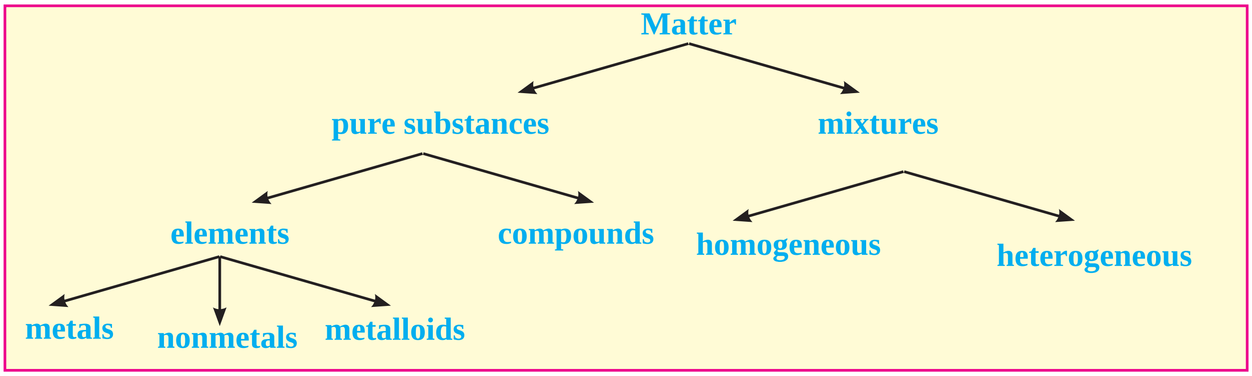
1. Some Basic Concepts of Chemistry

1.1 Introduction : Chemistry is the study of matter, its physical and chemical properties and the physical and chemical changes it undergoes under different conditions.

1.2.1. Matter : You have learnt earlier that matter occupies space and has mass. **Matter** can be further **classified** into pure substances and mixtures **on the basis of chemical composition**.



Chemistry is a central science. Its knowledge is required in the studies of physics, biological sciences, applied sciences, and earth and space sciences. The scope of chemistry is in every aspect of life, for example, the air we breathe, the food we eat, the fluids we drink, our clothing, transportation and fuel supplies, and so on.

Though it is an ancient science, due to development and advancement in science and technology, chemistry has developed as modern science. Technological development in sophisticated instruments expanded our knowledge of chemistry which, now, has been used in applied sciences such as medicine, dentistry, engineering, agriculture and in daily home use products.

1.2 Nature of Chemistry : Chemistry is traditionally classified further into five branches : organic, inorganic, physical, bio and analytical. Organic chemistry is the study of the properties and reactions of compounds of carbon. Inorganic chemistry is the study of all substances which are not organic. Physical chemistry is the study of principles underlying chemistry. It deals with the studies of properties of matter. It is study of atoms, molecules, and fundamental concepts related to electrons, energies and dynamics therein. It provides basic framework for all the other branches of chemistry.

Let us understand first what are pure substances and mixtures.

1.2.2 Pure substances versus mixtures :

Pure substances have a definite chemical composition. They always have the same properties regardless of their origin. Mixtures have no definite chemical composition and hence no definite properties.

Examples of pure substances : Pure metal, distilled water, etc.

Examples of mixtures : Paint (mixture of oils, pigment, additive), concrete (a mixture of sand, cement, water)

Can you tell?

Which are mixtures and pure substances from the following ?

i. sea water ii. gasoline iii. skin iv. a rusty nail v. a page of the textbook. vi. diamond

Pure substances are further divided into elements and compounds. **Elements are pure substances which can not be broken down into simpler substances by ordinary chemical changes.** Elements are further classified as metals, nonmetals and metalloids.

1. Metals :

i. have a lustre (a shiny appearance).
ii. conduct heat and electricity.
iii. can be drawn into wire (are ductile).
iv. can be hammered into thin sheets (are malleable).

Examples : gold, silver, copper, iron. Mercury is a liquid metal at room temperature.

2. Non-metals :

i. have no lustre. (exception : diamond, iodine)
ii. are poor conductors of heat and electricity. (exception : graphite)
iii. can not be hammered into sheets or drawn into wire, because they are brittle.

Examples : Iodine, nitrogen, carbon, etc.

3.Metalloids : Some elements have properties intermediate between metals and non-metals and are called metalloids or semi-metals. Examples include arsenic, silicon and germanium.

Compounds are the pure substances which can be broken down into simpler substances by ordinary chemical changes. In a compound, two or three elements are combined in a fixed proportion.

Mixture contains two or more substances in no fixed proportions and may be separated by physical methods. Mixtures are further divided into homogeneous and heterogeneous. Solutions are homogeneous mixtures, because the molecules of constituent solute and solvent are uniformly mixed throughout its bulk. In heterogeneous mixtures the molecules of the constituents are not uniformly mixed throughout the bulk. For example : Suspension of an insoluble solid in a liquid.

1.2.3 States of matter : You are also aware that matter exists in three different states namely gas, liquid and solid. You are going to learn about these states in unit 3 (chapter 10).

In solids, constituent atoms or molecules (particles) are tightly held in perfect order and therefore solids possess definite shape and volume. Liquids contain particles close to each other and they can move around within the liquid. While in gases, the particles are far apart as compared to those in liquid and solid state.

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

Can you tell?

Classify the following as element and compound.
i. mercuric oxide ii. helium gas iii. water iv. table salt v. iodine vi. mercury vii. oxygen viii. nitrogen

1.3 Properties of matter and their measurement :

Fig : 1.1 Burning of magnesium wire

Different kinds of matter have characteristic properties, which can be classified into two categories as **physical properties** and **chemical properties**.

Physical properties are those which can be measured or observed without changing the chemical composition of the substance. Colour, odour, melting point, boiling point, density, etc. are physical properties. Chemical properties are the properties where substances undergo a chemical change and thereby exhibit change in chemical composition. For example, coal burns in air to produce carbon dioxide or magnesium wire burns in air in the presence of oxygen to form magnesium oxide. (Fig. 1.1)

1.3.1 Measurement of properties : Many properties of matter are quantitative in nature. When you measure something, you are comparing it with some standard. The standard quantity is reproducible and unchanging.

Many properties of matter such as mass, length, area, pressure, volume, time, etc. are quantitative in nature. Any quantitative measurement is expressed by a number followed by units in which it is measured. For example, length of class room can be represented as 10 m. Here 10 is the number and 'm' denotes metre-the unit in which the length is measured.

The standards are chosen orbitrarily with some universally acceped criteria. **''The arbitrarily decided and universally accepted standards are called units.''**

There are several systems in which units are expressed such as CGS (centimetre for length, gram for mass and second for time), FPS (foot, pound, second) and MKS (metre, kilogram, second) systems, etc.

SI units :

In 1960, the general conference of weights and measure, proposed revised metric system, called **International System of units,** that is, **SI units**.

The metric system which originated in France in late eighteenth century, was more convenient as it was based on the decimal system. Later, based on a common standard system, the International System of Units (SI units) was established.

The SI system has seven base units as listed in Table 1.1. These are fundamental scientific quantities. Other units like speed, volume, density, etc. can be derived from these quantities.

Table 1.1 SI Fundamental units

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
Electric current	I	ampere	A
Thermodynamic temperature	T	Kelvin	K
Amount of substance	n	mole	mol
Luminous intensity	I_v	candela	cd

1.3.2 Physical properties

i. Mass and weight : We know that matter has mass. So mass is an inherent property of matter. It is the measure of the quantity of matter a body contains. The mass of a body does not vary as its position changes. On the other hand, the weight of a body is result of the mass and gravitational attraction. The weight of a body varies because the gravitational attraction of the earth for a body varies with the distance from the centre of the earth.

Hence, **the mass of a body is more fundamental property than its weight**.

The basic unit of mass in the SI system is the kilogram as given in Table 1.1. However, a fractional quantity 'gram' is used for weighing small quantities of chemicals in the laboratories. Therefore, in terms of grams it is defined (1kg = 1000 g = 10^3 g)

ii. Length : In chemistry we come across 'length' while expressing properties such as the atomic radius, bond length, wavelenght of electromagnetic radiation, and so on. These quantities are very small therefore fractional units of the SI unit of length are used for example, nanometre (nm), picometre (pm). Here 1nm = 10^{-9} m, 1 pm = 10^{-12} m.

iii. Volume : It is the amount of space occupied by a three dimensional object. It does not depend on shape. For measurement of volume of liquids and gases, a common unit, litre (L) which is not an SI unit is used.

1 L = 1 dm^3 = 1000 mL = 1000 cm^3

1000 cm^3 = 10 cm × 10 cm × 10 cm of volume

SI unit of volume is expressed as $(metre)^3$ or m^3.

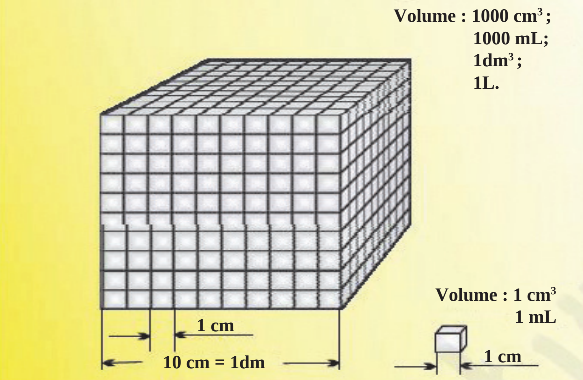


Fig. 1.2 : Litre and SI unit of volume

Different kinds of glassware are used to measure the volume of liquids and solutions. For example, graduated cylinder, burette, pipette, etc. A volumetric flask is used to prepare a known volume of a solution. Figure 1.3 shows the types of apparatus used in laboratory for measuring volume of liquids.

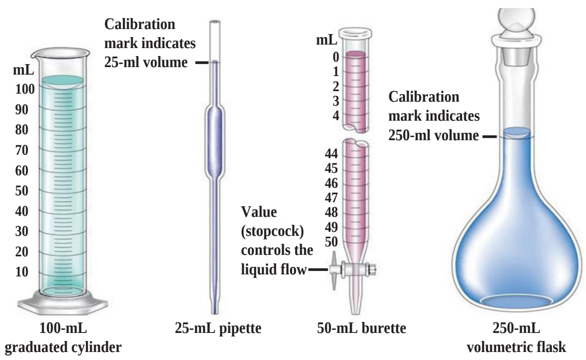


Fig. 1.3 : Volumetric glass apparatus

iv. Density : Density of a substance is its mass per unit volume. It is determined in the laboratory by measuring both the mass and the volume of a sample. The density is calculated by dividing mass by volume. It is the characteristic property of a substance.

So SI unit of density can be obtained as follows:

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

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

CGS units it is $\frac{\text{g}}{\text{mL}}$ or g mL^{-1} or g cm^{-3}

v. Temperature : Temperature is a measure of the hotness or coldness of an object. There are three common scales to measure temperature, namely ^{0}C (degree Celsius), ^{0}F (degree Fahrenheit) and K (Kelvin). Here K is the SI unit. Figure 1.4 shows the thermometers based on these scales.

Generally, the thermometer with celsius scale are calibrated from 0 ^{0}C to 100 ^{0}C where these two temperatures are respectively the freezing point and the boiling point of water at atmospheric pressure. These are represented on fahrenheit scale as 32^0 F to 212^0 F.

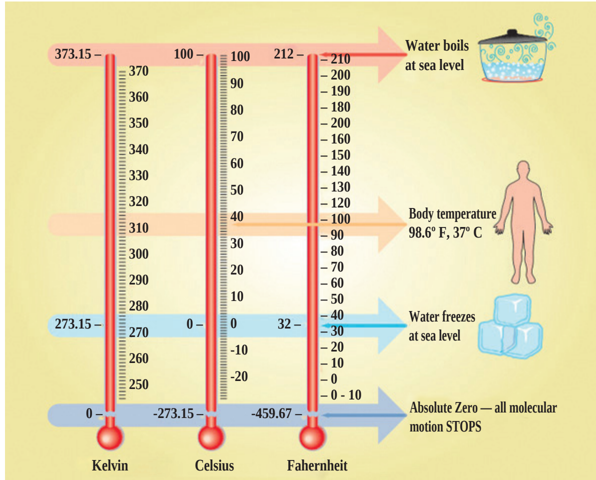


Fig 1.4 : Thermometers of different temperature scale

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

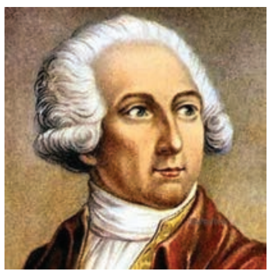
$$^0\text{F} = \frac{9}{5}(^0\text{C}) + 32$$

The Kelvin scale is related to Celsius scale as follows :

K = ^{0}C + 273.15

1.4 Laws of Chemical Combination : The elements combine with each other and form compounds. This process is governed by five basic laws discovered before the knowledge of molecular formulae.

1.4.1 Law of conservation of mass : Antoine



Lavoisier (1743-1794) a French scientist is often referred to as the father of modern chemistry. He carefully performed many combustion experiments, namely, burning of phosphorus and mercury, both in the presence of air. Both resulted in an increase in weight. After several experiments he found that the weight gained by the phosphorus was exactly the same as the weight lost by the air. He observed that,

Total mass of reactants

= Total mass of products

When hydrogen gas burns and combines with oxygen to yield water, the mass of the water formed is equal to the mass of the hydrogen and oxygen consumed. Thus, the law of conservation of mass states that **'mass can neither be created nor destroyed.'**

1.4.2 Law of Definite Proportions :

French chemist, Joseph Proust performed experiments on two samples of cupric carbonate. One of the samples was natural in origin and the other was a synthetic one. He found that the composition of elements present in it was same for both the samples as shown below :

Cupric Carbonate	% of copper	% of oxygen	% of carbon
Natural sample	51.35	38.91	9.74
Synthetic sample	51.35	38.91	9.74

This led Joseph Proust to state the law of definite proportion as follows :

'A given compound always contains exactly the same proportion of elements by weight.' Irrespective of the source, a given compound always contains same elements in the same proportion. The validity of this law has been confirmed by various experiments. This law is sometimes referred to as **Law of definite composition**.

1.4.3 Law of multiple proportions :

This law was proposed by John Dalton in 1803. It has been observed that two or more elements may form more than one compound. Law of multiple proportions summarizes many experiments on such compounds. **When two elements A and B form more than one compounds, the masses of element B that combine with a given mass of A are always in the ratio of small whole numbers.** For example, **i.** 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

Here, it is found that, the two masses of oxygen i.e. 16 g and 32 g which combine with a fixed mass of hydrogen (2g) are in the ratio of small whole numbers, i. e. 16:32 or 1:2.

ii. Nitrogen and oxygen combine to form two compounds, nitric oxide and nitrogen dioxide.

Nitrogen + Oxygen ⟶ Nitric Oxide
14 g 16 g 30 g

Nitrogen + Oxygen ⟶ Nitrogen Dioxide
14 g 32 g 46 g

Here, you find that the two masses of oxygen i.e. 16 g and 32 g when combine with a fixed mass of Nitrogen (14 g) are in the ratio of small whole numbers i.e. 16:32 or 1:2.

(Similar examples such as CO and CO_2 (1:2 ratio), SO_2 and SO_3 (2:3 ratio), can be found.)

1.4.4 Gay Lussac Law of Gaseous Volume :

This law was put forth by Gay Lussac in 1808. The law states 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**.

Illustration : i. Under the same conditions of temperature and pressure, 100 mL of hydrogen combines with 50 mL of oxygen to give 100 mL of water vapour.

Hydrogen (g) + Oxygen (g) ⟶ Water(g)
100 mL 50 mL 100 mL
(2 vol) (1 vol) (2 vol)

Thus, the volumes of hydrogen gas and oxygen gas which combine together i.e. 100 mL and 50 mL producing two volumes of water vapour which amounts to 100 mL bear a simple ratio of 2:1:2

ii. Under the same condition of temperature and pressure, 1 L of nitrogen gas combines with 3 L of hydrogen gas to produce 2 L of ammonia gas.

Nitrogen (g) + Hydrogen (g) $\longrightarrow$ Ammonia(g)

1 L	3 L	2 L
(1 vol)	(3 vol)	(2 vol)

Thus, the volume of nitrogen gas and hydrogen gas which combine together i.e. 1 L and 3 L and volume of ammonia gas produced i. e. 2 L bear a simple ratio of 1:3:2.

Remember

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

Can you tell?

If 10 volumes of dihydrogen gas react with 5 volumes of dioxygen gas, how many volumes of water vapour would be produced?

1.5 Avogadro Law : In 1811, Avogadro proposed that **equal volumes of all gases at the same temperature and pressure contain equal number of molecules**.

If we consider the reaction of hydrogen and oxygen to produce water vapour.

Hydrogen (g) +	Oxygen (g) $\longrightarrow$	Water (g)
100 mL	50 mL	100 mL
(2 vol)	(1 vol)	(2 vol)
(Gay Lussac Law)		
2n molecules	n molecules	2n molecules
(Avogadro law)		
2 molecules	1 molecule	2 molecules

We see that 2 volumes of hydrogen combine with 1 volume of oxygen to give 2 volumes of water vapour, without leaving any unreacted oxygen. According to Avogadro law, if 1 volume contains n molecules, then 2n molecules of hydrogen combine with n molecules of oxygen to give 2n molecules of water.

Therefore, 2 molecules of hydrogen gas combine with 1 molecule of oxygen to give 2 molecules of water vapour. Avogadro could explain the above result by considering the molecules to be polyatomic. If hydrogen and oxygen were considered as diatomic, as recognized now, then the above results are easily understandable.

Remember

Avogadro made a distinction between atoms and molecules, which is quite understandable in the present time.

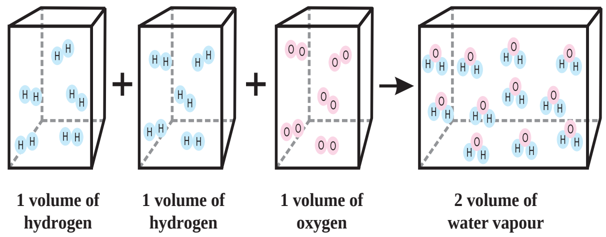


Fig. 1.5 : two volume of hydrogen react with one volume of oxygen to give two volumes of water vapour

1.6 Dalton's Atomic Theory : In 1808, Dalton published "A New System of chemical philosophy" in which he proposed the following features, which later became famous as Dalton's atomic theory.

1. Matter consists of **tiny, indivisible particles** called **atoms**.
2. All the atoms of a given elements have **identical** properties including **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 only the **reorganization of atoms**. Atoms are neither created nor destroyed in a chemical reaction.

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

Can you recall?

What is an atom and a moleule ? What is the order of magnitude of mass of one atom ? What are isotopes?

1.7 Atomic and molecular masses : You know about the terms atoms and molecules. Thus it is appropriate here to understand what we mean by atomic and molecular masses.

1.7.1 Atomic Mass : Every element has a characteristic atomic mass. Atomic mass is the mass of an atom. It is actually very very small. For example, the mass of one hydrogen atom is 1.6736×10^{-24} g. This is very small quantity and not easy to measure.

In the present system, mass of an atom is determined relative to the mass of a carbon - 12 atom as the standard and this has been agreed upon in 1961 by IU PAC. In this system, an atom of carbon-12 is assigned a mass of exactly 12.00000 atomic mass unit **(amu)** and all other atoms of other elements are given a relative atomic mass, to that of carbon - 12. The atomic masses are expressed in **amu**. **One amu is defined** as a **mass exactly equal to one twelth of the mass of one carbon-12 atom**. Later on the exact value of atomic mass unit in grams was experimentally established.

$$1 \text{ amu} = \frac{1}{12} \times \text{mass of one C-12}$$

$$= \frac{1}{12} \times 1.992648 \times 10^{-23} \text{ g}$$

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

Recently, amu has been replaced by **unified mass unit** called **dalton** (symbol '**u**' or '**Da**'), 'u' means unified mass.

Problem 1.1 : Mass of an atom of oxygen in gram is 26.56896×10^{-24} g. What is the atomic mass of oxygen in u ?

Solution : Mass of an atom of oxygen in gram is 26.56896×10^{-24} g, and

1.66056×10^{-24} g = 1 u

$\therefore$ 26.56896×10^{-24} g = ?

$$= \frac{26.56896 \times 10^{-24} \text{ g}}{1.66056 \times 10^{-24} \text{ g/u}} = 16.0 \text{ u}$$

Similarly mass of an atom of hydrogen = 1.0080 u

1.7.2 Average Atomic Mass : Many naturally occuring elements exist as mixture of more than one isotope. Isotopes have different atomic masses. The atomic mass of such an element is the weighted average of atomic masses of its isotopes (taking into account the atomic masses of isotopes and their relative abundance i.e. percent occurrance). This is called **average atomic mass** of an element. For example, carbon has the following three isotopes with relative abundances and atomic masses as shown against each of them.

Isotope	**Atomic mass (u)**	**Relative Abundance (%)**
^{12}C	12.00000	98.892
^{13}C	13.00335	1.108
^{14}C	14.00317	2×10^{-10}

From the above data, the average atomic mass of carbon

= (12 u) (98.892/100) + (13.00335 u) (1.108/100) + (14.00317) $(2 \times 10^{-10}/100)$

= 12.011 u

Similary, average atomic masses for other elements can be calculated.

Remember

- In the periodic table of elements, the atomic masses mentioned for different elements are actually their average atomic masses.
- For practical purpose, the average atomic mass is rounded off to the nearest whole number when it differs from it by a very small fraction.

Element	Isotopes	Average atomic mass	Rounded off atomic mass
Carbon	^{12}C, ^{13}C, ^{14}C	12.011 u	12.0 u
Nitrogen	^{14}N, ^{15}N	14.007 u	14.0 u
Oxygen	^{16}O, ^{17}O, ^{18}O	15.999 u	16.0 u
Chlorine	^{35}Cl, ^{37}Cl	35.453 u	35.5 u
Bromine	^{79}Br, ^{81}Br	79.904 u	79.9 u

Problem 1.2 : Calculate the average atomic mass of neon using the following data :

Isotope	Atomic mass	Natural Abundance
^{20}Ne	19.9924 u	90.92%
^{21}Ne	20.9940 u	0.26 %
^{22}Ne	21.9914 u	8.82 %

Solution : Average atomic mass of Neon (Ne)

$$= \frac{\text{Atomic mass of } ^{20}\text{Ne} \times \% + \text{Atomic mass of } ^{21}\text{Ne} \times \% + \text{Atomic mass of } ^{22}\text{Ne} \times \%}{100}$$

$$= \frac{(19.9924u)(90.92) + (20.9940u)(0.26) + (21.9914u)(8.82)}{100} = 20.1707 \text{ u}$$

1.7.3 Molecular Mass : Molecular mass of a substance is the sum of average atomic masses of all the atoms of elements which constitute the molecule. Molecular mass of a substance is the mass of one molecule of that substance relative to the mass of one carbon-12 atom. It is obtained by multiplying average atomic mass of each element by the number of its atoms and adding them together.

For example, the molecular mass of carbon dioxide (CO_2) is

= 1(average atomic mass of C)

+ 2 (average atomic mass of O)

= 1 (12.0 u) + 2 (16.0 u) = 44.0 u

Some more examples of calculations of molecular mass.

i. $H_2O = 2 \times 1 \text{ u} + 16 \text{ u} = 18 \text{ u}$

ii. $C_6H_5Cl = (6 \times 12 \text{ u}) + (5 \times 1 \text{ u}) + (35.5 \text{ u})$

$= 112.5 \text{ u}$

iii. $H_2SO_4 = (2 \times 1 \text{ u}) + (32 \text{ u}) + (4 \times 16 \text{ u}) = 98 \text{ u}$

Problem 1.3 : Find the mass of 1 molecule of oxygen (O_2) in amu (u) and in grams.

Solution : Molecular mass of O_2 = 2 × 16 u

∴ mass of 1 molecule = 32 u

∴ mass of 1 molecule of O_2

$= 32.0 \times 1.66056 \times 10^{-24}$ g

$= 53.1379 \times 10^{-24}$ g

1.7.4 Formula Mass

Some substances such as sodium chloride do not contain discrete molecules as the constituent units. In such compounds, cationic (sodium) and anionic (chloride) entities are arranged in a three dimensional structure.

In sodium chloride crystal, one $Na^{\oplus}$ ion is surrounded by six $Cl^{\ominus}$ ions, all at the same distance from it and vice versa. Therefore, NaCl is the formula used to represent sodium chloride, though it is not a molecule. Similarly, a term 'formula mass' is used for such ionic compounds, instead of molecular mass. **The formula mass of a substance is the sum of atomic masses of the atoms present in the formula.**

Problem 1.4 : Find the formula mass of

i. NaCl ii. $Cu(NO_3)_2$

i. Formula mass of NaCl

= average atomic mass of Na

+ average atomic mass of Cl

= 23.0 u + 35.5u = 58.5 u

ii. Formula mass of $Cu(NO_3)_2$

= average atomic mass of Cu + 2 × (average atomic mass of nitrogen + average atomic mass of three oxygen)

= (63.5) + 2(14 + 3 x 16) = 187.5 u

Try this

Find the formula mass of $CaSO_4$

If atomic mass of Ca = 40.1 u,

S =32.1 u and O = 16.0 u

1.8 Mole concept and molar mass

Can you recall?

1. One dozen means how many items ?
2. One gross means how many items ?

Mole : Expressing large count of objects is made easy by using quantitative adjectives such as dozen, gross. You know that even a small amount of any substance contains very large number of atoms or molecules. We use a quantitative adjective 'mole' to express the large number of submicroscopic entities like atoms, ions, electrons, etc. present in a substance.

Definiton : One mole is the amount of a substance that contains as many entities or particles as there are atoms in exactly 12 g (or 0.012 kg) of the carbon -12 isotope.

Let us calculate the number of atoms in 12.0000 g of Carbon-12 isotopes. Mass of one carbon-12 atom (determined by mass spectrometer) = 1.992648×10^{-23} g,

Mass of one mole carbon atom = 12 g

$\therefore$ Number of atoms in 12 g of carbon -12

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

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

Thus one mole is the amount of a substance that contains 6.0221367×10^{23} particles/ entities (such as atoms, molecules or ions).

Note that the name of the unit is **mole** and the symbol for the unit is **mol**.

Remember

The number 6.0221367×10^{23} is known as **Avogadro's Constant 'N_A'** in the honour of Amedo Avogadro.

In SI system, mole (Symbol mol) was introduced as seventh base quantity for the amount of a substance.

Example :

1 mole of oxygen atoms = 6.0221367×10^{23} atoms of oxygen

1 mole of water molecules = 6.0221367×10^{23} molecules of water

1 mole of sodium chloride = 6.0221367×10^{23} formula units of NaCl

Molar Mass : The mass of one mole of a substance (element/compound) in grams is called its molar mass. The molar mass of any element in grams is numerically equal to atomic mass of that element in u.

Element	Atomic mass (u)	Molar mass (g mol^{-1})
H	1.0 u	1.0 g mol^{-1}
C	12.0 u	12.0 g mol^{-1}
O	16.0 u	16.0 g mol^{-1}

Similarly molar mass of any substance, existing as polyatomic molecule, in grams is numerically equal to its molecular mass or formula mass in u.

Polyatomic substance	Molecular/ formula mass (u)	Molar mass (g mol^{-1})
O_2	32.0 u	32.0 g mol^{-1}
H_2O	18.0 u	18.0 g mol^{-1}
NaCl	58.5 u	58.5 g mol^{-1}

Molar mass of O atoms

$= 6.022 \times 10^{23}$ atom/mol $\times$ 16 u/atom

$\times 1.66056 \times 10^{-24}$ g/u = 16.0 g/mol

Problem 1.5 : Calculate the number of moles and molecules of urea present in 5.6 g of urea.

Solution : Mass of urea = 5.6 g

Molecular mass of urea, NH_2CONH_2

= 2 (average atomic mass of N) + 4 (average atomic mass of H) + 1(average atomic mass of C) + 1(average atomic mass of O)

= 2 × 14 u +1 × 12 u +4 ×1 u + 1 × 16 u

= 60 u

$\therefore$ molar mass of urea = 60 g mol^{-1}

Number of moles

$$= \frac{\text{mass of urea in g}}{\text{molar mass of urea in g mol}^{-1}}$$

$$= \frac{5.6 \text{ g}}{60 \text{ g mol}^{-1}} = 0.0933 \text{ mol}$$

Number of molecules = Number of moles × Avogadro's constant

Number of molecules of urea

$= 0.0933 \times 6.022 \times 10^{23}$ molecules/mol

$= 0.5618 \times 10^{23}$ molecules

$= 5.618 \times 10^{22}$ molecules

Ans : Number of moles = 0.0933 mol

Number of molecules of urea

$= 5.618 \times 10^{22}$ molecules

Problem 1.6 : Calculate the number of atoms in each of the following
i. 52 moles of Argon (Ar)
ii. 52 u of Helium (He)
iii. 52 g of Helium (He)

Solution :

i. 52 moles of Argon

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

$\therefore$ 52 moles of Ar

$$= 52 \text{ moles} \times \frac{6.022 \times 10^{23}}{1\text{mol}} \text{ atoms}$$

$= 313.144 \times 10^{23}$ atoms of Argon

ii. 52 u of Helium

Atomic mass of He = mass of 1 atom of He = 4.0u

$\therefore$ 4.0 u = 1 He $\therefore$ 52 u = ?

$$= 52 \text{ u} \times \frac{1\text{atom}}{4.0 \text{ u}} = 13$$

atoms of He

iii. 52 g of He

Mass of 1 mole of He = 4.0 g

Number of moles of He

$$= \frac{\text{mass of He}}{\text{mass of 1mole of He}}$$

$$= \frac{52 \text{ g}}{4.0 \text{ g mol}^{-1}} = 13 \text{ mol}$$

Number of atoms of He

= Number of moles $\times 6.022 \times 10^{23}$

= 13 mol $\times 6.022 \times 10^{23}$ atoms/mol

= 78.286×10^{23} atoms of He.

1.9 Moles and gases : Many substances exist as gases. If we want to find the number of moles of gas, we can do this more conveniently by measuring the volume rather than mass of the gas. Chemists have deduced from Avogardro law that "One mole of any gas occupies a volume of 22.4 dm^3 at **standard temperature (0^{0}C) and pressure (1 atm) (STP)**. The volume of 22.4 dm^3 at STP is known as molar volume of a gas.

Number of moles of a gas (n) =

$$\frac{\text{Volume of the gas at STP}}{\text{Molar volume of gas}}$$

Thus

Number of moles of a gas (n) =

$$\frac{\text{Volume of the gas at STP}}{22.4 \text{ dm}^3\text{mol}^{-1}}$$

Number of molecules = number of moles $\times$ 6.022×10^{23} molecules mol^{-1}

(Note : IUPAC has recently changed the standard pressure to 1 bar. Under these **new STP conditions** the molar volume of a gas is 22.71 L mol^{-1})

Problem 1.7 : Calculate the number of moles and molecules of ammonia (NH_3) gas in a volume 67.2 dm^3 of it measured at STP.

Solution :

Volume of NH_3 at STP = 67.2 dm^3

molar volume of a gas = 22.4 dm^3 mol^{-1}

Number of moles (n)

$$= \frac{\text{Volume of the gas at STP}}{\text{Molar volume of gas}}$$

$$\text{Number of moles of } NH_3 = \frac{67.2 \text{ dm}^3}{22.4 \text{ dm}^3 \text{ mol}^{-1}}$$

= 3.0 mol

Number of molecules = Number of moles $\times$ 6.022×10^{23} molecules mol^{-1}

Number of molecules of NH_3 = 3.0 mol $\times$ 6.022×10^{23} molecules mol^{-1}

= 18.066×10^{23} molecules

Try this

Calculate the volume in dm^3 occupied by 60.0 g of ethane at STP.

Exercises

1. Choose the most correct option.

A. A sample of pure water, whatever the source always contains .... by mass of oxygen and 11.1 % by mass of hydrogen.

a. 88.9 b. 18 c. 80 d. 16

B. Which of the following compounds can NOT demonstrate the law of multiple proportions ?

a. NO, NO_2 b. CO, CO_2

c. H_2O, H_2O_2 d. Na_2S, NaF

C. Which of the following temperature will read the same value on Celsius and Fahrenheit scales.

a. - 40^0 b. + 40^0

c. -80^0 d. -20^0

D. SI unit of the quantity electric current is ....

a. Volt b. Ampere

c. Candela d. Newton

E. In the reaction $N_2 + 3H_2 \longrightarrow 2NH_3$, the ratio by volume of N_2, H_2 and NH_3 is 1 : 3 : 2 This illustrates the law of

a. definite proportion

b. reciprocal proportion

c. multiple proportion

d. gaseous volumes

F. Which of the following has maximum number of molecules ?

a. 7 g N_2 b. 2 g H_2

c. 8 g O_2 d. 20 g NO_2

G. How many g of H_2O are present in 0.25 mol of it ?

a. 4.5 b. 18

c. 0.25 d. 5.4

H. The number of molecules in 22.4 cm^3 of nitrogen gas at STP is

a. 6.022 x 10^{20}

b. 6.022 x 10^{23}

c. 22.4 x 10^{20}

d. 22.4 x 10^{23}

I. Which of the following has the largest number of atoms ?

a. 1g Au (s) b. 1g Na (s)

c. 1g Li (s) d. 1g Cl_2 (g)

2. Answer the following questions.

A. State and explain Avogadro's law.

B. Point out the difference between 12 g of carbon and 12 u of carbon

C How many grams does an atom of hydrogen weigh ?

D. Calculate the molecular mass of the following in u.

a. NH_3 b. CH_3COOH c. C_2H_5OH

E. How many particles are present in 1 mole of a substance ?

F. What is the SI unit of amount of a substance ?

G. What is meant by molar volume of a gas ?

H. State and explain the law of conservation of mass.

I. State the law of multiple proportions.

3. Give one example of each

A. homogeneous mixture

B. heterogeneous mixture

C. element D. compound

4. Solve problems :

A. What is the ratio of molecules in 1 mole of NH_3 and 1 mole of HNO_3.

(Ans. : 1:1)

B. Calculate number of moles of hydrogen in 0.448 litre of hydrogen gas at STP

(Ans. : 0.02 mol)

C. The mass of an atom of hydrogen is 1.008 u. What is the mass of 18 atoms of hydrogen. (18.144 u)

D. Calculate the number of atom in each of the following (Given : Atomic mass of I = 127 u).

a. 254 u of iodine (I)

b. 254 g of iodine (I)

(Ans. : 2 atoms, 1.2044 x 10^{24} atoms)

E. A student used a carbon pencil to write his homework. The mass of this was found to be 5 mg. With the help of this calculate.

a. The number of moles of carbon in his homework writing.

(Ans : 4.16 x 10^{-4})

b. The number of carbon atoms in 12 mg of his homework writting

(Ans : 6.022 x 10^{20})

F. Arjun purchased 250 g of glucose ($C_6H_{12}O_6$) for Rs 40. Find the cost of glucose per mole.

(Ans : Rs 28.8)

G. The natural isotopic abundance of ^{10}B is 19.60% and ^{11}B is 80.40 %. The exact isotopic masses are 10.13 and 11.009 respectively. Calculate the average atomic mass of boron

(Ans. :10.81)

H. Convert the following degree Celsius temperature to degree Fahrenheit.

a. 40 ^0C b. 30 ^0C

(Ans. : A. 104 ^0F, B. 86 ^0F)

I. Calculate the number of moles and molecules of acetic acid present in 22 g of it.

(Ans. : 0.3666 mol, 2.2076 x 10^{23} molecules)

J. 24 g of carbon reacts with some oxygen to make 88 grams of carbon dioxide. Find out how much oxygen must have been used.

(Ans. : 64.0)

K. Calculate number of atoms is each of the following. (Average atomic mass : N = 14 u, S = 32 u)

a. 0.4 mole of nitrogen

b. 1.6 g of sulfur

(Ans. : A. 2.4088 x 10^{23}, B. 3.011 x 10^{22} atom)

L. 2.0 g of a metal burnt in oxygen gave 3.2 g of its oxide. 1.42 g of the same metal heated in steam gave 2.27 of its oxide. Which law is verified by these data ?

M. In two moles of acetaldehyde (CH_3CHO) calculate the following

a. Number of moles of carbon

b. Number of moles of hydrogen

c. Number of moles of oxygen

d. Number of molecules of acetaldehyde

(Ans. : A. 4 mol, B. 8 mol, C. 2 mol, D. 12.044 x 10^{23} molecules)

N. Calculate the number of moles of magnesium oxide, MgO in i. 80 g and ii. 10 g of the compound. (Average atomic masses of Mg = 24 and O = 16)

(Ans. i. 2 mol ii. 0.25 mol)

O. What is volume of carbon dioxide, CO_2 occupying by i. 5 moles and ii. 0.5 mole of CO_2 gas measured at STP.

(Ans. i. 112 dm^3 ii. 11.2dm^3)

P. Calculate the mass of potassium chlorate required to liberate 6.72 dm^3 of oxygen at STP. Molar mass of $KClO_3$ is 122.5 g mol^{-1}.

(Ans. 24.5 g)

Q. Calculate the number of atoms of hydrogen present in 5.6 g of urea, $(NH_2)_2CO$. Also calculate the number of atoms of N, C and O.

(Ans. : No. of atoms of H = 2.24 x 10^{23}, N =1.124 x 10^{23} and C = 0.562 x 10^{23}, O = 0.562 x 10^{23})

R. Calculate the mass of sulfur dioxide produced by burning 16 g of sulfur in excess of oxygen in contact process. (Average atomic mass : S = 32 u, O = 16 u)

(Ans. 32 g)

5. Explain

A. The need of the term average atomic mass.

B. Molar mass.

C. Mole concept.

D. Formula mass with an example.

E. Molar volume of gas.

F. Types of matter (on the basis of chemical composition).

Activity :

Collect information of various scientists and prepare charts of their contribution in chemistry.