2

Structure and Properties of Matter

All the objects around us whether living or non-living are matter. Water we drink, food we eat, air we breathe, chair we sit on, are all examples of matter. **Matter is anything that has mass and takes up space**. Matter appears in a huge variety of forms such as rocks, trees, computer, clouds, people, etc. Matter embraces each and everything around us. Therefore, in order to understand the world, it would be necessary to understand the matter.

Each pure kind of matter is called **substance.** Here, pure we mean the same through out. Thus, aluminium is one substance and water is another. Please remember, the scientific meaning of substance is a little different from its every day meaning and we shall discuss it a little later in this lesson.

OBJECTIVES

After completing this lesson, you will be able to:

- define various states of matter as solid, liquid and gas, and distinguish one from the other based on their properties;
- classify the matter based on their composition as element, compound and mixture;
- differentiate between atoms and molecules;
- state Dalton's atomic theory and explain various laws of chemical combinations;
- define isotopes, atomic mass and molecular mass;
- express chemical reaction in form of a balanced chemical equation;
- define mole concept and molar quantities such as molar mass and molar volume;
- apply mole concept to a chemical reaction and show a quantitative relationship between masses of reactants and products;
- define Gay Lussac's law of combining volume and Avogadro's law;
- solve numerical problems based on various concepts covered above;

2.1 CLASSIFICATION OF MATTER

Earlier Indian and Greek philosophers and scientists attempted to classify the matter in the form of five elements - Air, Earth, Fire, Sky and Water. This classification was more of philosophical nature. In modern science, however, there are two main ways of classifying the matter :

i) **Based on physical states:** All matter, at least in principle, can exist in three states, solid, liquid and gas.

ii) Based on composition and properties: The classification of matter includes elements, compounds and mixtures.

2.1.1 PHYSICAL STATE OF MATTER

A given kind of matter may exist in different physical forms under different conditions. Water, for example, at one atmospheric pressure, may exist as solid, liquid or gas with change of temperature. Sodium metal is normally solid, but it melts to a silvery liquid when heated to 98 ^{o}C. Liquid sodium changes to a bluish gas if the temperature is raised to 883 ^{o}C. Similarly, chlorine, which is normally a gas can exist as a yellow liquid or solid under appropriate conditions. These three different forms of matter differ from each other in their properties. Solids are rigid with definite shapes. Liquids are less rigid than solids and are fluid, i.e. they are able to flow and take the shape of their containers. Like liquids, gases are fluids, but unlike liquid, they can expand indefinitely.

Can you think of other differences between a gas and liquid? A gas can be compressed easily whereas a liquid cannot. You might be aware, natural gas is compressed and supplied as fuel for vehicles in the name of CNG (compressed natural gas). It is not possible to compress a liquid. It is still more difficult to compress a solid. All these three forms of matter (solid, liquid and gas) are generally referred as **states of matter.** Taking fluidity/ rigidity and compressibility, we can write characteristic properties of solid, liquid and gas in the Table 2.1.

Table 2.1: Characteristics of different states of matter

States of matter	Fluidity/rigidity	Compressibility
Solid	Rigid	Negligible
Liquid	Fluid	Very low
Gas	Fluid	High

As mentioned, a substance can exist in three forms depending upon temperature and pressure. Water at room temperature (25 ^{o}C) exists in liquid form and at 0 ^{o}C and 1 atmospheric pressure as solid. If we go on increasing temperature of water at constant pressure, more and more of it will go into vapour form and at 100 ^{o}C will start boiling,. If we continue heating at this temperature (100 ^{o}C), entire liquid water will be converted into vapour. This is true with most of the liquids. Definitely melting and boiling points of different substances will be different. Can you think why this variation in their melting point and boiling point occurs? You will study later on that **intermolecular forces** are different in different liquids, and therefore their boiling points and melting points are different. In gaseous form, intermolecular forces are very weak and unable to keep molecules together in aggregation. However, in case of solids, these forces are very strong and capable of keeping molecules in fixed positions. This is the reason solids are rigid and hard and cannot be compressed. Liquids have properties intermediate to solid and gases as intermolecular forces between molecules in liquid are definitely more than gases and less than solids but strong enough to keep the molecules in aggregation (Fig. 2.1). Due to weak intermolecular forces in gases, molecules in gases can move freely and can occupy any space available to them. This property of gases is responsible for their **effusion/diffusion**. Molecules in gases are far apart and therefore when pressure is applied they can be brought closer and gases can be compressed.

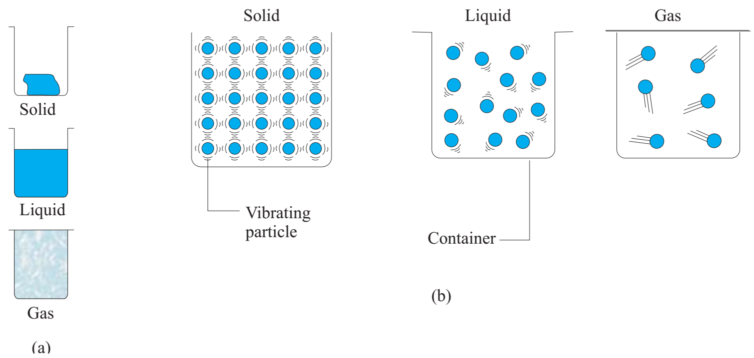

Fig. 2.1. Gases, liquids and solids (a) Bulk appearance and (b) the molecular picture.

ACTIVITY 2.1

Fill gas in a balloon and tightly tie its mouth. Now hold it with both hands and compress. What do you find? Balloon can be compressed easily.

2.2 CLASSIFICATION OF MATTER BASED ON COMPOSITION - ELEMENTS, COMPOUNDS AND MIXTURES

Another method of classification of matter is based on its **composition**. A **substance** is matter that has a definite or constant **composition** and has distinct properties. Examples are aluminium sliver, water, carbon dioxide, nitrogen, oxygen etc. Substances differ from one another in composition and can be identified by their properties like colour, smell, taste, appearance, etc. Aluminium has uniform composition. Similarly water has uniform composition. No doubt there are also matter which do not have uniform composition. Such matter are called mixtures. Some examples of mixtures are air, soft drink, milk, and cement. Mixtures are either homogeneous or heterogeneous. Suppose you add 5g of sugar to water kept in a glass tumbler. After stirring, the mixture obtained is uniform through out. This mixture is **homogeneous** through out and is called **solution**. Air is solution of several gases (oxygen, nitrogen, water vapour, carbon dioxide etc). Suppose you mix sand with iron filings, sand grains and the iron filings remain visible and separate. This type of mixture in which the composition is not uniform, is called a **heterogeneous mixture.** If you add oil to water, it creates another heterogeneous mixture because the liquid thus obtained does not have a uniform composition.

We can create homogeneous and heterogeneous mixtures and if need arises we can separate them into pure components by physical means without changing the identities of the components. We can recover sugar from its water solution by heating and evaporating the solution to dryness. From a mixture of iron filings and sand, we can separate iron filings using magnet. After separation we can see that the components have the same composition and properties as they did to start with.

2.2.1 Elements

Oxygen and magnesium, these two substances which have uniform composition through out are elements. Antoine Laurent Lavoisier (1743-94), a French chemist was first to explain an element. He defined **an element as basic form of matter that cannot be broken down into simpler substances even by chemical means.** Elements serve as the **building blocks** for various types of other substances, starting from water up to extremely complex substances like protein. Oxygen, nitrogen, magnesium, iron, gold all are example of element which you have already studied in your lower classes. Today more than 112 elements are known and we know various details about them. **An element consists of only one kind of atoms.** These elements are represented by suitable symbols, as you must have read in your previous classes. Fig. 2.2 shows the most abundant element in earth crust and in the human body. As can be seen from the figure, only five elements (oxygen, silicon, aluminium, iron and calcium) comprise over 90 per cent of Earth's crust. Out of these five, oxygen is the most abundant element in our body.

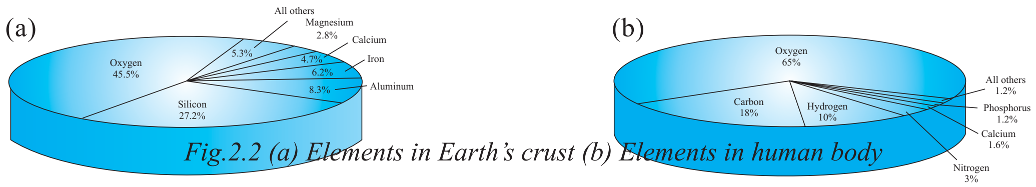

Fig.2.2 (a) Elements in Earth's crust (b) Elements in human body

2.2.2 Compounds

Most elements can interact with one or more other elements to form compounds. **A compound is a substance that consists of two or more different elements chemically united in a definite ratio.** A pure compound, whatever its source, always contains definite or constant proportions of the elements by mass. As you have read, water is composed of two elements: hydrogen and oxygen. Property of water is completely different from its constituent elements: hydrogen and oxygen which are gases. Similarly when sulphur is ignited in air, sulphur and oxygen (from air) combine to form **sulphur dioxide.** All sample of pure water contain these two elements combined in the ratio of one is to eight (1: 8) by mass. For example, 1.0 g of hydrogen will combine with 8.0g of oxygen. This regularity of composition by mass will be discussed later on as **law of constant composition).** This composition does not change whether we take water from river of India or of United States or the ice caps on Mars. Unlike mixtures compounds can be separated only by chemical means into their pure components.

In conclusion, the relationship among elements, compounds and other categories of matter are summarised in Fig. 2.3. We have just read that elements are made of one kind of atoms only. Now we shall discuss how concept of an atom emerged and how far this forms the basis of our other studies in science.

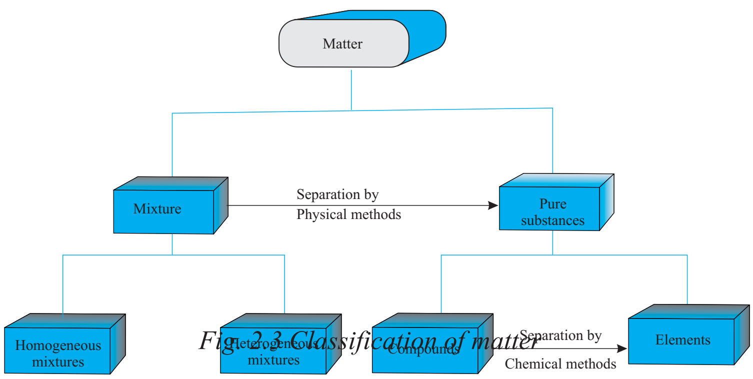

Fig. 2.3 Classification of matter

CHECK YOUR PROGRESS 2.1

1. Which of the following matter fall(s) in the category of substance?
   (i) Ice (ii) Milk (iii) Iron (iv) Air (v) Water (vi) Hydrochloric acid
2. Which one of the following is solution?
   (i) Mercury (ii) Air (iii) Coal (iv) Milk

2.3 DALTON'S ATOMIC THEORY

In the fifth century B.C. Indian philosopher Maharshi Kanad postulated that if one goes on dividing matter (Padarth), he would get smaller and smaller particles and a limit will come when he will come across smallest particles beyond which further division will not be possible. He (Kanad) named the particles **Parmanu.** More or less during the same period Greek philosophers, Leuappus and Democritus suggested similar ideas. This idea was not accepted at that time but it remained alive. Not much experimental work could be done until Lavoisier gave his law: **Law of conservation of mass** and **law of constant proportions** sometimes in 1789. English scientist and school teacher, John Dalton (1766-1844) provided the basic theory about the nature of matter: *All matter whether element, compound or mixture is composed of small particles called atoms.*

Dalton's theory can be summarized as follows:

- Elements are composed of extremely small indivisible particles called atoms.
- All atoms of a given element are identical, having the same size, mass and chemical properties. The atoms of one element are different from the atoms of all other elements.
- Compounds are composed of atoms of more than one element. In any compound the ratio of the numbers of atoms of any two of the elements present is either an integer or a simple fraction.
- A chemical reaction involves only the separation, combination, or rearrangement of atoms; it does not result in their creation or destruction.

In brief, an atom is the smallest particle of an element that maintains its chemical identity throughout all chemical and physical changes. Most of the earlier findings and concepts related to law of conservation of mass and law of constant proportions (Fig. 2.4) could be explained to a great extent. Dalton's theory also predicted the **law of multiple proportions**. However, today we know that atoms are not truly indivisible; they are themselves made up of particles (protons, neutrons, electrons, etc), which you will learn later on.

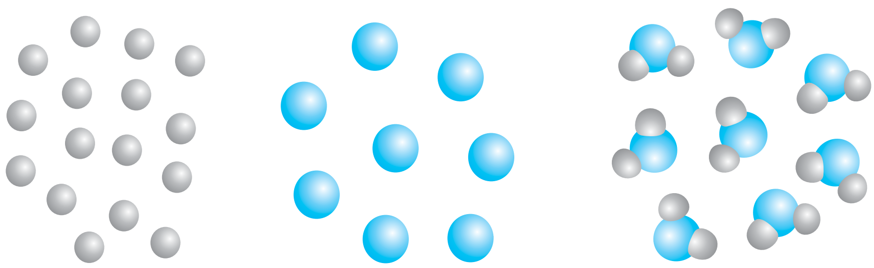

Fig. 2.4 Law of constant proportions

Modern technology has made it possible to take photograph of atoms. The scanning tunnelling microscope (STM) is a very sophisticated instrument. It can produce image of the surfaces of the elements, which show the individual atoms (Fig.2.5).

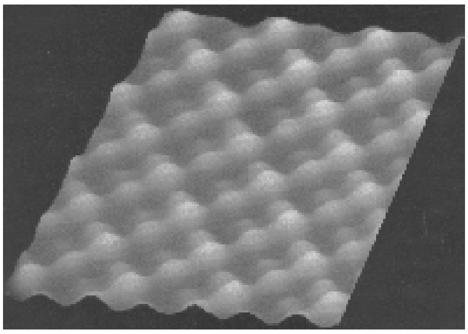

Fig.2.5 Image from a scanning tunneling microscope

Now let us see how atoms and molecules are related with each other.

2.4 ATOMS AND MOLECULES

We have just seen, the first chemist to use the name '**atom**' was John Dalton**.** Dalton used the word **'atom'** to mean the smallest particle of an element. He then went on explaining how atoms could react together to form molecules; which he called 'compound atoms'. Today we know what a molecule is. *A molecule is an aggregate of two or more than two atoms of the same or different elements in a definite arrangement held together by chemical forces or chemical bonds. We can also define a molecule as smallest particle of an element or of a compound which can exist alone or freely under ordinary conditions and shows all properties of that substance (element or compound).* A molecule will be **diatomic** if there are two atoms, for example, chlorine (Cl_2), carbon monoxide, CO; will be triatomic if there are three atoms, for example, water (H_2O) or carbon dioxide, (CO_2), will be tetratomic and pentatomic if there are four and five atoms respectively. In general, a molecule having atoms more than four will be called **polyatomic.** There are eight atoms in a molecule of sulphur and nine atoms in a molecule of ethyl alcohol and we write formulas as S_8 and C_2H_5OH respectively (Fig. 2.6). Only a few years back, a form of carbon called buck-minsterfullerene having molecular formula, C^{60} was discovered. The details you will study in lesson 20.

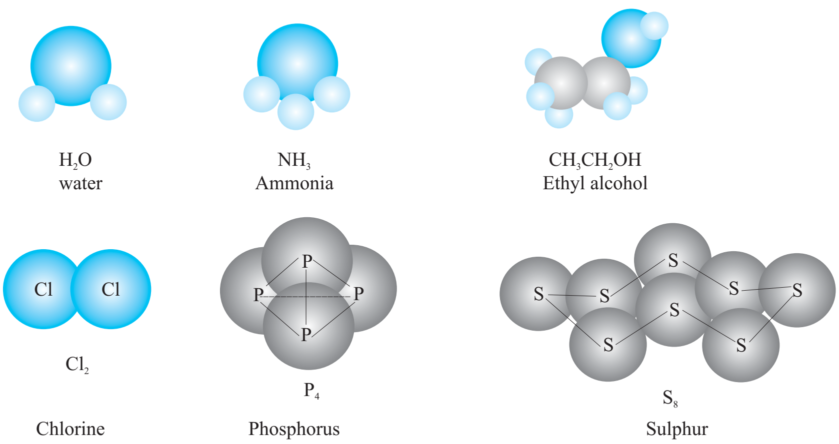

Fig. 2.6 Atomic structure of some molecules

2.5 CHEMICAL FORMULAE OF SIMPLE COMPOUNDS

A molecule is represented by using symbols of elements present in it. This representation is called **molecular formula** of the compound. Thus, a molecular formula of a substance tells us how many atoms of each kind are present in one molecule. In Fig. 2.6, you will find that atoms in a molecule are not only connected in definite ways but also exhibit definite spatial arrangements. *Properties of molecules depends upon the ways atoms are connected and on spatial configuration of the molecules.* CO_2 and H_2O both are triatomic molecules but they have entirely different properties. CO_2 is a linear molecule and is a gas but H_2O is a bent molecule and a liquid. Sodium chloride (common salt) contains equal number of sodium and chlorine atoms and is represented by the formula, NaCl. Sulphuric acid, H_2SO_4 contains three elements : hydrogen, oxygen and sulphur.

2.5.1 Valency and formulation

Every element has a definite capacity to combine with other elements. *This combining capacity of an element is called its valency.* In normal course, hydrogen has 1, oxygen has 2, nitrogen has 3 and carbon has 4 valency. Valency of an element depends upon how it combines with other elements. This will depend upon the nature of the element. Sometime an element shows more than on valency. We say element has *variable valency.* For example, nitrogen forms several oxides: N_2O, N_2O_2, N_2O_3, N_2O_4 and N_2O_5. If we take valency of oxygen 2 then valency of nitrogen in these oxides will be 1,2,3,4 and 5 respectively. Very soon, you will learn in lesson 3 that valency of an element depends on its electronic configuration. Valency of F, Cl, Br and I is normally taken as 1. In NaCl, valency of Na is also 1. All alkali metals such as K, Cs, Rb have 1 valency. Valency of oxygen is 2 and that of phosphorus is 5, we can write the formula of phosphorus pentaoxide as P_2O_5.

Thus, we can write the formula of water (H_2O), ammonia (NH_3), carbon dioxide (CO_2), magnesium oxide (MgO), phosphorus pentaoxide (P_2O_5), hydrochloric acid gas (HCl), phosphorus tribromide (PBr_3) etc. if we know the elements constituting these compounds and their (elements) valencies. Since valencies are not always fixed (as P has different valencies in P_2O_5 and in PBr_3 in the above example), sometimes we face problem. Writing formula of a compound is easy only in **binary compounds** (i.e. compound made of only two elements). However, when we have to write formula of a compound which involves more than two elements (i.e. of polyatomic molecules), it is somewhat cumbersome task. You will learn later on that basically there are two types of compounds: **covalent compounds** and **ionic compounds**. Covalent compounds are of the type H_2O, NH_3 etc. An electrovalent or ionic compound is made of two charged constituents. One positively charged called **'cation'** and other negatively charged called **'anion'**. Here again we should know the charge (valency) of both types of ions for writing formula of an ionic compound. Compounds like sodium nitrate ($NaNO_3$), potassium chloride (KCl), potassium sulphate (K_2SO_4), ammonium choride (NH_4Cl), sodium hydroxide (NaOH) etc. are made of two or more than two elements. For writing the formula of the compounds we should know the charge (valency) of positively and negatively charged constituents of the compounds in such cases. Remember in an ionic compound, sum of the charges of cation and anion should be zero. A few examples of cation and anions along with their valency are provided in Table 2.2.

Table 2.2 Valency of some common cations and anions which form ionic compounds

Anions	Valency	Cations	Valency
Chloride ion, Cl^-	-1	Potassium ion, K^+	+1
Nitrate ion, NO^-_3	-1	Sodium ion, Na^+	+1
Carbonate ion,	-2	Magnesium ion, Mg^{2+}	+2
Sulphate ion, SO_4^{2-}	-2	Calcium ion, Ca^{2+}	+2
Bicarbonate ion, HCO^-_3	-1	Aluminium ion, Al^{3+}	+3
Hydroxide ion, OH^-	-1	Lead ion, Pb^{2+}	+2
Nitrite ion, NO^-_2	-1	Iron ion, Fe^{3+}(Ferric)	+3
Phosphate ion, $PO^3{}_4^-$	-3	Iron ion (Ferrous) Fe^{2+}	+2
Acetate ion CH_3COO^-	-1	Zinc ion, Zn^{2+}	+2
Bromide ion, Br^-	-1	Copper ion, Cu^{2+}	+2
Iodide ion, I^-	-1	Mercury, Hg^{2+} (mercuric)	+2
Sulphide ion, S^{2-}	-2	Ammonium ion, NH_4^+	+1

Suppose you have to write the formula of potassium sulphate which is an electrovalent compound and made of potassium and sulphate ions. Here, charge on potassium ion is +1 and that on sulphate ion is –2. Therefore, for one sulphate ion two potassium ions will be required. We can write,

$$[K^+]_2 \ [SO_4^{2-}]_1 \ = K_2SO_4$$

Similarly for writing formula of sodium nitrate, charge (valency) of sodium ion is +1 and that of nitrate ion is -1, therefore, for one sodium ion, one nitrate ion will be required and we can write.

$$[Na^+]_1 \ [NO_3^-]_1 \ = NaNO_3$$

Now, it is clear that digit showing charge of cation goes to anion and digit showing charge of anion goes to cation. For writing formula of calcium phosphate we take charge of each ion into consideration and write the formula as discussed above as,

$$[Ca^{2+}]_3 \ [PO^{-3}{}_4]_2 = Ca_3(PO_4)_2$$

Writing formula of a compound comes by practice therefore write formula of as many ionic compounds as possible based on the guidelines given above.

2.5.2 Empirical and molecular formula

Molecular formula of a substance is not always identical with the simplest formula that expresses the relative numbers of atoms of each kind in it. **Simplest formula** of an element is expressed by using its symbol as O for oxygen, S for sulphur, P for phosphorus and Cl for chlorine. Molecular formula of these substances are O_2, S_8, P_4 and Cl_2 respectively. The simplest formula of a compound is called its **empirical formula**. *The empirical formula of a compound is the chemical formula that shows the relative number of atoms of each element in the simplest ratio.* In contrast, the molecular formula tells us the actual number of atoms of each element in a molecule. It may be the same as the empirical formula or some other integral multiple of the empirical formula. Empirical and molecular formulae of a few compounds are given in Table 2.3.

Table 2.3: Empirical and molecular formulae

Substance	Empirical formula	Molecular formula
water	H_2O	H_2O
ammonia	NH_3	NH_3
ethane	CH_3	C_2H_6
hydrogen peroxide	HO	H_2O_2
carbon dioxide	CO_2	CO_2
hydrazine	NH_2	N_2H_4

Formula of an ionic substance is always an empirical formula. For example, NaCl is empirical formula not a molecular formula of sodium chloride. You will study later on that ionic substances do not exist in molecular form.

CHECK YOUR PROGRESS 2.2

1. Give one evidence of modern technology which supports Dalton's atomic theory.
2. Write formula of the following compounds
   (i) Ferric phosphate
   (ii) Barium chloride
   (iii) Calcium carbonate
   (iv) Phosphorous tribromide
   (v) Magnesium sulphate
   (vi) Sodium phosphate
   (vii) Sulphur trioxide
3. Write differences between an atom and a molecule.
4. Write empirical formulae of the following molecules:

   C_2H_4, HCl, HNO_3

2.6 LAWS OF CHEMICAL COMBINATIONS

French chemist, **Antoine Laurent Lavoisier** (1743-1794) experimentally showed that *matter can neither be created nor destroyed in a chemical reaction.* This experimental finding was known as **law of conservation of mass**. In fact, this could be possible due to precise measurement of mass by Lavoisier. Law of conservation of mass helped in establishing the **law of definite composition** or **law of constant proportions**. This law states that *any sample of a pure substance always consists of the same elements combined in the same proportions by mass.* For instance, in water, the ratio of the mass of hydrogen to the mass of oxygen is always 1:8 irrespective of the source of water. Thus, if 18.0 g of water are decomposed, 2.0 g of hydrogen and 16.0 g of oxygen are always obtained. Also, if 2 g of hydrogen are mixed with 16.0 g of oxygen and mixture is ignited, 18.0 g of water are obtained after the reaction is over. In the water thus formed or decomposed, hydrogen to oxygen mass ratio is always 1:8. Similarly in ammonia (NH_3), nitrogen and hydrogen will always react in the ratio of 14:3 by mass.

John Dalton thought about the fact that an element may form more than one compound with another element. He observed that for a given mass of an element, the masses of the other element in two or more compounds are in the ratio of simple whole number or integers. In fact this observation helped him in formulation of his fundamental theory popularly known as Dalton's 'Atomic theory' which is discussed in Section 2.3. Let us take two compounds of nitrogen and hydrogen : (i) ammonia (NH_3) and (ii) hydrazine (N_2H_4). In ammonia, as discussed above, 3.0 g of hydrogen react with 14 g of nitrogen. In hydrazine, 4.0 g of hydrogen react with 28 g of nitrogen or 2.0 g of hydrogen reacts with 14.0 g of nitrogen. It can be seen that for 14 g of nitrogen, we require 3.0 g of hydrogen in NH_3 and 2.0 g of hydrogen in hydrazine (N_2H_4). This leads to the ratio

That is, masses of hydrogen which combine with the fixed mass of nitrogen in ammonia and in hydrazine are in the simple ratio of 3:2. This is known as **law of multiple proportions**.

2.6.1 Gay Lusaac's law of combining volume and Avogadro's hypothesis

The French chemist **Gay Lusaac** experimented with several reactions of gases and came to the conclusion that the volume of reactants and products in large number of gaseous chemical reactions are related to each other by small integers provided the volumes are measured at the same temperature and pressure. For example, in reaction of hydrogen gas with oxygen gas which produces water vapour, it was found that two volumes of hydrogen and one volume of oxygen give two volumes of water vapour

To be more specific, if 100 mL of H_2 gas combines with exactly 50 mL of O_2 gas we shall obtain 100 mL of H_2O vapour provided all the gases are measured at the same temperature and pressure (say 100 ^oC and 1 atm pressure).

As you know, the law of definite proportions is with respect to mass. Gay Lussac's findings of integer ratio in volume relationship is actually the **law of definite proportions by volume.**

The Gay Lussac's law was further explained by the work of Italian physicist and lawyer Amedeo Avogadro in 1811. Avogadro's hypothesis which was experimentally established and given the status of a law later on, states as follows:

The volume of a gas (at fixed pressure and temperature) is proportional to the number of moles (or molecules of gas present). Mathematically we can express the statement as

$$V \propto n$$

You will study in section 2.8 that 1 mole of a substance is 6.022×10^{23} particles/molecule of that substance.

Where V is volume and n is the number of moles of the gas. (It is clear from the relationship that more volume will contain more number of molecules). **Avogadro's law** can be stated in another simple way

"Equal volumes of all gases under the same conditions of temperature and pressure contain the same number of molecules"

For example,

$2H_2(g)$	+	$O_2(g)$	$\rightarrow$	$2H_2O(g)$
2 volumes		1 volume		2 volumes (Gay Lussac's law)
2 mol of H_2		1 mol of O_2		2 mol of H_2O (Avogadro's law)

Multiplying both sides of equation by the same number, equation does not change. Now let us multiply by 6.022×10^{23}, we get

	$2 \times 6.022 \times 10^{23}$ molecules of H_2	$1 \times 6.022 \times 10^{23}$ molecules of O_2	$2 \times 6.022 \times 10^{23}$ molecules of H_2O

Similarly,

	H_2 (g) +	Cl_2 (g) →	2HCl(g)
	1 volume	1 volume	2 volume
	1 mol of H_2	1 mol of Cl_2	2 mol of HCl
or	6.022×10^{23} molecules of H_2	6.022×10^{23} molecules of Cl_2	$2 \times 6.022 \times 10^{23}$ molecules of HCl

Experimentally, it has been found that at standard temperature (0 °C) and standard pressure (1 bar) volume of 1 mol of most of the gases is 22.7 litres. Since this volume is of 1 mol of a gas, it is also called **molar volume**. Volume of liquids and solids does not change much with temperature and pressure and same is true with its molar volume. If we know molar mass and density of a solid or of a liquid, we can easily calculate its molar volume by the relationship, volume $= \dfrac{\text{molar mass}}{\text{density}}$

2.7 ISOTOPES AND ATOMIC MASS

As you might have read in your earlier classes that an atom consists of several fundamental particles: electrons, protons and neutrons. An electron is negatively charged and a proton is positively charged particle. Number of electrons and protons in an atom is equal. Since charge on an electron is equal and opposite to charge of a proton, therefore, *an atom is electrically neutral.* Protons remain in the *nucleus* in the centre of the atom and nucleus is surrounded by negatively charged electrons.

The number of protons in the nucleus is called **atomic number**, and is denoted by Z. There are also neutral particles in the nucleus and they are called **neutrons**. Mass of a proton is nearly equal to the mass of neutron. Total mass of nucleus is equal to the sum of masses of protons and neutrons. The total number of protons and neutrons is called **mass number** or the **nucleons number** denoted by A. By convention, atomic number is written at the bottom left corner of the symbol of the atom and mass number is written at the top left corner. For example, we write, $^{4}_{2}He$, $^{7}_{3}Li$ and $^{12}_{6}C$for helium, lithium and carbon respectively. The symbol $^{12}_{6}C$ indicates that there is a total of 12 particles (nucleons) in the nucleus of carbon atom, 6 of which are protons. Thus, there must be 12 – 6 = 6 neutrons. Similarly, $^{16}_{8}O$ has 8 protons and 8 electrons and there are 8 neutrons. Also atomic number, Z differentiates the atom of one element from the atoms of another. Also *an element may be defined as a substance whose atoms have the same atomic number*. Thus, all atoms of an element have nuclei containing the same number of protons and having the same charge. But the nuclei of all the atoms of a given element do not necessarily contain the same number of neutrons. For example, atoms of oxygen, found in nature have the same number of protons which makes it different from other elements, but their neutrons are different. This is the reason that the masses of the atoms of the same elements are different. For example, one type of oxygen atom contains 8 protons and 8 neutrons in the nucleus, second type 8 protons and 9 neutrons and third type contains 8 proton and 10 neutrons. We represent them as $^{16}_{8}O$, $^{17}_{8}O$ and $^{18}_{8}O$. *Atoms of an element that have same atomic number(Z) but different mass number(A) are called isotopes.*

2.7.1. Atomic mass

The mass of an atom is related to the number of protons, electrons, and neutrons it has. Atom of an element is extremely small and therefore it is not easy to weigh it. No doubt, it is possible to determine the mass of one atom relative to another experimentally. For this, it is necessary to assign a value to the mass of one atom of a given element so that it can be used as a **standard**. Scientists agreed to chose an atom of carbon isotope (called carbon-12). Carbon-12 has six protons and six neutrons and has been assigned a mass of exactly 12 atomic mass unit (amu now known as u). Thus one **atomic mass unit** is defined as a mass exactly equal to one twelfth of the mass of one carbon-12 atom.

$$\text{Mass of one carbon-12 atom} = 12 \text{ amu or } 12 \text{ u}$$

$$\text{or } 1\text{amu} = \frac{\text{mass of one carbon atom}}{12}$$

Mass of every other element is determined relative to this mass. Further, it has been found by experiment that hydrogen atom is only 0.0840 times heavier than C-12 atoms. Then on carbon-12 scale, atomic mass of hydrogen = 0.0840×12.00 u = 1.008 u.

Similarly, experiment shows that an oxygen atom is, on the average, 1.3333 times heavier than C-12 atom. Therefore,

Atomic mass of oxygen = 1.3333×12.00 u = 16.0 u

Atomic mass of a few elements on C-12 scale is provided in Table 2.4.

If you see Table 2.4, you will find that atomic mass is not a whole number. For example, atomic mass of carbon is not 12 u but 12.01 u. This is because most naturally occurring elements (including carbon) have more than one isotope. Therefore when we determine atomic mass of an element we generally measure or calculate average mass of the naturally occurring mixture of isotopes. Let us take one example. Carbon has two natural isotope C-12 and C-13 and their natural abundance is 98.90 per cent*, 1.10 per cent, respectively . Atomic mass of C-13 has been determined to be 13.00335 u. Therefore, average atomic mass of carbon

= (0.9890) (12.000 u) + (0.010) (13.00335 u)

= 11.868 + 0.1430 = 12.01 u

Thus, '*atomic mass*' of an element means average atomic mass of that element. These days actual masses of atoms have been determined experimentally using **mass spectrometer**. You will learn about this in your higher classes.

Table 2.4 Atomic masses* of some common elements

Element	Symbol	Mass(u)	Element	Symbol	Mass(u)
Aluminium	Al	26.98	Magnesium	Mg	24.31
Argon	Ar	39.95	Manganese	Mn	54.94
Arsenic	As	74.92	Mercury	Hg	200.59
Barium	Ba	137.34	Neon	Ne	20.18
Boron	B	10.81	Nickel	Ni	58.71
Bromine	Br	79.91	Nitrogen	N	14.01
Caesium	Cs	132.91	Oxygen	O	16.00
Calcium	Ca	40.08	Phosphorus	P	30.97
Carbon	C	12.01	Platinum	Pt	195.09
Chlorine	Cl	35.45	Potassium	K	39.10
Chromium	Cr	52.00	Radon	Rn	(222)**
Cobalt	Co	58.93	Silicon	Si	28.09
Copper	Cu	63.54	Silver	Ag	107.87
Fluorine	F	19.00	Sodium	Na	23.00
Gold	Au	196.97	Sulphur	S	32.06
Helium	He	4.00	Tin	Sn	118.69
Hydrogen	H	1.008	Titanium	Ti	47.88
Iodine	I	126.90	Tungston	W	183.85
Iron	Fe	55.85	Uranium	U	238.03
Lead	Pb	207.19	Vanadium	V	50.94
Lithium	Li	6.94	Xenon	Xe	131.30
			Zinc	Zn	65.37

* During calculation we convert per cent into fraction by dividing by 100. Thus, 98.90 per cent becomes 0.9890.

*Atomic masses are average atomic masses. They are given correct up to second decimal places. In practice, we use round figures and for this rounding off is necessary.

**Radioactive

2.7.2 Molecular mass

You have just read that a molecule can be represented in form of a formula popularly known as **molecular formula**. Molecular formula may be of an element or of a compound. Molecular formula of a compound is normally used for determing the **molecular mass** of that substance. If the substance is composed of molecules (for example, CO_2, H_2O or NH_3), it is easy to calculate the molecular mass. *Molecular mass is the sum of atomic masses of all the atoms present in that molecule.* Thus the molecular mass of CO_2 is obtained as

C 1 × 12.0 u = 12.0 u
2O 2 × 16.0 u = 32.0 u
CO_2 Total = 44.0 u

We write molecular mass of CO_2 = 44.0 u

Similarly, we obtain molecular mass of ammonia, NH_3 as follows :

N 1 × 14.0 u = 14.0 u
3H 3 × 1.08 u = 3.24 u
NH_3 Total = 17.24 u

Molecular mass of ammonia, NH_3 = 17.24 u.

For substances which are not molecular in nature, we talk of **formula mass**. For example, sodium chloride, NaCl is an ionic substance. For this, we write formula mass which is calculated similar to molecular mass. In case of NaCl, formula mass = mass of 1 Na atom + mass of 1 Cl atom = 23 u + 35.5 u = 58.5 u.

You will learn about such compounds later on in your lesson 5.

CHECK YOUR PROGRESS 2.3

1. Silicon has three isotopes with 14, 15 and 16 neutrons respectively. What is the mass number and symbol of these three isotopes?
2. Calculate molecular mass of the following compounds C_3H_8, PCl_5, SO_3

2.8 MOLE CONCEPT

When we mix two substances, we get one or more new substance(s). For example when we mix hydrogen and oxygen and ignite the mixture, we get a new substance water. This can be represented in the form of an equation,

$$2H_2\,(g) + O_2(g) \rightarrow 2H_2O\,(l)$$

In above equation, 2 molecules (4 atoms) of hydrogen react with 1 molecule (2 atoms) of oxygen and give two molecules of water. Similarly, we always like to know how many atoms/molecules of a particular substance would react with atoms/molecules of another substance in a chemical reaction. No matter how small they are. The solution to this problem is to have a convenient unit of matter that contains a known number of particles (atoms /molecules). The chemical counting unit that has come into use is the **mole.**

The word mole was apparently introduced in about 1896 by Wilhelm Ostwald who derived the term from the Latin word 'moles' meaning a 'heap' or 'pile'. The mole whose symbol is 'mole' is the SI base unit for measuring **amount of substance**. It is defined as follows:

'A mole is the amount of pure substance that contains as many particles (atoms, molecules, or other fundamental units) as there are atoms in exactly 0.012 kg of C-12 isotope'.

In simple terms, mole is the number of atoms in exactly 0.012 kg (12 grams) of C-12. Although mole is defined in terms of carbon atoms but the unit is applicable to any substance just as 1 dozen means 12 or one gross means 144 of any thing. Mole is scientist's counting unit like dozen or gross. By using mole, scientists (particularly chemists) count atoms and molecules in a given substance. Now it is experimentally found that the number of atoms contained in exactly 12 grams of C-12 is 602,200 000 000 000 000 000 000 or 6.022×10^{23}. This number (6.022×10^{23}) is called **Avogadro constant** in honour of **Amedeo Avogadro** an Italian lawyer and physicist and is denoted by symbol, N_A. We have seen that

Atomic mass of C = 12 u

Atomic mass of He = 4 u

We can see that one atom of carbon is three times as heavy as one atom of helium. On the same logic 100 atoms of carbon are three times as heavy as 100 atoms of helium. Similarly 6.02×10^{23} atoms of carbon are three times as heavy as 6.02×0^{23} atoms of helium.

But 6.02×10^{23} atoms of carbon weigh 12 g, therefore 6.02×10^{23} atoms of helium will weigh $1/3 \times 12g = 4g$. We can take a few more examples of elements and can calculate the mass of one mole atoms of the element. Numerically it is equal to its atomic mass expressed in gram. Mass of one mole of a substance is called its **molar mass**. Mass of one mole atoms of oxygen will be 16 g. Mass of one mole of fluorine will be 19 g. Now if we take mass of one mole molecule of oxygen it would be 32 g because there are two atoms in a molecule of oxygen (O_2). When we do not mention atom or molecule before mole, we always mean one mole of that substance in its natural form. For example, if we simply say one mole of oxygen, it means that we are referring one mole molecule of oxygen as oxygen occurs in nature as molecular oxygen. If we take an example of a molecule of a compound, we find that same logic is applicable. For example, mass of one mole molecule of water will be 18 g as molecular mass of water is 18u.

Remember molar mass is always expressed as grams per mole or g /mol or g mol^{-1}. For example,

Molar mass of oxygen (O_2) = 32 g mol^{-1}

Molar mass of lead (Pb) = 207 g mol^{-1}

We have just seen in Section 1.6 that atoms of two different elements combine with one another in the ratio of small whole number. A modern interpretation of this observation is that atoms or molecules combine with one another in the ratio of 1:1, 1:2 or 1:3 or any other simple ratio i.e. they combine 1 mol for 1 mol or 1 mol for 2 mol or 1 mol for 3 mol, and so on. Thus mole concept is the cornerstone of quantitative science for chemical reactions which you will study in your higher classes.

Table 2.5 Molecular and molar mass of some common substances

Formula	Molecular mass(u)	Molar mass (g/mol)
O_2	32.0	32.0
H_2	2.0	2.0
Cl_2	71.0	71.0
P_4	123.9	123.9
CH_4	16.0	16.0
CH_3OH	32.0	32.0
NH_3	17.0	17.0
CO_2	44.0	44.0
HCl	36.5	36.5
C_6H_6	78.0	78.0
SO_2	64.0	64.0
CO	28.0	28.0
C_2H_5OH	46.0	46.0

Example 2.1: How many grams are there in 3.5 mol of sulphur?

Solution: For converting mass into mole and vice visa, we always need the molar mass. Molar mass of sulphur is 32.0 g mol^{-1}. Therefore, number of grams of sulphur in 3.50 mol of sulphur is

$$3.50 \text{ mol sulphur} \times \left(\frac{32.0g}{1\,\text{mol}}\right) = 112.0 \text{ g sulphur}$$

Example 2.2: Calculate number of moles present in 48 g of oxygen.
Solution: Molar mass of oxygen = 32 g mol^{-1}
Oxygen in natural form will be molecular oxygen, O_2

$$\text{Therefore, number of moles of oxygen} = \left(\frac{48g}{32\,\text{g mol}^{-1}}\right) = 1.5 \text{ mol}$$

CHECK YOUR PROGRESS 2.4

1. Sulphur is a non-metallic element. How many atoms are present in 16.3 g of S?
2. Molar mass of silver is 107.9 g. What is the mass of one atom of silver?

2.9 CHEMICAL EQUATIONS

A chemical equation is a shorthand description of a reaction carried out in a laboratory or elsewhere. It gives the formulas for all the reactants and products. For example

$$C + O_2 \rightarrow CO_2 \qquad \text{..... (1)}$$
$$2H_2 + O_2 \rightarrow 2H_2O \qquad \text{.......(2)}$$

In a chemical reaction reactants are written on the left and products are written on the right side of the arrow. Arrow (→) indicates conversion of reactant(s) into product(s). In a chemical reaction atoms are neither created nor destroyed. This is known as **law of conservation of mass.** A chemical equation, therefore, should be consistent with this law. Total number of atoms of each element must be the same in the products and in the reactants. As shown in equation (2) above two molecules (four atoms) of hydrogen react with one molecule (two atoms) of oxygen and give two water molecules in which there are four hydrogen atoms and two oxygen atoms. Since number of atoms of the involved elements is equal on both side of the arrow in the equation, we say the equation is **balanced.** A balanced chemical equation is quite meaningful in science (chemistry) as it gives a lots of information. In order to make an equation more informative, we also indicate the physical states of the reactants and products. We write in parenthesis 's' if the substance is solid, '*l*' if the substance is liquid and 'g' if the substance is a gas. Accordingly, equation (1) and (2) can be written as,

$$C\,(s) + O_2\,(g) \rightarrow CO_2\,(g)$$
$$2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(l)$$

2.9.1 Balancing of a chemical equation

Balancing of a chemical equation is essential as we can derive meaningful information from this. Before balancing a chemical equation, please ensure that correct formulas of *reactants* and *products* are known. Let us consider burning of methane in oxygen to give carbon dioxide and water. First write reactants and products,

$$\underbrace{CH_4 + O_2}_{\text{reactants}} \rightarrow \underbrace{CO_2 + H_2O}_{\text{products}} \quad \text{(unbalanced equation)}$$

In this equation, hydrogen and carbon appear in only two formulas each, while oxygen appears three times. So we begin by balancing the number of carbon and hydrogen atoms. Here if we examine both sides, carbon appears in methane on left and in carbon dioxide on the right side. Therefore, all carbon in methane, CH_4, must be converted to carbon dioxide. One molecule of CH_4, however, contains four hydrogen atoms, and since all the hydrogen atoms end up in water molecule, two water molecules must be produced for each methane molecules. Therefore, we must place coefficient 2 in front of the formula for water to give

$$CH_4 + O_2 \rightarrow CO_2 + 2H_2O \quad \text{(unbalanced)}$$

Now we can balance oxygen, since there are four oxygen atoms on right hand side of equation (two in CO_2 and two in two molecules of H_2O). Therefore, we must place 2 in front of the formula for oxygen, O_2. By doing this we get equal atoms of oxygen on both sides of equation.

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O \qquad \text{(balanced)}$$

Now number of atoms of each element is equal on both sides of the chemical equation. In order to make the chemical equation more informative, indicate states of each reactant and product.

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

Balancing of equation comes only by practice and therefore let us take one example.

Example2.3: Bottled gas sold as cooking gas contains butane, C_4H_{10} as the major component. Butane when burns in sufficient oxygen (present in air) gives carbon dioxide and water. Write a balanced chemical equation to describe the reaction.

Solution: Work out the balanced equation in steps

Step 1: Write an unbalanced equation showing correct formulas of reactants and products

$$\underset{\text{butane}}{C_4H_{10}} + \underset{\text{oxygen}}{O_2} \rightarrow \underset{\substack{\text{Carbon}\\\text{dioxide}}}{CO_2} + \underset{\text{water}}{H_2O} \quad \text{(unbalanced equation)}$$

Now balance C and H as they appear only in two places.

Step II: Balance the number of C atoms.

Since 4 carbon atoms are in the reactant, therefore, $4CO_2$ must be formed.

$$C_4H_{10} + O_2 \rightarrow 4CO_2 + H_2O \quad \text{(unbalanced)}$$

Step III: Balance the number of hydrogen atoms

There are 10 hydrogen atoms in butane and each water molecule requires 2 hydrogen atoms, therefore, 5 water molecules will be formed.

$$C_4H_{10} + O_2 \rightarrow 4CO_2 + 5H_2O \quad \text{(unbalanced)}$$

Step IV: Balance the number of O atoms

There are 8 oxygen atoms in the carbon dioxide and 5 oxygen atoms with H_2O molecules. Therefore, 13 atoms or 13/2 molecules of oxygen will be required.

$$C_4H_{10} + 13/2\ O_2 \rightarrow 4CO_2 + 5H_2O$$

Normally we do not write fractional coefficient in equation as one may interpret that molecules can also be available in fraction. Therefore, we multiply both sides by 2 and get the final balanced equation

$$2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O \qquad \text{(balanced)}$$

We can also write states of the substances involved.

$$2C_4H_{10}\,(g) + 13O_{2\,(g)} \rightarrow 8CO_2\,(g) + 10H_2O(l)$$

Remember :

(i) Use the simplest possible set of whole number coefficients to balance the equation.

(ii) Do not change subscript in formulas of reactants or products during balancing as that may change the identity of the substance. For example, $2NO_2$ means two molecules of nitrogen dioxide but if we double the subscript we have N_2O_4 which is formula of dinitrogen tetroxide, a completely different compound.

(iii) Do not try to balance an equation by arbitrarily selecting reactant(s) and product(s). A chemical equation represents a chemical reaction which is ***real.*** Thus real reactants and products only can be taken for balancing.

2.9.2 Uses of balanced equations

A balanced chemical equation gives a lot of meaningful information. First it gives the number of atoms and molecules taking part in the reaction and corresponding masses in atomic mass units (amu or u). Second it gives the number of moles taking part in the reaction, with the corresponding masses in grams or in other convenient units.

Let us consider the reaction between hydrogen and oxygen once again

$2H_2$ (g)	+	O_2 (g)	$\rightarrow$	$2H_2O(l)$
2 molecules of hydrogen		1 molecule of oxygen		2 molecules of water
4.0 u		32.0 u		36 u

But in normal course we deal with a large number of molecules, therefore, we can consider the above reaction as follows:

$2H_2$	+	O_2	$\rightarrow$	$2H_2O$
2 molecules of hydrogen		1 molecule of oxygen		2 molecules of water

Suppose we multiply entire chemical equation by 100, we can write

2×100 molecules of hydrogen	+	1×100 molecules of oxygen	$\rightarrow$	2×100 molecules of water

If we multiply entire equation by Avogadro constant, 6.022×10^{23} , we get

$2 \times 6.022 \times 10^{23}$ molecules molecules of hydrogen	+	$1 \times 6.022 \times 10^{23}$ molecules of oxygen	$\rightarrow$	$2 \times 6.022 \times 10^{23}$ of water

Since 6.022×10^{23} molecules is 1 mole, therefore, we can also write

or 2 mol of hydrogen + 1 mol of oxygen → 2 mol of water

Therefore, equation can be written as

$$\underset{\text{2 mol of hydrogen}}{2H_2} + \underset{\text{1 mol of oxygen}}{O_2} \rightarrow \underset{\text{2 mol of water}}{2H_2O}$$

Or 4.0 g of hydrogen + 32.0 g of oxygen → 36 g of water

Thus a chemical equation can also be interpreted in terms of masses of reactants consumed and product(s) formed. This relationship in chemical reaction is very important and provides a quantitative basis for taking definite masses of reactants to get a desired mass of a product.

Example 2.4: In the reaction

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

How much CO_2 will be formed if 80 g of methane gas (CH_4) is burnt?

Solution:

	$CH_4(g)$ +	$2O_2(g)$ →	$CO_2(g)$ +	$2H_2O(l)$
	1 mol	2 mol	1 mol	2 mol
or	16 g	64 g	44 g	36 g

We can see in the above equation that 16 g of CH_4 gives 44 g of CO_2.

Therefore, for getting 80 g of CH_4, the mass of CO_2 required will be

$$= \frac{44\,g \times 80\,g}{16\,g} = 44 \times 5\,g = 220\,g \text{ of } CO_2$$

CHECK YOUR PROGRESS 2.5

1. Balance the following equations.
   (i) $H_3PO_3 \rightarrow H_3PO_4 + PH_3$
   (ii) $Ca + H_2O \rightarrow Ca(OH)_2 + H_2$
   (iii) $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$
2. Name the following compounds.
   Na_2O, Cu_2Cl_2, BaO, Na_2SO_4

LET US REVISE

- Matter is anything that has mass and occupies space. It can be classified on the basis of its (i) physical state as solid, liquid or gas, and (ii) chemical composition/constitution as element, compound or mixture.
- An *element* is basic form of matter that cannot be broken down into simpler substances even by chemical reaction. A *compound* is a substance composed of two or more different types of elements chemically combined in a definite proportion by mass. A mixture contains more than one substance (element or compound) mixed in any proportion.

- A *solution* is a homogeneous mixture of two or more than two substances. Major component of the solution is called *solvent*.
- According to *law of constant proportions*, a sample of a pure substance always consists of the same elements combined in the same proportions by mass.
- When an element combines with another element and forms more than one compound, then different masses of one element that combine with a fix mass of another element are in ratio of simple whole number or integer. This is the *law of Multiple proportions*.
- John Dalton introduced the idea of an atom as an indivisible particle of matter. An atom is the smallest particle of an element which can exist and retains all the chemical properties of that element.
- A molecule is the smallest particle of an element or of a compound which can exist freely under ordinary conditions and shows all properties of that substance.
- A molecule can be expressed in form of a *chemical formula* using symbols of constituent elements.
- A *molecular formula* shows the actual number of atoms of different elements in a molecule of an element or of compound. In other words, composition of any compound can be represented by its formula. For writing formula of a compound valence or valency of the elements is used. Valency is *combining capacity* of an element and is related to its electronic configuration.
- An *empirical formula* shows the simplest whole number ratio of the atoms of different elements present in a compound.
- Atoms of the isotope ^{12}C are assigned a atomic mass unit of 12 and the relative masses of all other atoms are obtained by comparison with the mass of a carbon-12.
- The *mole* is the amount of substance which contains the same number of particles (atoms, ions or molecules) as there are atoms in exactly 0.012 kg of ^{12}C.
- *Avogadros constant* is defined as the number of atoms in exactly 12 g of C-12 and is equal to 6.022×10^{23} mol^{-1}.
- Mass of one mole atoms or one mole molecules of a substance is its *molar mass* and volume of one mole of the substance is its *molar volume.*
- A chemical equation is a shorthand description of a reaction. A balance chemical equation provides quantitative information about reactants consumed and products formed in a chemical reaction. A balance chemical equation obeys *law of conservation of mass* and *law of constant proportions.*

TERMINAL EXERCISES

1. There are many examples of homogeneous and heterogeneous mixtures in the world around you. How would you classify: sea-water, air (unpolluted), smoke, black coffee, tea, soil, soda water and wood ash?

2. Characterize gases, liquids and solids in terms of compressibility, fluidity and density.
3. What is atomic theory proposed by Dalton? Describe how it explains the great variety of different substances.
4. Give normal state (solid, liquid or gas) of each of the following:

   (i) Nitrogen (ii) Copper (iii) Bromine (iv) Oxygen

   (v) ethyl alcohol (vi) hydrogen peroxide
5. Label each of the following as a substance, a heterogeneous mixture, or a solution.

(i)	bromine	(iv)	soil (in front of your home)	(vii)	river water
(ii)	petrol	(v)	stone	(viii)	Coal
(iii)	concrete	(vi)	beach sand	(ix)	Soda water
6. Write the number of protons, neutrons and electrons in each of the following:

   $${}^{19}_{9}F,\ {}^{18}_{8}O,\ {}^{40}_{20}Ca$$
7. Give the symbol for each of the following isotopes

   (i) Atomic number 19, mass number 40

   (ii) Atomic number 18, mass number 40

   (iii) Atomic number 7, mass number 15
8. Boron has two isotopes with masses of 10.01294 and 11.00931 u and abundance of 19.77% and 80.23%. What is the average atomic mass of boron?

   (Ans.10.81 u)
9. How does an element differ from a compound? How are elements and compounds different than mixture?
10. How will you define a solution based on its composition?
11. Charge of one electron is 1.6022×10^{-19} coulomb. What is the total charge on 1 mol of electron? If there is same amount of charge on one proton, calculate total charge on 1 mol of protons.
12. How many molecules of O_2 are in 8.00 g of O_2? If the O_2 molecules were completely split into O (oxygen atom), how many moles of atoms of oxygen would be obtained?

    (Ans. Number of molecules in 8 g of O_2 =1.5055×10^{23} molecules

    Number of atoms in 8 g of O_2 = 3.0110×10^{23} atoms)
13. Assume that a human body is 80% water. Calculate the number of the molecules of water that are present in the body of a person who has mass of 65 kg.

    (Ans. 1.7×10^{27} molecules of water)
14. Using atomic masses given in the table of this lesson calculate the molar masses of each of the following compounds:

    CO,CH_4, NaCl, NH_3 and HCl

15. Average atomic mass of carbon is 12.01 u. Find the number of moles of carbon in (i) 2.00 g of carbon and (b) 3.00×10^{21} atoms of carbon.

16. Balance the following equations

    (i) $H_2O_2 \rightarrow H_2O + O_2$

    (ii) $S + O_2 \rightarrow SO_3$

    (iii) $C_2H_2 + O_2 \rightarrow CO + H_2O$

    (iv) $MnO_2 + HCl \rightarrow MnCl_2 + Cl_2 + H_2O$

17. Classify the following molecules as mono, di, tri, tetra, penta and hexatomic molecules.

    H_2, P_4, SF_4, SO_2, PCl_3, C_2H_2, CH_3OH, PCl_5, H_2O_2, HCl, Cl_2O

18. What is meant by molecular formula? Hydrogen peroxide has the molecular formula H_2O_2. What mass of oxygen can be formed from 17 g of H_2O_2 if decomposition of H_2O_2 takes place.

19. Write 'true' or 'false'.

    A balanced chemical equation shows

    (i) the formulas of the products

    (ii) the molar proportions in which the products are formed

    (iii) that a reaction can occur

    (iv) the relative number of atoms and molecules which react

    (v) that a reaction is exothermic

20. What is the mass of

    (i) 6.02×10^{23} atoms of O

    (ii) 6.02×10^{23} atoms of P

    (iii) 6.02×10^{23} molecules of P_4

    (iv) 6.02×10^{23} molecules of O_2

    [Ans. (a) 16.0 g (b) 31.0 g (c) 124.0 g (d) 32 g]

21. How many atoms are there in

    (i) two moles of iron

    (ii) 0.1 mol of sulphur

    (iii) 18 g of water, H_2O

    (iv) 0.44 g of carbon dioxide, CO_2

    [Ans. (a) 1.204×10^{24} (b) 6.02×10^{22} (c) 1.8×10^{24} and (d) 1.8×10^{22}]

22. Define the following

    (i) Law of constant proportions

    (ii) Law of multiple proportions

(iii) Avogadro's Law

(iv) Gay Lussacis Law

(v) Dalton's atomic theory

23. Convert into mole

(i) 12 g of oxygen gas (O_2)

(ii) 20 g of water (H_2O)

(iii) 22 g pf carbon dioxide (CO_2)

(Ans. (a) 0.375 mol (b) 1.11 mol (c) 0.50 mol)

ANSWERS TO CHECK YOUR PROGRESS

2.1 1. (i), (iii), (v) and (vi)

2. (ii)

2.2 1. refer text

2. (i) $FePO_4$ (ii) $BaCl_2$ (iii) $CaCO_3$ (iv) PBr_3 (v) $MgSO_4$ (vi) Na_3PO_4 (v) SO_3

3. refer text

4. CH_2, HCl, HNO_3

2.3 1. $^{28}_{14}Si$, $^{29}_{14}Si$, $^{30}_{14}S$

2. C_3H_8 = 44 u

PCl_5 = 207.5 u

SO_3 = 80 u

2.4 1. 3.08×10^{23} S atom

2. 1.77×10^{-22} g of Ag

2.5 1. (i) $4H_3PO_3 \rightarrow 3H_3PO_4 + PH_3$

(ii) $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$

(iii) $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

2. Sodium oxide, Cuperous chloride, Barium oxide, Sodium sulphate

GLOSSARY

Atomic mass: The average mass of an atom in a representative sample of atoms of an element.

Compound: Matter that is composed of two or more different kinds of elements chemically combined in definite proportions.

Chemical reaction: A process in which substances are changed into other substances through rearrangement/combination of atoms.

Diffusion: The gradual mixing of the molecules of two or more substances owing to random molecular motion.

Element: Matter that is composed of one kind of atoms, each atom of a given kind having the same properties (Mass is one such property).

Heterogeneous mixture: A mixture which has no uniformity in composition.

Homogeneous mixture: A mixture with the same composition throughout

Isotopes: Isotopes are atoms having the same atomic number, Z but different mass number, A.

Mass number: Number of protons plus number of neutrons in the nucleus of an atom of an element.

Matter: Anything that has mass and occupies space.

Mole: Mole is amount of substance that contains as many elementary particles as there are atoms in 0.012 kg of C-12 isotope.

Molar mass: The mass (in gram) of one mole of a substance.

Molar volume: The volume of one mole of a substance.

Molecular mass: The sum of atomic masses (in u) of all the atoms of a molecule.