UNIT - 1

SOME BASIC CONCEPTS OF CHEMISTRY

- Chemistry is that branch of science dealing with study of composition, structure and properties of matter. It deals with the study of the changes in which different forms of matter undergo under different conditions.

- chemistry also had branches that look at the laws governing these changes.

Importance and scope of chemistry -

1) Supply of food -

- The study of chemistry provided the world with chemical fertilizers such as urea, calcium superphosphate, sodium nitrate and Ammonium sulphate.

- It has helped to protect the crops from insects and harmful bacteria by the use of certain effective insecticides, and pesticides.

- Chemistry also led to the discovery of preservatives, which helps to preserve food products for a long period.

2) <u>contribution to improved Health and sanitation facilities</u>:

chemistry provided manking with a large number of life - saving drugs.

3) <u>Increase in comfort, Pleasure and Luxuries</u>:

Because of the advancements in science and the discoveries of chemistry, we lead a more comfortable life today. This is examplified by the large-scale production synthetic fibres, Building materials and supply of metals.

Nature of matter

We know that matter comprises of particles and molecules. Base on its physical state, we can divide the nature of matter into three major categories.

1) solids -

- Solids are all those substances having their particles very close to each other. There exist strong intermolecular forces between these particles.

- The particles are firmly held in their positions. These particles have only vibratory motion.

- solids have definite shape and definite volume.

example : Wood, iron etc.

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2) Liquids -

liquids comprise of all those substances with weak intermolecular forces. The particles are capable of minimum movement. They have a definite volume. However, they do not have a definite shape. They usually take the shape of the container. in which we place them.

examples : Water, milk etc.

3) Gases -

Gases are those forms of matter having very weak forces between their molecules. Hence in gases, the molecules are free to move. The distance between molecules is large as compared to solids and liquids.

- Gases have neither fixed shape nor a definite volume. They tend to completely occupy the container in which they are placed

example - air, oxygen, hydrogen etc.

- We can change the state of the matter from one form to another by changing the conditions of pressure and temperature.

- We must note that the nature of matter also depends on its composition as well.

Classification of matter -

- If matter consists of more than one type of particles the it is a mixture. on the other hand, if it consists of a single type of particles then it is a pure substance.

- We can further classify mixtures into homogeneous and heterogeneous mixtures.

- In a homogeneous mixture, the compound completely mix with each other. This means particles of components of the mixtures are uniformly distributed throughout the bulk of the mixture. and its composition is uniform throughout.

example : sugar solution, air etc.

- In heterogeneous mixture, the composition is not uniform and some-times different components are visible.

example : mixture of salt and sugar, Grain and pulses. etc.

Pure substances -

Pure substances have characteristics different from mixtures. constituent particles of pure substances have fixed composition.

examples : copper, silver, gold etc.

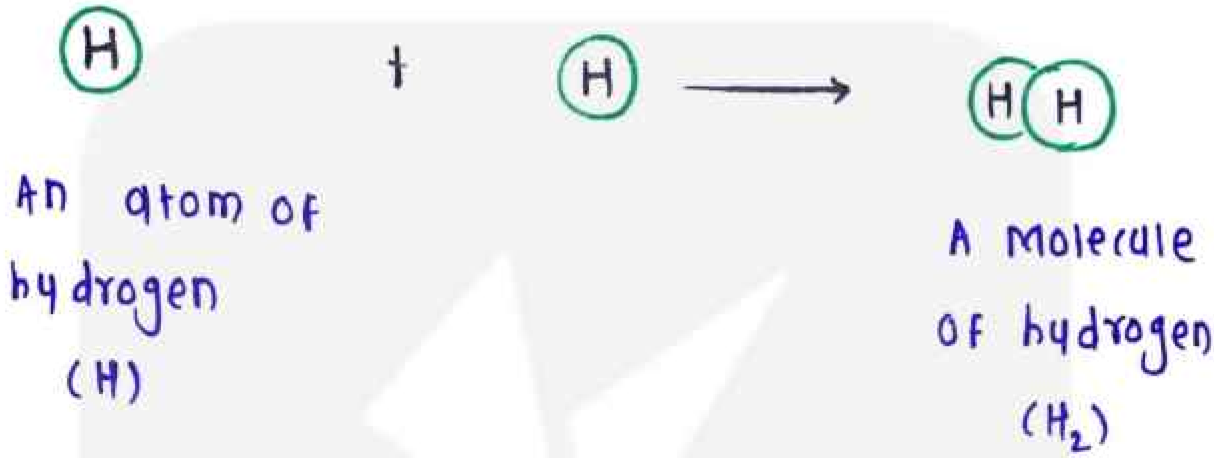
- pure substances can further be classified into elements and compounds. particles of an element consist of only one type of atoms. These particles may exist as atoms and molecules. For example :- sodium, copper, silver etc.

- When two or more atoms of different elements combine together in a definite ratio, the molecule of a compound is obtained.

- The constituents of a compound cannot be separated into simpler form by physical methods. They can be separated into simpler substances by chemical methods.

examples :- Water, carbon dioxide, sugar etc.

- A representation of atoms and molecules as follow :



- A water molecule comprises two hydrogen atoms and one oxygen atom. similarly, a molecule of carbon dioxide contains two oxygen atoms combined with one carbon atom.

- Thus, the atoms of different elements are present in a compound in a Fixed and definite ratio and this ratio is characteristic of a perticular compound. Also the properties of a compound are different from those of its cons-tituent elements.

- example : In H_2O, hydrogen and oxygen are gases. whereas, the compound formed by the combination is water. i.e., liquid.

- hydrogen burns with a pop sound and oxygen is a supporter of combustion, but water is used as a fire extinguisher.

Physical and chemical properties of matter :

- **Physical properties :**

These are the properties that we can measure without changing the chemical composition of the substance. The physical properties include mass, volume, density , refractive index etc.

- **chemical properties :**

These are the properties that we can evaluate at the cost of matter itself. As an example, we can test the sweet taste by eating it. It requires a chemical change to happen.

Measurement of Physical Properties -

We need to measure all physical quantities. We can express the value of physical quantity as the product of the numerical value and the unit in which it is expressed.

The International System of Units [SI] -

Fundamental Units :

Fundamental units are those which can neither be derived from one another nor they can be further resolved into any other units.

- The SI system has seven base units and they are listed below.

Base Physical Quantity	symbol for Quantity	Name of SI unit	symbol for SI unit.
length	l	metre	m
mass	m	kilogram	kg
Time	t	second	s
Thermodynamic Temperature	T	kelvin	K
Electric current	I	ampere	A
Amount of substance	n	mole	mol
luminous Intensity	I_v	candela	cd

- The other physical quantaties, such as speed, volume, density etc can be derived from these quantities.

The SI system allows the use of prefixes to indicate the multiples or submultiples of a unit.

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 the substance is constant, whereas, its weight may vary from one place to another due to change in gravity.

Volume -

- Volume is the place occupied by a substance. It has the units of $(length)^3$. So in SI system, volume has units of m^3.

- In chemistry laboratories, smaller volumes are used. Hence, volume is often denoted as 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.

Density -

The two properties - mass and volume discussed above are related as follows:

$$\text{Density} = \text{mass}/\text{volume}$$

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

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

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

Temperature -

- There are three common scales to measure temperature - °C (degree celcius), °F (degree fahrenheit) , and K (kelvin)

- kelvin is the SI unit.

- The temperatures on the two scales are related to each other by the following relationship :

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

- The kelvin scale is related to celsius scale as follows :

$$K = {^{\circ}C} + 273.15$$

- Temperature below $0^{\circ}C$ is possible i.e. negative values in Celsius scale But in kelvin scale, negative temp. is not possible.

Uncertainty in Measurement

Too often, we came across values that are close to each other and their average values. In such cases, we can say that the measurement is absolutely correct or precise. However, at times you may not experience this. At all those times, you have to mention the uncertainty in measurement.

- specifying this uncertainty is important as it will help you study the overall effect on output.

Scientific Notation -

- We know that atoms and molecules have extreamly low masses. However, we must not forget that they are present in massive numbers. scientiest have to deal with numbers that are large as 123,456, 789,101,110,000,000 and more.

- sometimes they also have to deal with numbers as small as 0.00000000166 g.

- To handle these numbers we use scientific notations. like $m \times 10^n$. Here we signify m times ten raised to the power of n. We can also see that n is an exponent which has positive and negative values and m is number which varies from 1.000 and 9.999

- similarly, we can write the scientific notion of 578.677 as 5.78677×10^2. In this, the decimal had to be moved to the left by the two places and same is the exponent of 10 in the scientific notation.

Multiplication and Division -

These two operations follow the same rules which are there for exponential numbers :

$$(5.6 \times 10^5) \times (6.9 \times 10^8) = (5.6 \times 6.9)(10^{5+8})$$
$$= (5.6 \times 6.9) \times 10^{13}$$
$$= 3.864 \times 10^{14}$$

Addition and subtraction -

- For adding 6.65×10^4 and 8.95×10^3, exponent is made same for both the numbers. Thus, we get $(6.65 \times 10^4) + (0.895 \times 10^4)$

Then these numbers can be added as follows:

$= (6.65 + 0.895) \times 10^4$

$= 7.545 \times 10^4$.

- similarly the substraction of two numbers can be done

Dimensional Analysis -

Dimensional Analysis is one of the simplest ways to carry out the calculations involving different units. Here, we convert a quantity expressed in one unit into an equivalent quantity with a different way.

- We do this by the use of conversion factor which expresses the relationship between units.

original quantity x conversion factor = equivalent quantity

(in former unit) (in other units)

- This is based on the fact that ratio of each fundamental quantity in one unit with their equivalent quantity in other unit is equal to one.

example – A jug contains 2L of milk calculate the volume of the milk in m^3.

Solution : since $1L = 1000\ cm^3$

and $1\ m = 100\ cm$, which gives

$$\frac{1\ m}{100\ cm} = 1 = \frac{100\ cm}{1\ m}$$

To get m^3 from the above unit factors, the first unit factor is taken and it is cubed.

$$\left(\frac{1m}{100cm}\right)^3 = \frac{1\ m^3}{10^6\ cm^3} = (1)^3 = 1.$$

Now $2L = 2 \times 1000\ cm^3$

- The above is multiplied by the unit factor

$$2 \times 1000\ cm^3 \times \frac{1 m^3}{10^6\ cm^3} = \frac{2 m^3}{10^3} = 2 \times 10^{-3} m^3$$

Laws of chemical combinations -

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

1) Law of conversion of mass -

French chemist, Antoine Lavoiser in 1789, studied this law.

- This law states that "In all physical and chemical changes, the total mass of the reactants is equal to that of Products" or "mass can neither be created nor destroyed"

- The mass and energy are interconv-ertable but the total sum of the mass and energy during any physical or chemical change remains constant.

2) Law of Definite proportion -

French chemist, J.L. Proust in 1799, discovered this law. It states that, " A chemical compound is always found to be made up of the same elements combined together in the fixed proportion by mass "

- For example, a sample of pure water from various sources or any country is always made up of only hydrogen and oxygen. These elements are always in the same fixed ratio of 1 : 8 by mass.

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 other element, are in ratio of small whole numbers"

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

Hydrogen + oxygen ⟶ Water
(2 g) (16 g) (18 g)

Hydrogen + oxygen ⟶ hydrogen peroxide
(2 g) (32 g) (34 g)

- The masses of oxygen in above reactions i.e 16 g and 32 g, which combined with a fixed mass of hydrogen bear a simple ratio = 16 : 32 = 1 : 2

4) Gay Lussac's Law of Gaseous Volume -

- When gases react together they always do so in volumes which bear a simple ratio to one another and to the volume of the products, if these are also gases. This holds true provided all measurement of volumes are done under similar conditions of temperature and pressure.

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

Thus, the volume of hydrogen and oxygen which combine bear a simple ratio of 2 : 1

- Gay Lussac's discovery in integer ratio in volume relationship is actually the law of definite proportion of volume.

5) Avogadro's Law -

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

- Avogadro made a distinction between atoms and molecules which is quite understable in which, If we consider reaction hydrogen and oxygen to produce water.

• We see that two volumes of hydrogen combines with one volume of oxygen to give two volumes of water without leaving any unreacted oxygen.

Dalton's Atomic Theory -

• In 1808 Dalton published 'A New system of chemical philosophy' in which he proposed the following:

1) Matter consist of indivisible atoms.

2) All atoms of a given element have identicle 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 involved reorganisation of atoms. These are neither created nor destroyed in chemical reaction.

- Dalton's theory could explain the laws of chemical combination. However, it could not explain the laws of gaseous volumes.

- It could not provide the reason for combining of atoms.

Atomic and Molecular Masses.

Atomic Mass -

- The atomic mass is such small a particle that is cannot seen or isolated. Therefore, it is impossible to determine the actual mass of a single atom by weighing it. Today, we have sophisticated techniques. eg. mass spectrometry for determining the atomic masses fairly accurate.

- The current 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 as ^{12}C. In this system, ^{12}C is assigned a mass exactly 12 atomic mass unit (amu) and masses of all other atoms are given related to this standard.

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

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

mass of an atom of hydrogen

$$= 1.6736 \times 10^{-24} \text{ g}$$

Thus, in turms 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.

- At present 'amu' has been replaced by 'u' which is unified mass.

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 the each elements by that 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 -

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

$= 16.043\,u$

- Similarly, molecular mass of H_2O

$= 2 \times$ atomic mass of hydrogen $+ 1 \times$ atomic mass of oxygen

$= 2(1.008\,u) + 16.00\,u$

$= 18.02\,u$

Formula Mass -

- It is the sum of the atomic weights of the various atoms present in the molecule of the substance.

For example : We can calculate the formula mass of $Na_2S = 2(23) + 1(32)$

$= 78.$

Mole concept and Molar masses

- Atoms and molecules are extremely small in size and their numbers in even a small amount of any substance is really very large. To handle such large numbers, we use the idea of mole to count. at microscopic level.

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

- In order to determine this number precisely, the mass of carbon-12 atom was determined by a mass spectrometer and found to be equal to 1.992648×10^{-23} g. Knowing that one mole of carbon weighs 12 g, the number of atoms in it is equal to:

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

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

- This number of entities in 1mol is so important that it is given a saparate name and symbol. It is known as 'Avogado constant'

- Having defined the mole, it is easier to know the mass op one mole of a substance or a constituent entities.

- 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 mol^{-1}

Molar mass of sodium chloride = 58.5 g mol^{-1}

Percentage composition -

- The percentage composition of any given element is nothing but the ratio of the amount of each element present in the compound to the total amount of individual elements present in the compound multiplied by 100
- Here we measure the quantity in terms of grams of the elements present in the solutions.

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

- example - Molar mass of water = 18.02 g

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

$$= 11.18$$

- mass % of oxygen = $\frac{16.00}{18.02} \times 100$

$= 88.79$

Empirical Formula for molecular Formula

- This formula helps in showing the lowest whole number of moles and the relative number of atoms of each element in a compound.

- This formula shows the exact number of atoms in the compound.

Stoichiometry and Stoichiometric calculations

- The word 'stoichiometry' is derived from two greek words 'stoichen' and 'metron' which means element and measure respec-tively.

- stoichiometry, thus deals with the calculation of masses of the reactants and the products involved in a chemical reaction.

- let us consider the combustion of methane.

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

- The coefficients 2 for O_2 and H_2O are called stoichiometric coefficients. similarly the coefficient of CH_4 and CO_2 is one in each case

- They represent the numbers of molecules (and moles as well) taking part in the reaction or formed in the reaction.

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) react with 2 molecules of O_2 (g) to give one molecule of CO_2 (g) and 2 molecules of H_2O (g)

- 22.7 L of CH_4 (g) reacts with 45.4L of O_2 (g) to give 22.7 L of CO_2(g) and 45.4L 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 this relationships, the given data can be interconverted as follows:

mass $\rightleftharpoons$ moles $\rightleftharpoons$ No. of molecules

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

Limiting reagent

Limiting reagent is an important concept in chemistry for chemical cal-culations. It refers to reactants which is present in minimum stoichio-metric quantity for a given chemical reaction.

- This perticular reactant is fully consumed in the chemical reaction. so, the calculation related to various products or in sequence of reaction on the basis of limiting reagent.

Reactions in solutions

A majority of reactions in the laboratories are carried out in solutions. Therefore it is important to understand how the amount of substances is expressed when it is present in the 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:

1) Mass per cent -

It is obtained by using the following relation

$$\text{mass per cent} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

2) Mole Fraction -

It is the ratio of number of moles of a particular component to the total number of moles of the solution.

- If the substance 'A' dissolved in substance 'B' and their number of moles are n_A and n_B respectively, then the mole Fractions of A and B are :

- mole fraction of A = $\dfrac{\text{No. of moles of A}}{\text{No. of moles of sol}^n}$

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

- mole fraction of B = $\dfrac{\text{No. of moles of B}}{\text{No. of moles of sol}^n}$

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

3) 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

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

- Unit of molarity : mol L^{-1}

4) Molality -

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

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