As we know that sugar, salt, fabric, oxygen, copper, silver, water etc., are all different kinds of matter. We have discussed earlier that what is matter and how the matter is classified into solid liquid and gases on the basis of their physical states or into elements, compounds and mixture on the basis of their chemical properties. Now we will study it in detail.

1. ATOMS AND MOLECULES: ANCIENT THOUGHTS

Ancient Indian and Greek philosophers have always wondered about the unknown and unseen form of matter or we can say that what matter is ultimately made up of. It was around 500 BC that an Indian philosopher Maharishi Kanad had postulated that **"matter is divisible"** i.e. if we go on dividing the matter we will get smaller and smaller particles and ultimately we will get the smallest particle of matter which can not be divided any further. These indivisible particles were named by him as "**Parmanu**".

Around the same era ancient Greek philosopher Democritus and Leuccipus suggested the same idea. However he called the smallest indivisible particles as "atoms" (Greek : means uncutable)

Another Indian philosopher Pakudha Katyayama said that these particles normally exist in a combined form called molecules which give us various forms of matter.

Thus, we may conclude that matter is made up of small particles which may be atoms or molecules. Different kinds of atoms and molecules have different properties that's why different kinds of matter also show different properties.

Till eighteenth century there were no experimental work done to validate these philosophical ideas. By the end of eighteenth century, the fact had been established that pure substances can be either elements or compounds. Scientist became interested in finding out why and how elements combine and what happens when they combine?

Experimental studies were carried out to understand the laws according to which the elements combine to form compounds. These laws are called law of chemical combination.

2. LAWS OF CHEMICAL COMBINATION

As mentioned above, whenever reactants (elements) react together to form a compound they do so according to certain laws. These laws are called "laws of chemical combination". There are two important laws of chemical combination. These are

(i) Law of conservation of mass
(ii) Law of constant proportions

The laws of chemical combination are the experimental laws which have been established by scientists after performing a large number of experiments.

2.1 LAW OF CONSERVATION OF MASS

This law was put forward by **Antoine Lavoisier** in 1774. It deals with the masses of reactants and products involved in a chemical reaction.

It was observed by the scientists that if they carried out a chemical reaction in a closed container then there was no change in mass. This preservation of mass in chemical reaction lead to the formulation of law of conservation of mass (or matter).

"**Matter is neither created nor destroyed in a chemical reaction.**"

OR

"**In any chemical reaction the total mass of the reactants is equal to the total mass of the products.**"

OR

"**There is no change in mass during a chemical reaction.**"

The following experiment illustrate the validity of this law

2.1.1 Experiment : When matter undergoes a physical change

A piece of ice (water) is taken in a small conical flask. It is well corked and weighed. This flask is then heated gently to melt the ice (solid) into water (liquid)

$$\text{Ice} \xrightarrow{\text{Heat}} \text{Water}$$

Observation : After heating, the ice changes into water. When we weighed the flask the mass does not change though a physical change has taken place

Conclusion : From the above experiment, we lead to the conclusion that during a physical change mass does not change or in other words mass is conserved during the physical change.

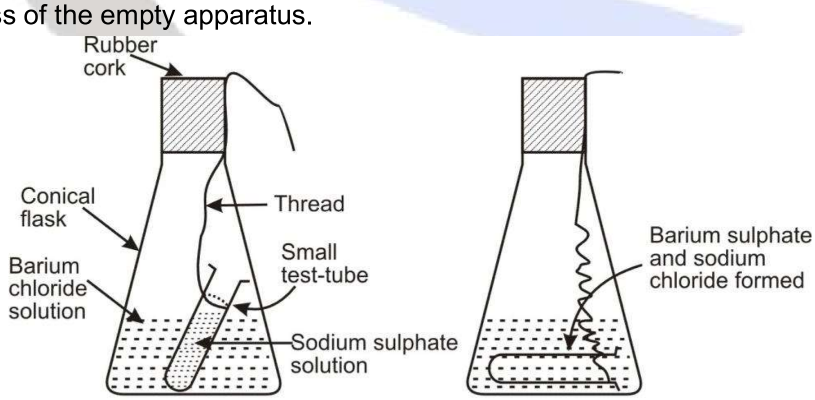
2.1.2 When matter undergoes a chemical change

When barium chloride reacts with sodium sulphate, barium sulphate and sodium chloride is formed.

Barium chloride +	Sodium sulphate ⟶	Barium sulphate +	sodium chloride
(solution)	(solution)	(white ppt)	(solution)

Experiment : Take a clean conical flask fitted with a cork and a small test tube having a long thread tied to its neck. Weigh the flask, cork and tube together to find the initial mass of the apparatus. Now prepare the two solutions.

(i) Approximately 5% solution of barium chloride is prepared by dissolving 5g of barium chloride in 100 ml of water. Take a small amount of barium chloride solution in the conical flask. Put some sodium sulphate solution in the small test tube. Suspend the test tube in the flask with the thread as shown in figure. Weigh the complete system. Subtract the mass of the empty apparatus.



(a) Before reaction (b) After reaction

The difference will give the mass of the reactants taken. Let the mass of the reactants be x.

Now loosen the cork so that the thread is loosened and the test tube fall into the flask due to which sodium sulphate solution mixes with barium chloride solution to form a white precipitate. of barium sulphate and sodium chloride solution (figure). Weigh the complete apparatus again along with the contents. If we subtract the initial mass of the apparatus from this mass, we will get the mass of the products. Suppose the mass of the product is y.

Observation : It was observed that the mass of the reactants (x) comes out to be same as that of the products (y). This is in accordance with the law of conservation of mass.

Conclusion : From the above experiment we lead to the conclusion that during a chemical change, mass remains the same or unchanged or in other words mass is conserved during a chemical change.

2.2 LAW OF CONSTANT PROPORTIONS

Lavoisier, along with other scientists found that many compounds were composed of two or more elements and each such compound had the same elements in the same proportions, irrespective of from where the compound came.

These observations led to the 'law of constant proportions' which is also known as the 'law of definite proportions'. This law deals with the composition of elements present in a given compound. This law was given by J.L. Proust. According to this law

"**In a chemical substance the elements are always present in definite proportions by mass.**"

Or

"**A chemical compound is always made up of the same elements combined together in the same fixed proportion by mass.**"

For example : In water, hydrogen and oxygen combined together in the same fixed proportion of 1 : 8 by mass, irrespective of the source of water (like river, rain or tap water).

If we decompose 9 g of pure water by electrolysis i.e. passing electricity through it, then 1 gm of hydrogen and 8 gm of oxygen are obtained. Now, if we repeat this experiment by taking pure water from different sources the same masses of hydrogen and oxygen elements are obtained.

This experiment shows that water always consists of hydrogen and oxygen combined together in the same constant proportion of 1 : 8 by mass.

Law of constant proportions also help us to calculate the percentage of any element in the given compound using the following expression.

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

3. DALTON'S ATOMIC THEORY

After the above two laws of chemical combination were put forward, the next problem faced by scientist was to give appropriate explanation of these laws. This led John Dalton to put forward a theory in 1808 about the nature of matter. Dalton picked up the idea of divisibility of matter, which was till then just a theoretical idea. The theory is based on certain postulates called postulates (or assumptions) of Dalton's atomic theory. His theory provided an explanation for the law of conservation of mass and the law of definite proportion. According to Dalton's atomic theory all matter, whether an element, a compound or a mixture is composed of small particles called atoms.

3.1 POSTULATES OF DALTON'S ATOMIC THEORY

The main postulates of Dalton's atomic theory are as follows :

(i) All matter is made up of very small particles called atoms.
(ii) Atoms are indivisible particles, which can not be created or destroyed in a chemical reaction.
(iii) Atoms of a given element are identical in all respects i.e. size, shape, mass and chemical properties.
(iv) Atoms of different elements have different size and masses and also posses different properties.
(v) Atoms of the same or different elements combine in the ratio of small whole numbers to form compounds.
(vi) The relative number and kinds of atoms are constant in a given compound.
(vii) Atoms of the same elements or two different elements may combine in different ratios to form more than one compound.

3.2 EXPLANATION OF THE LAWS OF CHEMICAL COMBINATION

Dalton's atomic theory was the first modern attempt to describe the properties of matter in terms of atoms. This theory provides a simple explanation for the laws of chemical combination.

3.2.1 Explanation of law of conservation of mass

According to Dalton's atomic theory matter is made up of atoms and, number of various types of atoms in the products of a chemical reaction is the same as that of the reactants. As the same number of various atoms in products and reactants will have the same mass, so the total mass of products is equal to the total mass of reactants or the mass remain unchanged during a chemical reaction and this is the law of conservation of mass.

3.2.2 Explanation of law of constant proportion

According to one of the postulates of Dalton's atomic theory the number and kind of atoms in a compound is fixed. From this we can infer that a compound is always made up of the same elements combined together in the same proportion by mass and this is the law of constant proportion.

3.3 LIMITATIONS OR DRAWBACKS OF DALTON'S ATOMIC THEORY

With the advancement in scientific studies, Dalton's atomic theory suffered from the following drawbacks :

(i) Atom is no longer considered as the smallest indivisible particle.
(ii) According to Dalton's atomic theory, all the atoms of an element have exactly the same mass. Though, it is now known that atoms of the same elements may have different masses.
(iii) According to Dalton's atomic theory, atoms of different elements have different masses. However it is now known that even atoms of different elements can have the same mass.

(iv) Substances made up of the same kind of atoms may have different properties. For example, charcoal, graphite and diamond are all made up of carbon atoms but have different physical properties.

4. ATOM

As all the houses are made up of bricks similarly all matter whether an element, or a compound or a mixture is made up of the smallest indivisible particles called atoms. In other word atoms are the building blocks of all matter. In chemistry atom is defined as **the smallest particle of an element that can take part in a chemical reaction and which may or may not be capable of free existence**. Atoms of most of the elements are very reactive and do not exist in the free state, instead, they exist in combination with the atoms of the same element or another element. For example, H_2, N_2, O_2, HCl, NH_3, etc.

4.1 HOW BIG ARE THE ATOMS ?

Atoms are very-very small in size. They are so small that they can not be seen even under a microscope. To imagine about their size, it is very much interesting to note that if millions of atoms are stacked one above the other, the thickness produced may not be equal to the thickness of the sheet of a paper.

The size of an atom is indicated by its radius which is called '**atomic radius**". Atomic radius is measured in nanometers, which is represented by the symbol 'nm'.

$1\ nm = 10^{-9}\ m$

$1m = 10^{9}\ nm$

Atomic radii of some common elements

S. No.	Element	Atomic radius
1.	Hydrogen	0.037 nm
2.	Carbon	0.077 nm
3.	Nitrogen	0.074 nm
4.	Oxygen	0.073 nm
5.	Chloride	0.099 nm
6.	Sulphate	0.104 nm

Atoms are so small that they can not be seen even under the most powerful optical microscope. However it is only recently that a highly sophisticated microscope known as scanning tunneling microscope (STM), has magnified image of surface of elements showing atoms.

Relative size	
Radii (in meter)	Example
10^{-10}	molecule of water

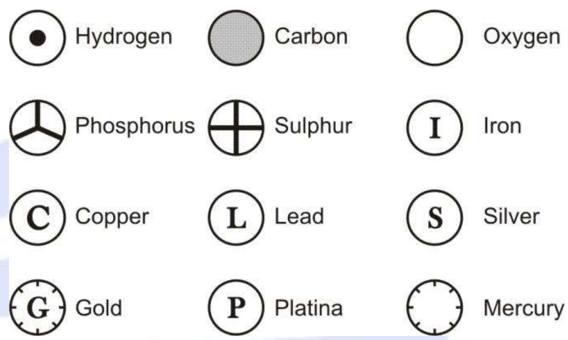
10^{-9}	atom of hydrogen
10^{-8}	molecule of Haemoglobin
10^{-4}	Grain of sand
10^{-3}	ant

4.2 WHAT ARE THE MODERN DAY SYMBOLS OF ATOMS OF DIFFERENT ELEMENTS ?

Symbol is a short method of representing anything. In case of elements a short method of representing the full name of an element is known as **symbol**.

4.2.1 Dalton's symbols of element

Dalton was the first scientist to suggest the symbols for elements in a very specific way. Dalton's symbol for an element represent the "element' as well as one atom of that element. Thus, we can say that the symbol used by him also represent the quantity of the element. A few of these symbols as proposed by Dalton are as follows :



Dalton's symbols were difficult to draw and inconvenient to use so they could not become popular.

4.2.2 Berzelius suggestion for symbols of elements

J.J. Berzelius, a Swedish chemist, suggested a more scientific method for representing an element. He suggested that the first one or two letter of the name of an element can be used as its symbols. This idea led to the development of modern symbols of elements.

4.2.3 Modern symbols of elements

In the beginning, the names of elements were derived from the name of the place where they were first found, discovered or after the name of the scientist who discovered it. For example the name copper was taken from Cyprus, Rutherfordrium after Rutherford etc. Some names were taken from specific colours. For example gold was taken from the English word meaning yellow. However as more and more elements were discovered an international committee was set up called **'International Union of Pure and Applied Chemistry' (IUPAC)** which approved the names of the different elements.

Though the names of most of the elements have been taken from English there are some elements which have been taken from Latin, German or Greek. In all cases the symbol of an element is the "first letter" or "the first letter and another letter" of the English or Latin name of the element. For example :

The symbol of Hydrogen is H

The symbol of oxygen is O

So, in the case of hydrogen and oxygen the first letter of their English names are taken as their symbols.

It should be noted that in a "two letter" symbol, the first letter is the capital letter but the second letter is the small letter. The necessity of adding another letter arises only in case of elements whose names start with the same letter. For example the name of the elements viz. carbon, chlorine calcium and copper starts with the common letter C.
Hence,
Carbon is represented by the symbol C
Chlorine is represented by the symbol Cl
Calcium is represented by the symbol Ca
Copper is represented by the symbol Cu
Symbols of elements derived from English names

English name of the element	Symbol	English name of the element	Symbol
1. Argon	Ar	15. Iodine	I
2. Arsenic	As	16. Lithium	Li
3. Aluminium	Al	17. Magnesium	Mg
4. Boron	B	18. Manganese	Mn
5. Barium	Ba	19. Nitrogen	N
6. Bromine	Br	20. Neon	Ne
7. Carbon	C	21. Nickel	Ni
8. Calcium	Ca	22. Oxygen	O
9. Chlorine	Cl	23. Phosphorus	P
10. Chromium	Cr	24. Platinum	Pt
11. Cobalt	Co	25. Sulphur	S
12. Fluorine	F	26. Uranium	U
13. Hydrogen	H	27. Zinc	Zn
14. Helium	He		

Symbols derived from Latin names

English name of the element	Latin name of the element	Symbol
1. Antimony	Stibium	Sb
2. Copper	Cuprum	Cu
3. Gold	Aurum	Au
4. Iron	Ferrum	Fe
5. Lead	Plumbum	Pb
6. Mercury	Hydragyrum	Hg
7. Potassium	Kalium	K
8. Silver	Argentum	Ag
9. Sodium	Natrium	Na
10. Tin	Stannum	Sn

4.3 ATOMIC MASS

The most remarkable concept that Dalton's atomic theory proposed was that of the atomic mass. According to Dalton, each element had a characteristic atomic mass. We know that atoms are extremely small particles. Hence determining the mass of an individual atom was a relatively difficult task. For example, actual mass of an atom of hydrogen is found to be $1.673 \cdot 10^{\square 24}$ g which is extremely small. However, it was found easy to compare the masses of atoms of different elements with some reference atom. The masses thus obtained are called **relative atomic mass**.

The reference which was chosen earlier was hydrogen atom as it is the lightest element. Its mass was taken as 1. However, by using hydrogen as the reference element the masses of atoms of other elements came out to be fractional. Due to this the reference was changed to oxygen (taken as 16) or in other words 1/16th of the mass of an atom of naturally occurring oxygen was taken as one unit. This was considered relevent due to the following two reasons :

- Oxygen reacted with a large number of elements and formed compounds.
- This atomic mass unit gave masses of most of the elements as whole numbers.

However, in 1961, a universally accepted atomic mass unit carbon-12 isotope was chosen as the standard reference for measuring atomic masses. It is called carbon twelve (C – 12) and is represented as ^{12}C. The atomic mass can be defined as :

"One atomic mass unit is a mass unit equal to exactly one twelfth (1/12th) the mass of one atom of carbon-12. The relative atomic masses of all elements have been found with respect to an atom of carbon -12".

Thus, 1/12th of the mass of an atom of carbon -12 isotope represents one unit of mass on the atomic scale. This is called one atomic mass unit (amu). Now it is represented simply by 'u' which stands for unified mass.

Atomic mass of an element may therefore also be defined as **"the number of times an atom of that element is heavier than 1/12th of the mass of an atom of C-12 isotope".**

For example, an atom of magnesium is found to be two times heavier than an atom of C-12 i.e. 24 times heavier than 1/12th of the mass of C-12 atom. Hence, atomic mass of magnesium = 24 amu.

Atomic masses of some common elements

Element	Symbol	Atomic mass	Element	Symbol	Atomic mass
1. Hydrogen	H	1	14. Sulphur	S	32
2. Helium	He	4	15. Chlorine	Cl	35.5
3. Lithium	Li	7	16. Argon	Ar	40
4. Boron	B	11	17. Potassium	K	39
5. Carbon	C	12	18. Calcium	Ca	40
6. Nitrogen	N	14	19. Iron	Fe	56
7. Oxygen	O	16	20. Copper	Cu	63.5
8. Fluorine	F	19	21. Zinc	Zn	65
9. Neon	Ne	20	22. Silver	Ag	108
10. Sodium	Na	23	23. Platinum	Pt	195
11. Magnesium	Mg	24	24. Gold	Au	197
12. Aluminium	Al	27	25. Lead	Pb	207
13. Phosphorus	P	31	26. Uranium	U	238

5. HOW DO ATOMS EXIST?

Atoms of most elements are not able to exist independently. Only the atoms of a few elements exist in free state.

Atoms usually exist in two ways :

(i) In the form of molecules

(ii) In the form of ions

When they form molecules or ions they become stable. Molecules and ions aggregate in large numbers to form matter which we see around us.

Though we can not see individual atoms, molecules or ions but we can see the matter. For example, we cannot see the Na^+ and Cl^- ions but we can see the sodium chloride compound (common salt).

6. MOLECULE

A molecule is in general a group of two or more atoms that are chemically bonded together i.e. tightly held together by attractive forces.

Or

A molecule is the smallest particle of an element or a compound which can exist freely and possesses all the properties of that substance.

Atoms of the same element or of different elements can join together to form molecules.

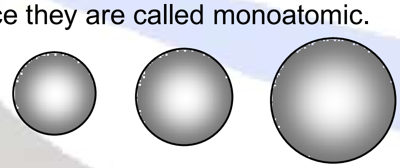
6.1 MOLECULE OF AN ELEMENT

The molecules of an element are constituted by the same type of atoms.

Or

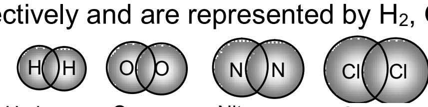
The molecules of an element contain two (or more) similar atoms chemically combined together. Molecules of many elements such as Argon (Ar), Helium (He), etc. are made up of only one atom of that element. But this is not the case with most of the elements. Depending upon whether the molecule contains one, two, three or four atoms they are called monoatomic, diatomic triatomic, tetra atomic or polyatomic. A few examples of molecules of different types are as follows:

(i) Monoatomic molecules : Noble gases like Helium, Neon, etc. exist as single atoms i.e. He, Ne etc. Hence they are called monoatomic.



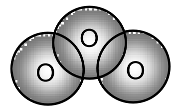
Helium (He) Neon (Ne) Argon (Ar)

(ii) Diatomic molecules : Molecules of Hydrogen, Oxygen, Nitrogen contain two atoms of each element respectively and are represented by H_2, O_2, N_2 etc.



Hydrogen molecule (H_2) Oxygen molecule (O_2) Nitrogen molecule (N_2) Chlorine molecule (Cl_2)

(iii) Triatomic molecules : Molecules containing 3 atoms are called triatomic molecules. For example, ozone contains 3 atoms of oxygen element combined together.



(O_3)

(iv) Tetratomic molecules : Molecules containing 4 atoms of an element are called tetratomic molecules. Most common example is that of phosphorus represented by P_4.

(v) Polyatomic molecules : Molecules containing more than four atoms of particular element are called polyatomic molecules. For example, a molecule of sulphur contains 8 atoms of sulphur and is represented by S_8.

6.1.1 Atomicity

The number of atoms present in one molecule of a substance is known as its **atomicity**.

Let us look at the atomicity of some non-metals.

Atomicity of some elements		
Type of Element	Name	Atomicity
Non-metal	Argon Helium Oxygen Hydrogen Nitrogen Chlorine Phosphorus Sulphur	Monoatomic Monoatomic Diatomic Diatomic Diatomic Diatomic Tetra-atomic Poly-atomic

6.2 Molecules of compounds

The molecules of a compound consist of two or more atoms of different elements combined together in a definite proportion by mass to form a compound that can exist freely.

For example, carbon dioxide contains one atom of carbon and two atoms of oxygen combined together in a fixed ratio of 3 : 8 by mass.

Molecules of some compounds		
Compound	Combining Elements	Ratio by Mass
Water	Hydrogen, Oxygen	1:8
Ammonia	Nitrogen, Hydrogen	14:3
Carbon dioxide	Carbon, Oxygen	3:8

In the above table we have given the ratio by mass of atoms present in one molecule of that particular compound. The atomic masses of different elements are H = 1.0u, O = 16.0u, N = 14.0u, C = 12.0u. By comparing the data we can find out the ratio by number of atoms of elements in the molecule of the particular compound as follows :

S.No.	Compound	Element	ratio by mass	Atomic mass(u)	Mass ratio/ atomic mass	simplest ratio
1.	H_2O	H	1	1	$\frac{1}{1}=1$	2
		O	8	16	$\frac{8}{16}=\frac{1}{2}$	1

2.	NH_3	N	14	14	$\frac{14}{14}=1$	1
		H	3	1	$\frac{3}{1}=3$	3
3.	CO_2	C	3	12	$\frac{3}{12}=\frac{1}{4}$	1
		O	8	16	$\frac{8}{16}=\frac{1}{2}$	2

Thus ratio by number of atoms for water is H:O = 2:1, for ammonia is N:H = 1:3 and for carbon dioxide C:O = 1:2. Thus we can say that in a compound, the element are combined together in a simple whole number atomic ratio.

7. AN ION

The molecular compounds composed of metals and non-metals contain charged species. The charged species are known as **ions**. Ions can be defined as **a positively or negatively charged atom (or group of atoms)**. Depending upon the charge they carry, ions can be of two types :

(i) Cation : A positively charged ion is known as the **cation**. For example, Sodium ion (Na^+), Magnesium ion (Mg^{++}) etc. A cation is formed by the loss of one or more electrons by an atoms. This can be represented as follows :

$$Na \xrightarrow{-1\text{ electron}} Na^+$$

sodium ion (cation)

(ii) Anion : A negatively charged ion is known as **anion**. For example chloride ion (Cl^{-}), oxide ion (O^{--}) are anions as they are negatively charged.

An anion is formed by the gain of one or more electrons.

$$Cl \xrightarrow{+1\text{ electron}} Cl^-$$

Chlorine atom chloride ion (anion)

It is important to know that in any compound the total positive charge of cation is equal to the negative charge of anion, so that the ionic compound as a whole is electrically neutral.

7.1 SIMPLE IONS AND COMPOUND IONS (POLYATOMIC IONS)

7.1.1 Simple ions

Those ions which are formed from single atoms are called simple ions.

For example, Na^+, Mg^{2+}, Al^{3+}, etc.

7.1.2 Polyatomic ion

Ions formed from a group of atoms carrying a charge (either negative or positive) is known as a polyatomic ion or compound ion.

For example NH_4^+, CO_3^{2-}, SO_4^{2-}.

Some of the common ionic compound their formula and the ions present in them are given below :

Some Ionic compounds

Name	Formula	Ions present
1. sodium chloride	$NaCl$	Na^+ and Cl^-
2. Potassium chloride	KCl	K^+ and Cl^-
3. Ammonium chloride	NH_4Cl	NH_4^+ and Cl^-
4. Magnesium	$MgCl_2$	Mg^{2+} and Cl^-
5. Calcium chloride	$CaCl_2$	Ca^{2+} and Cl^-
6. Magnesium oxide	MgO	Mg^{2+} and O^{2-}
7. Calcium oxide	CaO	Ca^{2+} and O^{2-}
8. Aluminium oxide	Al_2O_3	Al^{3+} and O^{2-}
9. Sodium hydroxide	$NaOH$	Na^+ and OH^-
10. Copper sulphate	$CuSO_4$	Cu^{2+} and SO_4^{-2}
11. Calcium nitrate	$Ca(NO_2)_2$	Ca^{2+} and NO_3^-

As discussed earlier a compound is always made up of the same elements combined together in a fixed proportion by mass. For example Hydrogen chloride contains hydrogen and chlorine in a fixed ratio of 2 : 35.5 by mass. Similarly calcium oxide contains calcium and oxygen in the fixed ratio of 40 : 16 or 5 : 2 by mass.

Some ionic compounds		
Ionic Compound	Constituting Elements	Ratio by Mass
Calcium oxide	Calcium and oxygen	5:2
Magnesium Sulphide	Magnesium and sulphur	3:4
Sodium chloride	Sodium and chlorine	23:35:5

8. WRITING A CHEMICAL FORMULA

Before we learn how to write a chemical formula we should know what is a chemical formula? A chemical formula may be defined as the "Symbolic representation of its composition" or in other words.

"A chemical formula of a molecular compound represents the actual number of atoms present in one molecule of the compound".

For example: H_2O is the chemical formula of water, NH_3 is the chemical formula of ammonia. In other words, the formula of a compound tells us the "kind of atoms" as well as "the number of atoms" of various elements present in one molecule of the compound.

"Chemical formula of an ionic compound represents the cations and anions present in the structure of the compound".

For example: Na^+Cl^- represents that sodium chloride contains Na^+ and Cl^- ions in the ratio of 1:1.

To understand how to write the chemical formula of different compounds, we need to learn the symbols and combining capacity of the elements.

8.1 CONCEPT OF VALENCY

Valency can be defined as the combining capacity of that particular element.

To find out how the atoms of an element will combine with the atom (s) of another element to form a compound, valency can be used.

For example valency of oxygen is 2, this means that one atom of oxygen can combine with 2 atoms of hydrogen or in other words we can say that valency of hydrogen is one so 2 atoms of hydrogen can combine with one atom of oxygen to form water molecule.

The concept of valency can be understood more easily by the following examples. The valency of the atom of an element can be considered as hands of that atom.

We know that human beings have two hands and an octopus has eight. If one octopus has to catch hold of a few people in such a way that both arms of all the human and all the eight arms of octopus are locked. How many humans according to you an octopus can hold?

If we represent the octopus with O and humans with H than we can write a formula for this combination as OH_4 where 4 is the subscript which indicates the number of humans held by the octopus.

As molecular compounds are formed by the combination of non-metal atoms, the valencies of some of them are as follows:

Valencies of some common non-metal elements					
Element	Symbol	Valency	Element	Symbol	Valency
Hydrogen	H	1	Oxygen	O	2
Fluorine	F	1	Sulphur	S	2, 4, 6
Chlorine	Cl	1	Nitrogen	N	3, 5
Bromine	Br	1	Phosphorus	P	3, 5
Iodine	I	1	Carbon	C	4

For writing the chemical formula of an ionic compound, valency of an ion can be defined as – "**the units of positive or negative charge present on the ion.**"

For example: Na^+ ion has one unit positive charge

Ca^{2+} ion has two unit positive charge

Cl^- ion has one unit negative charge

SO_4^{2-} ion has two unit negative charge

Depending upon whether the ions has 1, 2, 3 or 4 unit charge (positive or negative), they are called monovalent, divalent, trivalent and tetravalent ions respectively.

Below is given a table showing names and symbols of some ions.

Names and symbols of some ions						
Valency	Name of ion	Symbol	Non-metallic element	Symbol	Polyatomic ions	Symbol
1.	Sodium	Na^+	Hydrogen	H^+	Ammonium	NH^{4+}
	Potassium	K^+	Hydride	H^-	Hydroxide	OH^-
	Silver	Ag^+	Chloride	Cl^-	Nitrate	NO_3^-

	Copper (I)*	Cu^{+}	Bromide	Br^{-}	Hydrogen carbonate	HCO_3^{-}
			Iodide	I^{-}		
2.	Magnesium	Mg^{2+}	Oxide	O^{2-}	Carbonate	CO_3^{2-}
	Calcium	Ca^{2+}	Sulphide	S^{2-}	Sulphite	SO_3^{2-}
	Zinc	Zn^{2+}			Sulphate	SO_4^{2-}
	Iron (II)*	Fe^{2+}				
	Copper (II)*	Cu^{2+}				
3.	Aluminium	Al^{3+}	Nitride	N^{3-}	Phosphate	PO_4^{3-}
	Iron (III)*	Fe^{3+}				

From the above table, it is seen that some elements show a number of valencies (called variable valencies). For example, copper shows a valency of 1 as well as 2, iron shows a valency of 2 as well as 3. These valencies are sometimes shown by Roman numeral in a bracket after the symbol of the element, e.g., Cu (I), Cu (II), Fe (II), Fe (III) etc. Some non-metal atoms also show variable valency.

8.2 RULES FOR WRITING THE CHEMICAL FORMULAE

While writing a chemical formulae of molecular or ionic compounds the following steps are to be followed:

(i) In case of simple molecular compounds (compounds made up of only two elements). The symbols of the two elements are written side by side and their respective valencies are written below their symbols.

(ii) In case of simple ionic compounds, the symbol of the cation or metal atom is written first followed by the symbol of the anion or non-metal atom and their respective valencies are written below their symbols. For example in CaO, symbol of calcium (Ca, a metal) must be written first followed by symbol of oxygen (which is a non-metal).

(iii) The valencies or charges on the ion must be balanced.

(iv) In case of compounds containing polyatomic ions, the formula of the polyatomic ion is written in brackets and the valencies are written below.

(v) In any of the above cases, if there is a common factor between the valencies of the cation and anion, the valencies are divided by the common factor.

(vi) Finally we apply cross-over of the valencies so that they appear on the lower right hand side of the symbols. However, 1 appearing on the lower right hand side of the symbol is omitted. Similarly we also omit the + and − signs of the charges of the ions.

8.2.1 Formulae of Simple Compounds

The simplest compounds, which are made up of two different elements are called binary compounds.

By applying the above mentioned rules, the chemical formulae of some simple molecular compounds can be written as follow:

Example 1: Steps for writing the formula of Hydrogen chloride:

(i) Elements present are: hydrogen and chlorine
(ii) Symbols of the elements: H Cl
(iii) Valency of the elements: 1 1
(iv) Cross-over of the valency:

H Cl
1 1

(v) We stated that one appearing on the lower right hand side of the symbol is omitted. So the formula of the compound would be HCl.

Example 2: Steps for writing the formula of Methane

(i) Elements present are: Carbon Hydrogen
(ii) Symbols of the elements: C H
(iii) Valency of the elements: 4 1
(iv) Cross-over of the valency:

C H
4 1

(v) So the chemical formula of the methane can be written as CH_4 (As we omit one appearing on the lower right hand side of carbon element).

Example 3: Steps for writing the chemical formula of Carbon dioxide

(i) Elements present are: Carbon Oxygen
(ii) Symbols of the elements: C O
(iii) Valency of the elements: 4 2
(iv) Dividing by common factor: $\frac{4}{2} = 2$ $\frac{2}{2} = 1$
(v) Cross-over of the valency:

C O
2 1

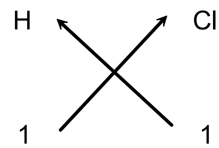
(vi) So the chemical formula of carbon dioxide can be written as CO_2 (as we omit one appearing on the lower right hand side of carbon element).

8.2.2 Chemical formulae of some simple ionic compound

Example 1: Steps for writing the chemical formula of Sodium chloride

(i) Elements present in the compound: Sodium Chlorine
(ii) Symbols of the elements: Na Cl
(iii) Charge on the ions +1 −1
(iv) Valency of the elements: 1 1

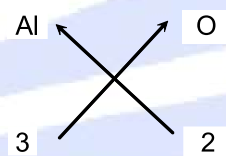
(v) Cross-over of the valency:



(vi) So the chemical formula of sodium chloride can be written as NaCl (As we omit 1 appearing on the lower right hand side of Na and Cl atoms and + and – sign of the charges of ions).

Example 2: Steps for writing the chemical formula of Aluminium oxide

(i)	Elements present in the compound:	Aluminium	Oxygen
(ii)	Symbols of the elements:	Al	O
(iii)	Charge on the ions	+3	−2
(iv)	Valency of the elements:	3	2
(v)	Cross-over of the valency:		



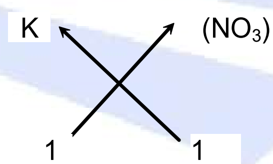
(vi) So the chemical formula of aluminium oxide can be written as Al_2O_3.

8.2.3 Chemical formulae of compounds containing polyatomic ions

While writing the chemical formulae of compounds containing polyatomic ions, same rules will apply except that the formula of polyatomic ion is written in brackets.

Example 1: Steps for writing the formula of potassium nitrate

(i)	Symbols of the ions:	K	(NO_3)
(ii)	Charge on the ions:	1+	1−
(iii)	Valency of the ions:	1	1
(iv)	Cross-over of the valency:		



(v) So the chemical formula of Potassium nitrate can be written as KNO_3 (As we omit one appearing on the right hand side of the ions).

Example 2: Steps for writing the chemical formula of Calcium nitrate

(i)	Symbols of the ions:	Ca	(NO_3)
(ii)	Charge on the ions:	2+	1−
(iii)	Valency of the ions:	2	1
(iv)	Cross-over of the valency:		

Ca $\quad$ (NO_3)

2+ $\quad$ 1–

(v) So the chemical formula of Calcium nitrate can be written as $Ca(NO_3)_2$ (As we omit one appearing on the right hand side of the Ca ions).

Example 3: Steps for writing the chemical formula of Ammonium sulphate

(i)	Symbols of the ions present:	(NH_4)	(SO_4)
(ii)	Charge on the ions:	1+	2□
(iii)	Valency of the ions:	1	2
(iv)	Cross-over of the valency:		

(NH_4) $\quad$ (SO_4)

1+ $\quad$ 2–

(v) So we can write the chemical formula of ammonium sulphate as $(NH_4)_2SO_4$ (As we omit one appearing on the right hand side of sulphate ion).

8.2.4 Some general examples

Example 1: An element E is trivalent. Write the chemical formula of its (i) chloride (ii) oxide (iii) sulphide.

Solution: (i) Steps for writing the chemical formula of Chloride of trivalent element E.

(i)	Elements present:	Element	Chloride
(ii)	Symbols:	E	Cl
(iii)	Valency:	3	1
(iv)	Cross-over of valency:		

E $\quad$ Cl

3 $\quad$ 1

(v) Thus, we can write the chemical formula as ECl_3.

(ii) Steps for writing the chemical formula of Oxide of trivalent element E.

(i)	Elements present:	Element	Oxygen
(ii)	Symbols:	E	O
(iii)	Valency:	3	2
(iv)	Cross-over of valency:		

E $\quad$ O

3 $\quad$ 2

(v) Thus we can write the chemical formula as E_2O_3.

(iii) Steps for writing the chemical formula of Sulphide of trivalent element E.

		Element	Sulphur
(i)	Elements present:	Element	Sulphur
(ii)	Symbols:	E	S
(iii)	Valency:	3	2

(iv) Cross-over of valency:

E S

3 2

(v) Thus we can write the chemical formula as E_2S_3

Example 2: An element M forms the oxide M_2O_3. What will be the formula of its phosphate?

Solution: (i) In M_2O_3 total charge on 2 oxide ion = 3 x (−2) = −6

(ii) As the compound M_2O_3 is neutral so the charge on two metal atoms should be = +6

4 Charge on one metal atom = + 6/2 = + 3

i.e. valency of metal atom = 3

Steps for writing the formula of metal phosphate.

(i)	Symbols of the elements/ion present:	M	(PO_4)
(ii)	Charge:	3+	3−
(iii)	Valency:	3	3
(iv)	Dividing by the common factor	$\frac{3}{3} = 1$	$\frac{3}{3} = 1$

(iv) Cross-over

M PO_4

1 1

(v) Thus we can write the chemical formula as MPO_4

9. MOLECULAR MASS AND MOLE - CONCEPT

The molecular mass of a substance is the sum of the atomic masses of all the atoms in a molecule of the substance. It is therefore the relative mass of a molecule expressed in atomic mass unit.

Molecular mass of a substance can be defined as follows:

"Molecular mass of a substance (element or compound) is the average relative mass of its molecule as compared with that of an atom of C-12 isotope taken as 12".

OR

"Molecular mass of a substance represents the number of times the molecules of that substance is heavier than $1/12^{th}$ of the mass of an atom of C-12 isotope".

For example the molecular mass of hydrogen is 2, which means that a molecule of hydrogen is two times heavier than the $1/12^{th}$ of the mass of an atom of C-12 isotope.

9.1 CALCULATION OF MOLECULAR MASS

As we know that molecules are made up of two or more atoms of same or different elements and each element/atom has a definite atomic mass, therefore, molecular mass of a molecule of a substance can be calculated by adding atomic masses of all the atoms present in one molecule of the substance. For example,

(a) A molecule of water has the formula H_2O. Hence molecular mass of
H_2O = (2 x atomic mass of hydrogen) + (1 x atomic mass of oxygen)
= (2 x 1.0 u + 1 x 16.04) – 18 u

(b) A molecule of sulphuric acid has the formula H_2SO_4. Hence molecular mass of
H_2SO_4 = (2 x atomic mass of hydrogen) + (1 x atomic mass of sulphur)
+ (4 x atomic mass of oxygen)
= (2 x 1.0 u) + (1 x 32.0 u) + (4 x 16.0 u)
= (2.0 u + 32.0 u + 64.0 u)
= 98.0 u

9.2 FORMULA UNIT MASS

Before describing the formula of unit mass, we should be aware of the meaning of 'Formula Unit' of an ionic compound.

An ionic compound consists of a very large but equal number of positively charged and negative charged ions (as an ionic compound is electrically neutral) arranged in a definite order in a crystal lattice. Thus the actual formula of an ionic compound should be written as $(C^{+})_n$ $(a)_n$ or $(C^{+}a^{-})_n$, where 'n' is a very large number. Thus formula unit can be defined as follows:

The simplest combination of ions that produces an electrically neutral unit, is called a 'formula unit' of the ionic compound. It is thought to be the smallest unit of that compound. For example the formula unit of sodium chloride is NaCl.

As we have discussed that ionic compounds do not exist in molecular form, so we can not give its molecular mass. In such case, the term "Formula unit mass" is used.

The formula unit mass of an ionic compound can be defined as follows:

"The formula unit mass or formula mass of an ionic compound is the sum of the atomic masses of all the atoms present in one formula unit of the compound".

9.3 GRAM ATOMIC MASS AND GRAM MOLECULAR MASS

In order to understand the "Mole concept" we should first know the meaning of the terms "gram atomic mass" and "gram molecular mass".

9.3.1 Gram Atomic Mass

"**Atomic mass expressed in grams is called gram atomic mass of that element.**"

Or

"**The amount of a substance whose mass in grams is numerically equal to its atomic mass, is called the gram atomic mass of that substance.**"

In other words, we can say that atomic mass and gram atomic mass of a substance is numerically equal, but their units are different.

For example:

(a) Atomic mass of Na = 23.0u

Gram atomic mass of Na = 23.0g

(b) Atomic mass of Cl = 35.5u

Gram atomic mass of Cl = 35.5g and so on.

9.3.2 Gram Molecular Mass

Like gram atomic mass, gram molecular mass can also be defined as follows:

"Molecular mass expressed in grams is called gram molecular mass of that element"

Or

"The amount of substance whose mass in grams is numerically equal to its molecular mass is called gram molecular mass of that substance".

In other words, molecular mass and gram molecular mass of a substance is numerically equal, only their units are changed.

For example:

(a) Molecular mass of H_2 = 2.0u

Gram molecular mass of H_2 = 2.0g

(b) Molecular mass of H_2O = 18.0u

Gram molecular mass of H_2O = 18.0g.

Thus to convert molecular mass or atomic mass into gram molecular mass or gram atomic mass, we have to replace 'u' by 'g'.

9.3.3 Gram Formula Unit Mass

Similarly we can define the gram formula unit mass as follows:

"Formula unit mass expressed in grams is called as gram formula unit mass".

For example:

Formula unit mass of NaCl = 23.0u + 35.5u

= 58.5u

Gram formula unit mass of Nacl = 58.5u

9.4 MOLE CONCEPT

In our daily life we buy some products in terms of numbers such as dozens (for 12), or gross (for 144) or in terms of mass such as kilograms (1000g) or quintals (100kg).

For example: We buy bananas, oranges, mangoes etc. in dozens where as rice, wheat, pulses etc. in kilograms.

1 dozen banana = 12 banana

1 dozen mango = 12 mangoes

1 kilogram wheat = 1000g wheat

1 quintal rice = 100 kilogram rice and so on.

We can not buy rice, pulses, sugar etc. in terms of numbers as they are very small in size. Similarly atoms and molecules are also very small in size. They are so small that we can not see them through our eyes.

So to express a definite amount of a chemical substance a new bigger unit "mole" was introduced. In Latin '**mole**' means 'heap' or 'collection' or 'pile'. The mole may be expressed in terms of mass or in terms of number.

9.4.1 Mole in Terms of Mass

"A mole of atoms is defined as" the amount of the substance i.e. element which has mass equal to its gram atomic mass".

Or

"A mole of atoms is equal to one gram atom of that particular element".

For example:

1 mole of Hydrogen (H) atom = 1g atom of H = 1.0g
1 mole of Oxygen (O) atom = 1g atom of O = 16.0g
1 mole of Nitrogen (N) atom = 1g atom of N = 14.0g
1 mole of Sodium (Na) atom = 1g atom of Na = 23.0g

similarly

"A mole of molecule is defined as that amount of the substance (element or compound) which has mass equal to gram molecular mass".

Or

"A mole of molecule is equal to one gram molecule of the substance".

For example,

1 mole of Hydrogen (H_2) molecule = 1 g molecule of H_2 = 2.0g
1 mole of Oxygen (O_2) molecule = 1 g molecule of O_2 = 32.0g
1 mole of Ammonia (NH_3) molecule = 1 g molecule of NH_3 = 17.0g
1 mole of Carbon dioxide molecule = 1 g molecule of CO_2 = 44.0g

9.4.2 Mole in Terms of Number

As we have discussed earlier that one dozen contain 12 number of a particular substance or one gross contain 144 number of the substance in consideration. Similarly the number of particles (atoms, molecules or ions) present in 1 mole of any substance is fixed. The concept of mole in terms of number was arrived as follows:

Consider any two chemical substance, suppose carbon and oxygen.

Atomic mass of carbon = 12.0u
Molecular mass of oxygen (O_2) = 32.0u

Thus the ratio of mass of 1 atom of C to 1 molecule of oxygen (O_2) = 12:32. Now if we take a million atoms of carbon and a million molecules of oxygen (O_2), the ratio of their masses will remain the same. Conversely, it implies that the masses of carbon and oxygen taken in the ratio of 12:32 will contain the same number of atoms or oxygen molecules. Thus we can say that if we take 32.0g of oxygen (i.e. 1 mole of oxygen), it will contain the same number of oxygen molecules as the number of atoms in 12g of carbon (i.e. 1mole of carbon atoms).

Thus is general mole can be defined as follows:

"A mole of particles (atoms, molecules or ions) is defined as that amount of the substance which contains the same number of particles as there are C-12 atoms in 12g of carbon".

Experimentally, it has been found that 12g of C-12 isotope contain 6.022×12^{23} atoms. This number is called **Avogadro's number or Avogadro's constant** and is represented by the symbol N_0. Thus

Avogadro's number (N_0) = 6.022×10^{23}.

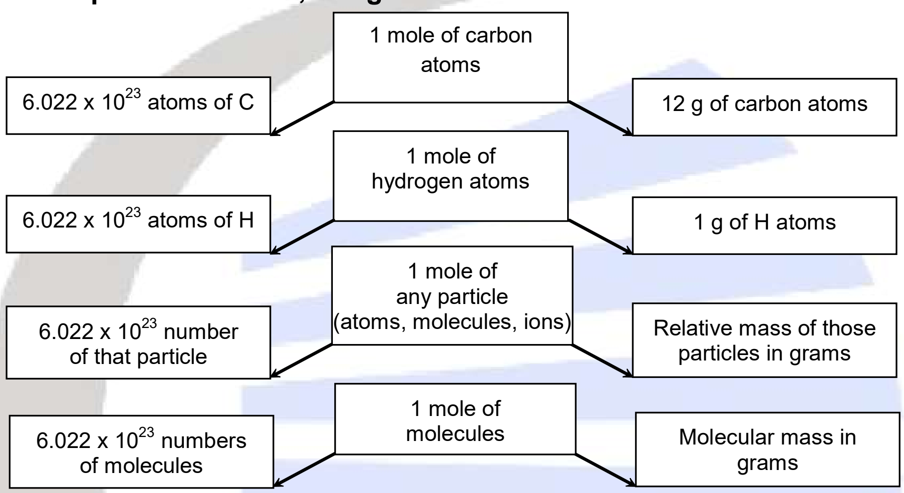
Thus a mole of particles can also be defined as follows:

"A mole of particles (atoms, molecules or ions) is that amount of the substance which contain 6.022 x 10^{23} particles".

For example:

1 mole of C atoms	= 6.022 x 10^{23} C atoms
1 mole of O atoms	= 6.022 x 10^{23} O atoms
1 mole of H atoms	= 6.022 x 10^{23} H atoms
1 mole of H_2O molecules	= 6.022 x 10^{23} H_2O molecules
1 mole of CO_2 molecules	= 6.022 x 10^{23} CO_2 molecules
1 mole of Na^+ ions	= 6.022 x 10^{23} Na^+ ions

9.4.3 Relationship Between Mole, Avogadro Number and Mass



9.4.4 Significance of the mole

A mole represents the following:

(i) It represents 6.022 x 10^{23} particles of the substance

(ii) The mass of one mole of an element is equal to the mass of 6.022 x 10^{23} atoms of that element.

(iii) One mole of a substance represents one gram, formula mass of that substance.

9.4.5 Formulae for calculation

(a) Number of moles = $\dfrac{\text{Mass of the element in grams}}{\text{Gram atomic mass of the element}}$

i.e. $$n = \frac{m}{M}$$

(b) Number of moles = $\dfrac{\text{Given number of atom or molecules}}{\text{Avogadro's number}}$

$$n = \frac{N}{N_0}$$