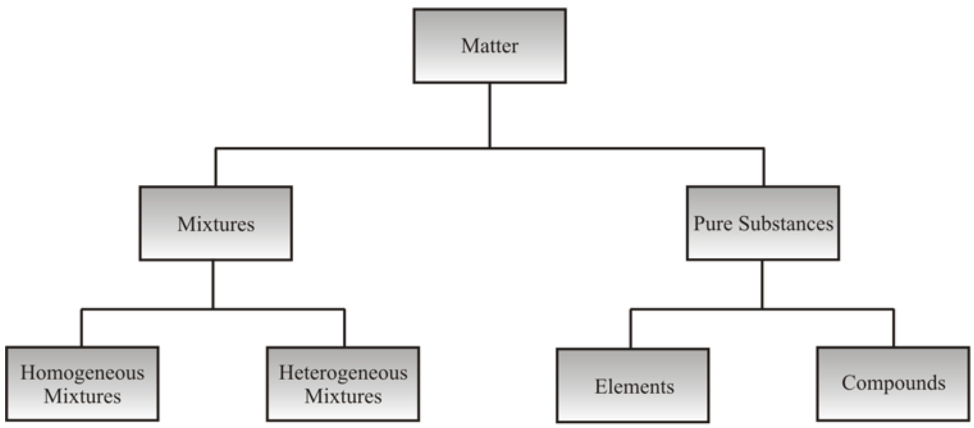
Different states of matter exhibit the following characteristics: (i) Solids have definite volume and definite shape. (ii) Liquids have definite volume but not the definite shape. They take the shape of the container in which they are placed. (iii) Gases have neither definite volume nor definite shape. They completely occupy the container in which they are placed.

These three states of matter are interconvertible by changing the conditions of temperature and pressure.

$$\text{Solid} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{liquid} \rightleftharpoons \text{Gas}$$

Classification of Matter

At the macroscopic or bulk level, matter can be classified as mixtures or pure substances. These can be further sub-divided as shown in Fig.

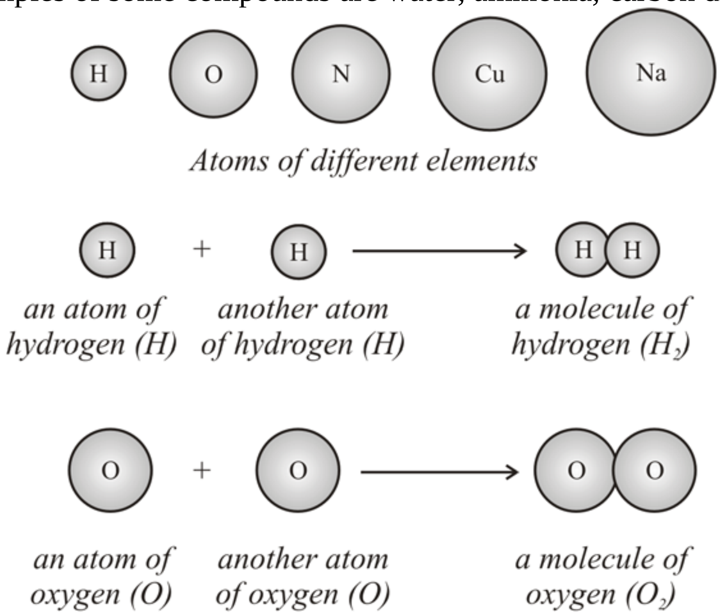


Classification of matter

- 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. In contrast to this, in 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.
- The components of a mixture can be separated by using physical methods such as simple hand picking, filtration, crystallisation, distillation etc.
- 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. Also, the constituents of pure substances cannot be separated by simple physical methods.
- Pure substances can be further classified into elements and compounds. An 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. 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.

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



A representation of atoms and molecules

- 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. Also, the properties of a compound are different from those of its constituent elements. For example, hydrogen and oxygen are gases whereas the compound formed by their combination i.e., water is a liquid.

Properties of Matter And Their Measurement

Every substance has unique or characteristic properties. These properties can be classified into two categories–physical properties and chemical properties. 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. The measurement or observation of chemical properties require a chemical change to occur. The examples of chemical properties are characteristic reactions of different substances; these include acidity or basicity, combustibility etc.

Different System of Measurement:

Two different system of measurement, i.e. the English System and the Metric System were being used in different parts of the world. The metric system which originated in France in late eighteenth century, was more convenient as it was based on the decimal system. The need of a common standard system was being felt by the scientific community. Such a system was established in 1960.

The International System of Units (S.I.):

- The International System of Units (in Freanch Le System International d'Unites – abbreviated as SI) was established by the 11^{th} General Conference on Weights and Measures (CGPM from Conference Gnerale des Poids at Measures).

- The SI system has seven base units and they are listed in Table. These units pertain to the seven fundamental scientific quantities. The other physical quantities such as peed, volume, density etc. can be derived from these quantities.

Table: Base Physical Quantities and their Units

Base Physical Quantity	Symbol for Quantity	Name of SI Unit	Symbol for SI Unit
Length Mass Time Electric current Thermodynamic temperature Amount of substance Luminous intensity	ℓ m t I T n I_v	metre kilogram second ampere kelvin mole candela	m kg s A K mol cd

- The definitions of the SI base units are given in Table.

Table: Definitions of SI Base Units

Unit of length	**metre**	The metre is the length of the path traveled by light in vacuum during a time interval of 1/299 792 458 a second.
Unit of Mass	**kilogram**	The kilogram is the unit of mass; it is equal to the mass of the international prototype of the kilogram.
Unit of time	**second**	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 the caesium-133 atom.
Unit of electric current	**ampere**	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.
Unit of thermodynamic temperature	**kelvin**	The kelvin, unit of thermodynamic temperature, is the fraction 1/273.16 of the thermodynamic temperature of the triple point of water.
Unit of amount of substance	**mole**	1. The mole is the amount of substance of a system which contains as many elementary entities as there are atoms in 0.012 kilogram of carbon-12; its system is “mol.” 2. When the mole is used, the elementary entities must be specified and may be

Unit of luminous intensity	**candela**	atoms, molecules, ions, electrons, other particles, or specified groups of such particles. The candela is the luminous intensity, in a given direction, of a source that emits monochromatic radiation of frequency 540 $\times 10^{12}$ hertz and that has a radiant intensity in that direction of 1/683 watt per steradian.

- The SI system allows the use of prefixed to indicate the multiples or submultiples of a unit. These prefixes are listed in Table.

Table: Prefixes used in the SI System

Multiple	Prefix	Symbol
10^{-24}	yocto	y
10^{-21}	zepto	z
10^{-18}	atto	a
10^{-15}	femto	f
10^{-12}	pico	p
10^{-9}	nano	n
10^{-6}	micro	μ
10^{-3}	milli	m
10^{-2}	centi	c
10^{-1}	deci	d
10	deca	da
10^{2}	hecto	h
10^{3}	kilo	k
10^{6}	mega	M
10^{9}	giga	G
10^{12}	tera	T
10^{15}	peta	P
10^{18}	exa	E
10^{21}	zeta	Z
10^{24}	yotta	Y

Mass and Weight

- 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.
- The mass of a substance can be determined very accurately in the laboratory by using an analytical balance.
- The SI unit of mass as given in Table is kilogram. However, its fraction gram (1 kg = 1000 g), is used in laboratories due to the smaller amounts of chemicals used in chemical reactions.

Volume

- Volume has the units of $(length)^3$. So in SI system, volume has units of m^3. But again, in chemistry laboratories, smaller volumes are used. Hence, volume is often denoted in cm^3 or dm^3 units.
- A common unit, litre (L) which is not an SI unit, is used for measurement of volume of liquids.

$$1 \text{ L} = 1000 \text{ mL}, 1000 \text{ cm}^3 = 1 \text{ dm}^3$$

- In the laboratory, volume of liquids or solutions can be measured by graduated cylinder, burette, pipette etc. A volumetric flask is used to prepare a known volume of a solution.

Density

Density of a substance is its amount of mass per unit volume. So SI units o density can be obtained as follows:

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

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

This unit is quite large and a chemist often expresses density in g cm^{-3}.

Temperature

- There are three common scales to measure temperature – ºC (degree Celsius), ºF (degree Fahrenheit) and K (kelvin). Here, K is the SI unit.
- Generally, the thermometer with Celsius scale are calibrated from 0º to 100º where these two temperatures are the freezing point and the boiling point of water respectively. The Fahrenheit scale is represented between 32º to 212º.
- The temperatures on two scales are related to each other by the following relationship:

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

The kelvin scale is related to celsius scale as follows:

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

Uncertainty in Measurement

1. Scientific Notation

- As chemistry is the study of atoms and molecules which have extremely low masses and are present in extremely large numbers, a chemist has to deal with numbers as large as 602, 200,000,000,000,000,000,000 for the molecules of 2 g of hydrogen gas or as small as 0.00000000000000000000000166 g mass of a H atom. Similarly other constants such as Planck's constant, speed of light, charges on particles etc., involve numbers of the above magnitude. It may look funny for a moment to write or count numbers involving so many zeros but it offers a real challenge to do simple mathematical operations of addition, subtraction, multiplication or division with such numbers.
- This problem is solved by using scientific notation for such 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}.
- Now, for performing mathematical operations on numbers expressed in scientific notations, the following points are to be kept in mind.

Multiplication and Division

These two operations follow the same rules which are there for exponential numbers, i.e.

$$(5.6 \times 10^{5}) \times (6.9 \times 10^{8}) = (5.6 \times 6.9)\ (10^{5+8})$$
$$= (5.6 \times 6.9) \times 10^{13}$$
$$= 39.64 \times 10^{13}$$

$$(9.8 \times 10^{-2}) \times (2.5 \times 10^{-6}) = (9.8 \times 2/5)\ (10^{-2+(-6)})$$
$$= (9.8 \times 2.5)\ (10^{-2-6})$$
$$= 24.50 \times 10^{-8}$$

$$\frac{2.7\times10^{-3}}{5.5\times10^{4}} = (2.7 \div 5.5)(10^{-3-4}) = 0.4909\times10^{-7}$$

Addition and Subtraction

For these two operations, first the numbers are written in such a way that they have same exponent. After that, the coefficient are added or subtracted as the case may be.

Thus, for adding 6.65×10^{4} and 8.95×10^{3}, $6.65 \times 10^{4} + 0.895 \times 10^{4}$ exponent is made same for both the numbers.

Then, these numbers can be added as follows $(6.65 + 0.895) \times 10^{4} = 7.545 \times 10^{4}$

Similarly, the subtraction of two numbers can be done as shown below:

$$2.5 \times 10^{-2} - 4.8 \times 10^{-3}$$
$$= (2.5 \times 10^{-2}) - (0.48 \times 10^{-2})$$
$$= (2.5 - 0.48) \times 10^{-2} = 2.02 \times 10^{-2}$$

2. <u>Significant Figures</u>

- Precision refers to the closeness of various measurements for the same quantity. However, accuracy is the agreement of a particular value to the true value of the result.
- The uncertainty in the experimental or the calculated values is indicated by mentioning the number of significant figures. Significant figures are meaningful digits which are known with certainty.
- Rules for determining the number of significant figures. These are stated below:

  (1) All non-zero digits are significant. For example in 285 cm, there are three significant figures and in 0.25 mL, there are two significant figures.

  (2) Zeros preceding to first non-zero digit are not significant. Such zero indicates the position of decimal point. Thus, 0.03 has one significant figure and 0.0052 has two significant figures.

  (3) Zeros between two non-zero digits are significant. Thus, 2.005 has four significant figures.

  (4) Zeros at the end or right of a number are significant provided they are on the right side of the decimal point. For example, 0.200 g has three significant figures. But, if otherwise, the terminal zeros are not significant if there is no decimal point. For example, 100 has only one significant figure, but 100. has three significant figures and 100.0 has four significant figures. Such numbers are better represented in scientific notation. We can express the number 100 as 1×10^{2} for one significant figure, 1.0×10^{2} for two significant figures and 1.00×10^{2} for three significant figures.

  (5) Counting numbers of objects, for example, 2 balls or 20 eggs, have infinite significant figures as these are exact numbers and can be represented by writing infinite number of zeros after placing a decimal i.e., 2 = 2.000000 or 20 = 20.000000 In numbers written in scientific notation, all digits are significant e.g., 4.01×10^{2} has three significant figures, and 8.256×10^{-3} has four significant figures.

  Example: The result cannot have more digits to the right of the decimal point than either of the original numbers.

$$\begin{array}{r} 12.11 \\ 18.0 \\ 1.012 \\ \hline \mathbf{31.122} \\ \hline \end{array}$$

Here, 18.0 has only one digit after the decimal point and the result should be reported only up to one digit after the decimal point which is 31.1.

Following points for rounding off the numbers:

1. If the rightmost digit to be removed is more than 5, the preceding number is increased by one. for example, 1.386
2. If the rightmost digit to be removed is less than 5, the preceding number is not changed, for example, 4.334 if 4 is to be removed then the result is rounded upto 4.33.
3. If the rightmost digit to be removed is 5, then the preceding number is not changed if it is an even number but it is increased by one if it is an odd number. For example, if 6.35 is to be rounded by removing 5, we have to increase 3 to 4 giving 6.4 as the result. However, if 6.25 is to be rounded off it is rounded off to 6.2.

3. Dimensional Analysis

Often while calculating, there is a need to convert units from one system to other. The method used to accomplish this is called factor label method or unit factor method or dimensional analysis. This is illustrated below.

Example

A piece of metal is 3 inch (represented by in) long. What is its length in cm?

We know that 1 in = 2.54 cm

From this equivalence, we can write

$$\frac{1\text{ in}}{2.54\text{ cm}} = 1 = \frac{2.54\text{ cm}}{1\text{ in}}$$

thus $\frac{1\text{ in}}{2.54\text{ cm}}$ equals 1 and $\frac{2.54\text{ cm}}{1\text{ in}}$ also equals 1. Both of these are called unit factors. If some number is multiplied by these unit factors (i.e. 1), it will not be affected otherwise.

Law of Chemical Combinations

The combination of elements to form compounds is governed by the following five basic laws.

1. Law of Conservation of Mass

It states that matter can neither be created nor destroyed. This law was put forth by Antoine Lavoisier in 1789.

2. Law of Definite Proportions

This law was given by, a French chemist, Joseph Proust. He stated that a given compound always contains exactly the same proportion of elements by weight.

Two samples of cupric carbonate – one of which was of natural origin and the other was synthetic one. He found that the composition of elements present in it was same for both the samples as shown below:

	% of copper	**% of oxygen**	**% of carbon**
Natural Sample	51.35	9.74	38.91
Synthetic Sample	51.35	9.74	38.91

3. Law of Multiple Proportions

This law was proposed by Dalton in 1803. According to this law, if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers.

For example, hydrogen combines with oxygen to form two compounds, namely, water and hydrogen peroxide.

Hydrogen +	Oxygen $\rightarrow$	Water
2g	16g	18g

Hydrogen +	Oxygen $\rightarrow$	Hydrogen Peroxide
2g	32g	34g

Here, the masses of oxygen (i.e. 16 g and 32 g) which combine with a fixed mass of hydrogen (2g) bear a simple ratio, i.e. 16:32 or 1: 2.

4. Gay Lussac's Law of Gaseous Volumes

This law was given by Gay Lussac in 1808. 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.

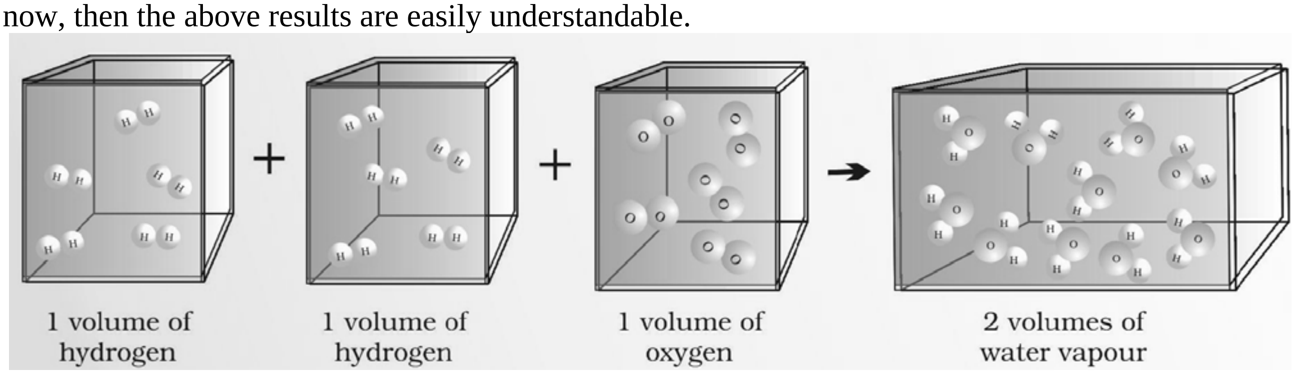
Thus, 100 mL of hydrogen combine with 50 mL of oxygen to give 100 mL of water vapour.

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.

5. Avogadro Law

In 1811, Avogadro proposed that equal volumes of gases at the same temperature and pressure should contain equal number of molecules. Avogadro made a distinction between atoms and molecules. If we consider again the reaction of hydrogen and oxygen to produce water, we see that two volumes of hydrogen combine with one volume of oxygen to give two volumes of water without leaving any unreacted oxygen. In fact, Avogadro could explain the above result by considering the molecules to be polyatomic. If hydrogen and oxygen were considered as diatomic as recognised now, then the above results are easily understandable.



Dalton's Atomic Energy

In 1808, Dalton published 'A New System of Chemical Philosophy' in which he proposed the following:

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 reorganization of atoms. These are neither created nor destroyed in a chemical reaction.

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

Atomic and Molecular Masses

Atomic Mass

The present system of atomic masses is based on carbon – 12 as the standard and has been agreed upon in 1961. Here, Carbon – 12 is one of the isotopes of carbon and can be represented as ^{12}C. In this system, ^{12}C is assigned a mass of exactly 12 atomic mass unit (amu) and masses of all other atoms are given relative to this standard.

One atomic mass unit is defined as a mass exactly equal to one twelfth the mass of one carbon – 12 atom.

And 1 amu $= 1.66056 \times 10^{-24}$ g

Mass of an atom of hydrogen $= 1.6736 \times 10^{-24}$ g

Thus, in terms of amu, the mass of hydrogen atom

$$= \frac{1.6736 \times 10^{-24}\,g}{1.66056 \times 10^{-24}\,g}$$

$= 1.0078$ amu

$= 1.0080$ amu

Similarly, the mass of oxygen – 16 (^{16}O) atom would be 15.995 amu.

Today, 'amu' has been replaced by 'u' which is known as unified mass.

Average Atomic Mass

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. For example, carbon has the following three isotopes with relative abundances and masses as shown against each of them.

Isotope	Relative Abundance (%)	Atomic Mass (amu)
^{12}C	98.892	12
^{13}C	1.108	13.00335
^{14}C	2×10^{-10}	14.00317

From the above data, the average atomic mass of carbon will come out to be:

$(0.98892)(12\ u) + (0.01108)(13.00335\ u) + (2 \times 10^{-12})(14.00317\ u) = 12.011\ u$

Similarly, average atomic masses for other elements can be calculated. In the periodic table of elements, the atomic masses mentioned for different elements actually represented their average atomic masses.

Molecular Mass

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. For example, molecular mass of methane which contains one carbon atom and four hydrogen atoms can be obtained as follows:

Molecular mass of methane,

$(CH_4) = (12.011\ u) + 4\ (1.008\ u) = 16.043\ u$

Formula Mass
Some substances such as sodium chloride do not contain discrete molecules as their constituent units.
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.
Thus, formula mass of sodium chloride = atomic mass of sodium + atomic mass of chlorine
= 23.0 u + 35.5 u = 58.5 u

Mole Concept and Molar Masses

- Just as we denote one dozen for 12 items, score for 20 items, gross for 144 items, we use the idea of mole to count entities at the microscopic level (i.e. atoms/molecules/ particles, electrons, ions, etc).
- In SI system, mole (symbol, mol) was introduced as seventh base quantity for the amount of a substance.
- 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.
- The mass of a carbon –12 atom was determined by a mass spectrometer and found to be equal to 1.992648×10^{-23} g.
- One mole of carbon weighs 12 g, the number of atoms in it is equal to:

$$\frac{12\ \text{g/mol}\ ^{12}\text{C}}{1.992648 \times 10^{-23}\ \text{g}/^{12}\text{C atom}}$$

$$= 6.0221367 \times 10^{23}\ \text{atoms/mol}$$

- It is known as 'Avogadro constant', denoted by N_A in honour of Amedeo Avogadro.
- Therefore, say that 1 mol of hydrogen atoms = 6.022×10^{23} atoms.
- 1 mol of water molecules = 6.022×10^{23} water molecules.
- 1 mol of sodium chloride = 6.022×10^{23} formula units of sodium chloride.
- The mass of one mole of a substance in grams is called its molar mass. The molar mass in grams is numerically equal to atomic/molecular/ formula mass in u.
  Molar mass of water = 18.02 g
  Molar mass of sodium chloride = 58.5 g

Percentage Composition

Let us understand it by taking the example of water (H_2O). Since water contains hydrogen and oxygen, the percentage composition of both these elements can be calculated as follows:

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

Molar mass of water = 18.02 g

$$\text{Mass \% of hydrogen} = \frac{2 \times 1.008}{18.02} \times 100 = 11.18$$

$$\text{Mass \% of oxygen} = \frac{16.00}{18.02} \times 100 = 88.79$$

Empirical Formula for Molecular Formula

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.

If the mass per cent of various elements present in a compound is known, its empirical formula can be determined.

Example:

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?

Solution:

Step 1. Conversion of mass per cent to grams.

Since we are having mass per cent, it is convenient to use 100 g of the compound as the starting material. Thus, in the 100 g sample of the above compound, 4.07g hydrogen is present, 24.27g carbon is present and 71.65 g chlorine is present.

Step 2. Convert into number moles of each element

Divide the masses obtained above by respective atomic masses of various elements.

Moles of hydrogen = $\frac{4.07\text{ g}}{1.008\text{ g}} = 4.04$

Moles of carbon = $\frac{24.27\text{ g}}{12.01\text{ g}} = 2.021$

Moles of chlorine = $\frac{71.65\text{ g}}{35.453\text{ g}} = 2.021$

Step 3. Divide the mole value obtained above by the smallest number

Since 2.021 is smallest value, division by it gives a ratio of 2:1:1 for H:C:Cl . In case the ratios are not whole numbers, then they may be converted into whole number by multiplying by the suitable coefficient.

Step 4. Write empirical formula by mentioning the numbers after writing the symbols of respective elements.

CH_2Cl is, thus, the empirical formula of the above compound.

Step 5. Writing molecular formula

(a) Determine empirical formula mass Add the atomic masses of various atoms present in the empirical formula.

For CH_2Cl, empirical formula mass is

$$12.01 + 2 \times 1.008 + 35.453 = 49.48\text{ g}$$

(b) Divide Molar mass by empirical formula mass

$$\frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{98.96\text{ g}}{49.48\text{ g}} = 2 = (n)$$

(c) Multiply empirical formula by n obtained above to get the molecular formula

Empirical formula = CH_2Cl, n = 2. Hence molecular formula is $C_2H_4Cl_2$.

Stoichiometry and Stoichiometric Calculations

The word ‘stoichiometry’ is derived from two Greek words - stoicheion (meaning element) and metron (meaning measure). Stoichiometry, thus, deals with the calculation of masses (sometimes volumes also) of the reactants and the products involved in a chemical reaction.

$$CH_4\text{ (g)} + 2O_2\text{ (g)} \rightarrow CO_2\text{ (gas)} + 2H_2O\text{ (g)}$$

Thus, according to the above chemical reaction,

- One mole of CH_4(g) reacts with two moles of O_2(g) to give one mole of CO_2(g) and two moles of H_2O(g).
- One molecule of CH_4(g) reacts with 2 molecules of O_2(g) to give one molecule of CO_2(g) and 2 molecules of H_2O(g)
- 22.4 L of CH_4(g) reacts with 44.8 L of O_2 (g) to give 22.4 L of CO_2 (g) and 44.8 L of H_2O(g)
- 16 g of CH_4 (g) reacts with 2 × 32 g of O_2 (g) to give 44 g of CO_2 (g) and 2 × 18 g of H_2O (g).

From these relationships, the given data can be interconverted as follows:

$$\text{mass} \rightleftharpoons \text{moles} \rightleftharpoons \text{no. of molecules}$$

$$\frac{\text{Mass}}{\text{Volume}} = \text{Density}$$

Limiting Reagent

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.

In performing stoichiometric calculations, this aspect is also to be kept in mind.

Reactions in Solution

It is important to understand as how the amount of substance is expressed when it is present in the form of a solution. The concentration of a solution or the amount of substance present in its given volume can be expressed in any of the following ways.

(a) Mass per cent or weight per cent (w/w %)
(b) Mole fraction
(c) Molarity
(d) Molality

Let us now study each one of them in detail.

(a) Mass per cent

It is obtained by using the following relation:

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

(b) Mole Fraction

It is the ratio of number of moles of a particular component to the total number of moles of the solution. If a substance 'A' dissolves in substance 'B' and their number of moles are n_A and n_B respectively; then the mole fractions of A and B are given as

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

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

$$= \frac{n_B}{n_A + n_B}$$

(c) Molarity

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

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

Suppose we have 1 M solution of a substance, say NaOH and we want to prepare a 0.2 M solution from it.

For such calculations, a general formula, $M_1 \times V_1 = M_2 \times V_2$ where M and V are molarity and volume respectively can be used.

(d) Molality

It is defined as the number of moles of solute present in 1 kg of solvent. It is denoted by m.

Thus, Molality (m) = $\frac{\text{No. of moles of solute}}{\text{Mass of solvent in kg}}$

domain E

SOLVED EXAMPLES

NCERT Solved Examples

NCERT 1: Calculate molecular mass of glucose $(C_6H_{12}O_6)$ molecule.

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

$= 6(12.011\,u) + 12(1.008\,u) + 6(16.00\,u)$

$= (72.066\,u) + (12.096\,u) + (96.00\,u)$

= 180. 162 u

NCERT 2: 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?

Solution: Step 1: Conversion of mass per cent to grams.

Since we are having mass per cent, it is convenient to use 100 g of the compound as the starting material. Thus, in the 100 g sample of the above compound, 4.07g hydrogen is present, 24.27 g carbon is present and 71.65 g chlorine is present.

Step 2: Convert into number moles of each element

Divide the masses obtained above by respective atomic masses of various elements.

Moles of hydrogen = $\dfrac{4.07}{1.008g} = 4.04$

Moles of carbon $= \dfrac{24.27g}{12.01g} = 2.021$

Moles of chlorine $= \dfrac{71.65g}{35.453g} = 2.021$

Step 3: Divide the mole value obtained above by the smallest number

Since 2.021 is smallest value, division by it gives a ratio of 2:1:1 for H:C:Cl.

In case the ratios are not whole numbers, then they may be converted into whole number by multiplying by the suitable coefficient.

Step 4: Write empirical formula by mentioning the numbers after writing the symbols of respective elements.

CH_2Cl is, thus, the empirical formula of the above compound.

Step 5: Writing molecular formula

(a) Determine empirical formula mass. Add the atomic masses of various atoms present in the empirical formula.

For CH_2Cl, empirical formula mass is $12.01 + 2 \times 1.008 + 35.453$ = 49.48 g

(b) Divide Molar mass by empirical formula mass $\dfrac{\text{Molar mass}}{\text{Empirical formula mass}} = \dfrac{98.96g}{49.48g}$

$= 2 = (n)$

(c) Multiply empirical formula by n obtained above to get the molecular formula Empirical formula $= CH_2Cl, n = 2.$ Hence molecular formula is $C_2H_4Cl_2$.

NCERT 3: 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 to one mole.

(ii) From the above equation, 1 mol of CH_4 (g) gives 2 mol of H_2O (g).

2 mol of water (H_2O) $= 2 \times (2+16)$

$= 2 \times 18 = 36g$

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

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

$= 2 \times 18g\,H_2O = 36g\,H_2O$

NCERT 4: How many moles 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).

[$\because$ 1 mol $CO_2(g)$ is obtained from 1 mole of $CH_4(g)$]

mole of CO_2 (g)

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

= 0.5 mol CO_2(g)

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

NCERT 5: 50.0 kg of $N_2(g)$ and 10.0 kg of $H_2(g)$ are mixed to produce $NH_3(g)$. Calculate the $NH_3(g)$ formed. Identify the limiting reagent in the production of $NH_3(g)$ in this situation.

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

Calculation of moles:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

moles of N_2

$= 50.0kg\,N_2 \times \frac{1000g\,N_2}{1kg\,N_2} \times \frac{1mol\,N_2}{28.0gN_2}$

$= 17.86 \times 10^2$ mol

moles of H_2

$= 10.00kg\,H_2 \times \frac{100g\,H_2}{1kg\,H_2} \times \frac{1mol\,H_2}{2.016g\,H_2}$

$= 4.96 \times 10^3$ mol

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

$$17.86\times10^{2}\,\text{mol}\,N_2\times\frac{3\text{mol}\,H_2(g)}{1\text{mol}\,N_2(g)}$$

$$=5.36\times10^{3}\,\text{mol}\,H_2$$

But we have Only $4.96\times10^{3}\,\text{mol}$ 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.,

$4.96\times10^{3}\,\text{mol}$

Since 3 mol $H_2(g)$ gives 2 mol $NH_3(g)$

$$4.96\times10^{3}\,\text{mol}\,H_2(g)\times\frac{2\,\text{mol}\,NH_3(g)}{3\,\text{mol}\,H_2(g)}$$

$$=3.30\times10^{3}\,\text{mol}\,NH_3(g)$$

$3.30\times10^{3}\,\text{mol}\,NH_3(g)$ is obtained.

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

1 mol $NH_3(g)=17.0\text{g}\,NH_3(g)$

$$3.30\times10^{3}\,\text{mol}\,NH_3(g)\times\frac{17.0\text{g}NH_3(g)}{1\text{mol}\,NH_3(g)}$$

$$=3.30\times10^{3}\times17\text{g}\,NH_3(g)$$

$$=56.1\times10^{3}\text{g}\,NH_3$$

$$=56.1\text{kg}\,NH_3$$

NCERT 6: A solution is prepared by adding 2 g of a substance A to 18 g of water. Calculate the mass percent of the solute.

Solution: Mass per cent of $A=\dfrac{\text{Mass of A}}{\text{Mass of solution}}\times100$

$$=\frac{2\text{g}}{2\text{g of A}+18\text{g of water}}\times100$$

$$=\frac{2\text{g}}{20\text{g}}\times100$$

= 10%

NCERT 7: 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)

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

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

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

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

$$= 0.4\,\text{M}$$

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

NCERT 8: 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}

Mas of NaCl

In 1 L solution $= 3 \times 58.5 = 175.5\text{g}$

Mass of

1L solution $= 1000 \times 1.25 = 1250\,\text{g}$

(since density = $1.25\text{g}\,\text{mL}^{-1}$)

Mass of

water in solution = 1250 – 175.5

= 1074.5 g

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

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

$$= 2.79\,\text{m}$$

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 solution. Note that molality of a solution does not change with temperature since mass remains unaffected with temperature.

Additional Solved Example

Example 1: Vanadium metal is added to steel to impart strength. The density of Vanadium is 5.96 g/cm^3. Expressin S.I. unit (kg/m^3).

Solution: 5.96 g $\text{cm}^{-3} = \dfrac{5.96 \times 10^{-3}\,\text{kg}}{10^{-6}\,\text{m}^3} = 5960\,\text{kg}\,\text{m}^{-3}$.

Example 2: (a) When 4.2 g of $NaHCO_3$ is added to a solution of CH_3COOH weighing 10g, it is observed that 2.2 gm of CO_2 is released into atmosphere. The residue is found to weigh 12.0 g. Show that these observations are in agreement with the law of conservation of mass.

(b) If 6.3 g of $NaHCO_3$ re added to 15.0 g CH_3COOH solution, the residue is found to weight 18.0 g. What is the mass of CO_2 released in the reaction?

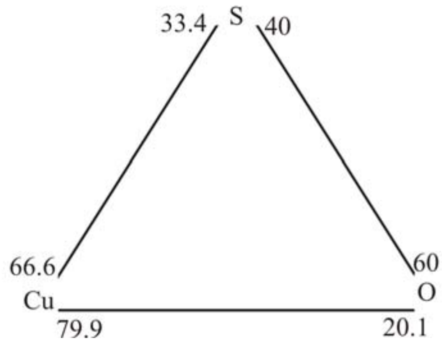
Solution: (a) According to law of conservation of mass the mass of reactants and products must be equal.

Thus,
Mass of reactants = 4.2 + 10 = 14.2 g
and mass of products = 2.2 + 12.0 = 14.2 g
Hence, the law of conservation of mass is proved.
(b) According to law of conservation of mass
mass of reactants = mass of products
$\therefore \quad 6.3 + 15.0 = 18.0 + x$
or $\quad x = 21.3 - 18.0 = 3.3g.$

Example 3: 1.375 g of cupric oxide was reduced by heating in a current of hydrogen and the weight of copper that remained was 1.098 g. In another experiment, 1.179 g of copper was dissolved in the nitric acid and the resulting copper nitrate converted into cupric oxide by ignition. The weight of cupric oxide formed was 1.476 g. Show that these results illustrate the law of constant composition.
Solution: First experiment:
Copper oxide = 1.375 g
Copper left = 1.098 g
$\therefore$ Oxygen present = 1.375 – 1.098 = 0.277 g
Hence % of oxygen in CuO $= \dfrac{0.277 \times 100}{1.375} = 20.14$
Second experiment:
Copper taken = 1.179 g
Copper oxide formed = 1.476 g
$\therefore$ Oxygen present = 1.476 – 1.179 g = 0.297 g
Hence % of oxygen of CuO $= \dfrac{0.297 \times 100}{1.476} = 20.12.$
Percentage of oxygen is the same in both the above cases, so the law of constant composition is illustrated.

Example 4: Carbon and oxygen are known to form two compounds. The carbon content in one of these is 42.9% while in the other it is 27.3%. Show that this data is in agreement with law of multiple proportions.
Solution: In compound I C = 42.9%; O = 57.1%
In compound II C = 27.3%; O = 72.7%
For a fixed mass of carbon (or parts) = 1 the ratio of oxygen in the two compounds would be
Oxygen in, I : II = $\dfrac{57.1}{42.9} : \dfrac{72.7}{27.3} = 1.33 : 2.66 = 1 : 2$
Since it is a simple whole number ratio, it illustrates law of multiple proportions.

Example 5: Copper sulphide contains 66.6% Cu, copper oxide contains 79.9% copper and sulphur trioxide contains 40% sulphur. Show that these data illustrates law of reciprocal proportions
Solution: In copper sulphide,
Cu : S mass ratio is 66.6 : 33.4



In sulphur trioxide,
O : S mass ratio is 60 : 40
Now in copper sulphide
33.4 parts of sulphur combines with Cu = 66.6 parts
40.0 parts of sulphur combines with Cu

$$= \frac{66.6 \times 40}{33.4} = 79.8 \text{ parts.}$$

Now, ratio of the masses of Cu and O which combine with same mass (40 parts) of sulphur separately is

79.8 : 60 …(i)

Cu : O ratio by mass in CuO is

79.9 : 20.1 …(ii)

Ratio I : Ratio II $= \frac{79.8}{60} \times \frac{20.1}{79.9} = 3:1$

Which is simple number ratio.
Hence, law of reciprocal proportion is proved.

Example 6: How many molecules of water of hydration are present in 252 mg of oxalic acid, $(H_2C_2O_4.2H_2O)$?

Solution: Gram molecular mass of $H_2C_2O_4.2H_2O$

= 126 g

Now, water molecules in 1 mol of oxalic acid

= 2 mol.

Or water molecules in 126 g of oxalic acid

$= 2 \times 6.023 \times 10^{23}$

$\therefore$ Water molecules in 252×10^{-3} g of oxalic acid

$$= \frac{2 \times 6.023 \times 10^{23} \times 252 \times 10^{-3}}{126} = 2.4 \times 10^{21}.$$

Example 7: What mass of zinc is required to produce hydrogen by reaction with HC*l* which is enough to produce 4 mol of ammonia according to the reactions

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$

$$3H_2 + N_2 \longrightarrow 2NH_3$$

Solution: The given equations are

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$
$$3H_2 + N_2 \longrightarrow 2NH_3$$

From the equations it is clear that
2 mol of NH_3 require = 3 mol of H_2,
and 1 mol H_2 require = 3 mol of Zn
or 3 mol of H_2 require = 3 mol of Zn
Thus, 2 mol of NH_3 require
$= 3\,\text{mol of Zn} = 3 \times 65\,\text{g of Zn}$
$\therefore$ 4 mol of NH_3 require
$$= \frac{3 \times 65}{2} \times 4 = 390\,\text{g of Zn.}$$

Example 8: How much magnesium sulphide can be obtained from 2.00 g of magnesium and 2.00 g of sulphure by the reaction $(H_2C_2O_4.2H_2O)$? Which is the limiting reagent? Calculate the amount of one of the reactants which remains unreacted.
Solution: First of all ach of the masses are expressed in moles:
$$2.00\,\text{g of mg} = \frac{2.00}{24.3} = 0.0824\,\text{moles of Mg}$$
$$2.00\text{ g of }S = \frac{2.00}{32.1} = 0.0624\,\text{moles of S}$$
From the equation, $Mg + S \rightarrow MgS,$ it follows that one mole of Mg reacts with one moles of S. We are given more moles of Mg than of S therefore, Mg is in excess and some of it will remain unreacted when the reaction is over. S is the limiting reagent and will control the amount of product. From the equation we note that one mole of S gives one mole of MgS, so 0.0624 mole of S will react with 0.0624 mole of Mg to form 0.0624 mole of MgS.
Molar mass of MgS = 56.4 g
$\therefore$ Mass of MgS formed
$= 0624 \times 56.4g = 3.52\,\text{g of MgS}$
Mole of Mg left unreacted
= 0.0824 – 0.0624 moles of Mg
$= 0.0200$ moles of Mg
Mass of Mg left unreacted
= Moles of Mg × molar mass of Mg
= $0.200 \times 24.3\,\text{g of Mg} =$ 0.489 of Mg.

Example 9: 2.746 gm of a compound gave on analysis 1.94 g of silver, 0.268 g of sulphur and 0.538 g of oxygen. Find the empirical formula of the compound. (At. masses: Ag = 108, S = 32, O = 16)
Solution: To calculate empirical formula:

Element	**Mass in (g)**	**At. Mass**	**No. of g-atom**	**Simple Ratio of g-atom**	**Simples Whole No. Rato**

Ag	1.94 g	108	$\frac{1.94}{108} = 0.0179$	$\frac{0.0179}{8.375\times10^{-3}} = 2.12$	2
S	0.268 g	32	$\frac{0.268}{32} = 8.375\times10^{-3}$	$\frac{8.375\times10^{-3}}{8.375\times10^{-3}} = 1$	1
O	0.538 g	16	$\frac{0.538}{16} = 0.0336$	$\frac{0.0336}{8.375\times10^{-3}} = 4.01$	4

Example 10: Caffeine, a stimulant found in coffee, tea, chocolate, and some medications, contains 49.48% carbon, 5.15% hydrogen, 28.87% nitrogen, and 16.49% oxygen by mass and has a molar mass of 194.2. Determine the molecular formula of caffeine.

Solution: We will first determine the mass of each element in 1 mol (194.2 g) of caffeine:

$$\frac{49.48\,\text{g C}}{100.0\,\text{g caffeine}}\times\frac{194.2\,\text{g}}{\text{mol}}=\frac{96.09\,\text{g C}}{\text{mol caffeine}}$$

$$\frac{5.15\,\text{g H}}{100.0\,\text{g caffeine}}\times\frac{194.2\,\text{g}}{\text{mol}}=\frac{10.0\,\text{g H}}{\text{mol caffeine}}$$

$$\frac{28.87\,\text{g N}}{100.0\,\text{g caffeine}}\times\frac{194.2\,\text{g}}{\text{mol}}=\frac{56.07\,\text{g N}}{\text{mol caffeine}}$$

$$\frac{16.49\,\text{g O}}{100.0\,\text{g caffeine}}\times\frac{194.2\,\text{g}}{\text{mol}}=\frac{32.02\,\text{g O}}{\text{mol caffeine}}$$

Now we will convert to moles,

C: $$\frac{96.09\,\text{g C}}{\text{mol caffeine}}\times\frac{1\,\text{mol C}}{12.011\,\text{g C}}=\frac{8.000\,\text{mol C}}{\text{mol caffeine}}$$

H: $$\frac{10.0\,\text{g H}}{\text{mol caffeine}}\times\frac{1\,\text{mol H}}{1.008\,\text{g H}}=\frac{9.92\,\text{mol H}}{\text{mol caffeine}}$$

N: $$\frac{56.07\,\text{g N}}{\text{mol caffeine}}\times\frac{1\,\text{mol N}}{14.01\,\text{g N}}=\frac{4.002\,\text{mol N}}{\text{mol caffeine}}$$

O: $$\frac{32.02\,\text{g O}}{\text{mol caffeine}}\times\frac{1\,\text{mol O}}{16.00\,\text{g O}}=\frac{2.001\,\text{mol O}}{\text{mol caffeine}}$$

Rounding the numbers to integers gives the molecular formula for caffeine: $C_8H_{10}N_4O_2$.

Example 11: Penicillin, the first of a now large number of antibiotics (antibacterial agents), was discovered accidentally by the Scottish bacteriologist Alexander Fleming in 1928, but he was never able to isolate it as a pure compound. This and similar antibiotics have saved millions of lives that might have been lost to infections. Penicillin F has the formula $C_{14}H_{20}N_2SO_4$. Compute the mass percentof each element.

Solution: The molar mass of penicillin F is computed as follows:

C: $$14\,\text{mol}\times12.011\frac{\text{g}}{\text{mol}}=168.15\,\text{g}$$

H: $20\,\text{mol}\times 1.008\,\dfrac{\text{g}}{\text{mol}} = 20.16\,\text{g}$

N: $2\text{ mol }\times 14.007\,\dfrac{\text{g}}{\text{mol}} = 28.014\,\text{g}$

S: $1\,\text{mol}\times 32.07\,\dfrac{\text{g}}{\text{mol}} = 32.07\text{g}$

O: $4\,\text{mol}\times 15.999\,\dfrac{\text{g}}{\text{mol}} = 63.996\,\text{g}$

Mass of 1 mol of $C_{14}H_{20}N_2SO_4 = 312.39\text{g}$

Mass percent of $C = \dfrac{168.15\,\text{g}\,C}{312.39\,\text{g}\,C_{14}H_{20}N_2SO_4}\times 100\% = 53.827\%$

Mass percent of $H = \dfrac{20.16\,\text{g}\,H}{312.39\,\text{g}\,C_{14}H_{20}N_2SO_4}\times 100\% = 6.453\%$

Mass percent of $N = \dfrac{28.014\,\text{g}\,N}{312.39\,\text{g}\,C_{14}H_{20}N_2SO_4}\times 100\% = 8.968\%$

Mass percent of $S = \dfrac{32.07\,\text{g}\,S}{312.39\,\text{g}\,C_{14}H_{20}N_2SO_4}\times 100\% = 10.27\%$

Mass percent of $O = \dfrac{63.996\,\text{g}\,S}{312.39\,\text{g}\,C_{14}H_{20}N_2SO_4}\times 100\% = 20.486\%$

PROBLEMS

Exercise I

Q.1 Calculate the molecular mass of the following.
(i) H_2O (ii) CO_2 (iii) CH_4

Q.2 Calculate the mass per cent of different elements present in sodium sulphate (Na_2SO_4).

Q.3 Determine the empirical formula of an oxide of iron which has 69.9% iron and 30.1% dioxygen by mass.

Q.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.

Q.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 82.0245 g mol^{-1}

Q.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%.

Q.7 How much copper can be obtained from 100 g of copper sulphate ($CuSO_4$) ?

Q.8 Determine the molecular formula of an oxide of iron in which the mass percent of iron and oxygen are 69.9 and 30.1 respectively.

Q.9 Calculate the atomic mass (average) of chlorine using the following data :

	% Natural Abundance	**Molar Mass**
$^{35}C\ell$	75.77	34.9689
$^{37}C\ell$	24.23	36.9659

Q.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.

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

Q.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?

Q.13 Pressure is determined as force per unit area of the surface. The SI unit of pressure, pascal is as shown below:
1 Pa = 1N m^{-2}

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

Q.14 What is the SI unit of mass? How is it defined?

Q.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

Q.16 What do you mean by significant figures?

Q.17 A sample of drinking water was found to be severely contaminated with chloroform, $CHC\ell_3$, supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass).
(i) Express this in percent by mass.
(ii) Determine the molality of chloroform in the water sample.

Q.18 Express the following in the scientific notation:
(i) 0.0048 (ii) 234,000 (iii) 8008 (iv) 6.0012

Q.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

Q.20 Round up the following up to three significant figures:
(i) 34.216 (ii) 10.4107 (iii) 0.04597 (vi) 2808

Q.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 combination is obeyed by the above experimental data? Give its statement.
(b) Fil in the blanks in the following conversions :
(i) 1 km = ………………… mm = …………………pm
(ii) 1 mg = ………………… kg = …………………ng
(iii) 1 mL = ………………… L = …………………dm^3

Q.22 If the speed of light is 3.0×10^8 m s^{-1}, calculate the distance covered by light in 2.00 ns.

Q.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

Q.24 Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation:
$N_2(g) + H_2(g) \rightarrow 2NH_3(g)$
(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 its mass?

Q.25 How are 0.50 mol Na_2CO_3 and 0.50 M Na_2CO_3 different?

Q.26 If ten volumes of dihydrogen gas reacts with five volumes of dioxygen gas, how many volumes of water vapour would be produced?

Q.27 Convert the following into basic units:
(i) 28.7 pm (ii) 15.15 pm (iii) 25365 mg

Q.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 $C\ell_2(g)$

Q.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).

Q.30 What will be the mass of one ^{12}C atom in g?

Q.31 How many significant figures should be present in the answer of the following calculations?
(i) $\frac{0.02856 \times 298.15 \times 0.112}{0.5785}$ (ii) 5×5.364 (iii) $0.0125 + 0.7864 + 0.0215$

Q.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%

Q.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.

Q.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 if found to weight 11.6 g. Calculate (i) empirical formula, (ii) molar mass of the gas, and (iii) molecular formula.

Q.35 Calcium carbonate reacts with aqueous $HC\ell$ to give $CaC\ell_2$ and CO_2 according to the reaction, $CaCO_3(s) + 2\,HC\ell\,(aq) \rightarrow CaC\ell_2\,(aq) + CO_2(g) + H_2O(I)$
What mass of $CaCO_3$ is required to react completely with 25 mL of 0.75M $HC\ell$?

Q.36 Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid according to the reaction
$4HC\ell(aq) + MnO_2(s) \rightarrow 2H_2O(\ell) + MnC\ell(aq) + C\ell_2(g)$
How many grams of $HC\ell$ react with 5.0 g of manganese dioxide?

Exercise II

Q.1 What is the difference between the empirical and molecular formulas of a compound? Can they ever be the same? Explain.

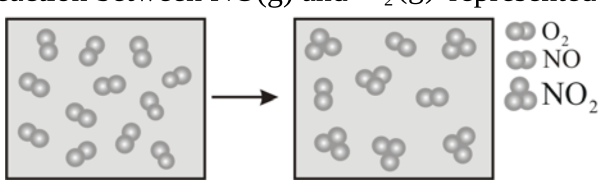
Q.2 Explain the law of conservation of mass, the law of definite proportion, and the law of multiple proportions.

Q.3 Hydrazine, ammonia, and hydrogen azide all contain only nitrogen and hydrogen. The mass of hydrogen that combines with 1.00 g of nitrogen for each compound is 1.44×10^{-1} g, 2.16×10^{-1} g, and 2.40×10^{-2} g, respectively. Show how these data illustrate the law of multiple proportions.

Q.4 What amount (moles) is represented by each of these samples?
(a) 1.03×10^{-4} mol (b) 4.52×10^{-3} mol (c) 3.41×10^{-2} mol

Q.5 Arrange the following substances in order of increasing mass percent of carbon.
(a) caffeine, $C_8H_{10}N_4O_2$ (b) ethanol, C_2H_5OH (c) surcrose, $C_{12}H_{22}O_{11}$

Q.6 Consider the reaction between NO(g) and $O_2(g)$ represented below.



What is the balanced equation for this reaction and what is the limiting reactant?

Q.7 A sample of an oxide of iron weighing 1.60 gm was heated in a stream of hydrogen gas until it was completely converted to 1.12 gm of metallic iron. What is the empirical formula of the iron oxide?

Q.8 Calculate the volume of O_2 at STP liberated by heating 12.25 g of $KClO_3$.
(At. wt. of K = 39, Cl = 35.5, .O = 16 u)

Q.9 Explain how compounds differ from elements? Give four differences

Q.10 Calculate the percentage of copper in sample of $CuCl_2$, (Atomic mass of Cu = 63.5 u, Cl = 35.5 u)