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Class 9

with CASE STUDY Approach for IIT-JEE/ NEET / Olympiad

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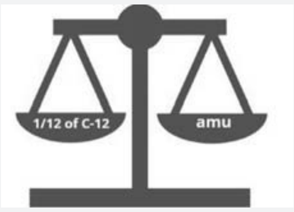


LAWS OF CHEMICAL COMBINATION

One of the most important aspects of the subject of chemistry is the study of chemical reactions. These chemical reactions take place according to certain laws, called the 'laws of chemical combination'.

(A) Law of conservation of mass
(B) Law of constant proportions
(C) Law of multiple proportions
(D) Law of reciprocal proportions
(E) Law of combining volumes (Gas Lussac's Law of Gaseous volumes)

The laws of chemical combination are the experimental laws which led to the idea of atoms being the smallest unit of matter. The laws of chemical combination played a significant role in the development of Dalton's atomic theory of matter.

The first four laws deal with the mass relationships whereas the fifth law deals with the volumes of the reacting gases.

(A) Law of Conservation of Mass or Matter :

This law was given by Antoine Lavoisier. The law of conservation of mass means that in a chemical reaction, the total mass of products is equal to the total mass of the reactants. There is no change in mass during a chemical reaction. Suppose we carry out a chemical reaction between A and B to form C and D, then.

$$A + B \longrightarrow C + D$$

$$\text{Mass of } (A + B) = \text{Mass of } (C + D)$$

or mass of reactants = mass of products.

Example :

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

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

The flask is again weighed. It is found that there is no change in the weight though a physical change has taken place.

(ii) When we burn a sample of metallic magnesium in air, the magnesium combines with oxygen from the air to form magnesium oxide, a white powder. This chemical reaction is accompanied by the release of large amount of heat energy and light energy. When we weigh the product of the reaction, magnesium oxide, we find that it is heavier than the original piece of magnesium. The increase in mass of the solid is due to the combination of oxygen with magnesium to form magnesium oxide. Now when we conduct same experiment in a closed container we find that now the mass of the magnesium oxide is exactly the sum of the masses of magnesium and oxygen that combined to form it. Similar statements can be made for all chemical reactions.

> **DID YOU KNOW?**
> *Thus according to this law, there is no increase or decrease in the total mass of matter during a chemical or a physical change.*

Illustration 1 :

4.90 g of $KClO_3$ when heated produce 1.92 g of oxygen and the residue (KCl) left behind weighs 2.96 g. Show that these results illustrate the law of conservation of mass.

Solution : Mass of $KClO_3$ taken = 4.90 g

Total mass of the products ($KCl + O_2$) = 2.96 + 1.92 = 4.88 g

Difference between the mass of the reactant and the total mass of the products = 4.90 – 4.88 = 0.02 g

This small difference may be due to experimental error.

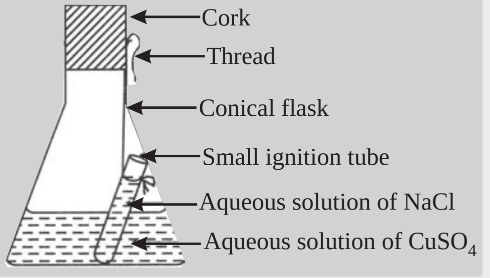
Thus, law of conservation of mass holds good within experimental errors.

Let's Do Activity

From your school chemistry lab collect the following apparatus

- *Ignition tube*
- *Conical flask*
- *Thread*
- *Rubber cork*
- *Aqueous solution of NaCl and $CuSO_4$*

By using above apparatus make an arrangement as shown in figure below.



Now first weigh the apparatus shown in figure and record your observation then tilt the ignition tube containing aqueous NaCl. Swirl the conical flask slowly for few minutes so that solution of both salts get mixed properly. After mixing weigh apparatus again and record your observation. On the basis of your observation draw a conclusion.

(B) Law of Constant (Definite) Proportions

Most elements can interact with other elements to form compounds. Hydrogen gas, for example, burns in oxygen gas to form water. Conversely, water can be decomposed into its component elements by passing an electrical current through it. Pure water, regardless of its source, consists of 11% hydrogen and 89% oxygen by mass. This macroscopic composition corresponds to the molecular composition, which consists of two hydrogen atoms combined with one oxygen atom.

The observation that the element composition of a pure compound is always the same is known as the law of constant composition (or the law of definite proportion). It was first put forth by the French chemist Joseph Louis Proust (1754-1826) in 1799. Although this law has been known for 200 years, the general belief persists among some people that a fundamental difference exists between compounds prepared in the laboratory and the corresponding compound found in nature. However, a pure compound has the same composition and properties regardless of its source. Both chemists and nature must use the same elements and operate under the same natural laws. When two materials differ in composition and properties, we know that they are composed of different compounds or that they differ in purity.

For example, carbon dioxide can be obtained by any of the following chemical reactions.

(i) by burning carbon in oxygen

$$C(s) + O_2(g) \longrightarrow CO_2(g)\uparrow$$

(ii) by heating sodium bicarbonate

$$2NaHCO_3(s) \xrightarrow{\Delta} Na_2CO_3(s) + CO_2(g)\uparrow + H_2O(l)$$

(iii) by heating calcium carbonate

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)\uparrow$$

(iv) by reacting calcium carbonate with hydrochloric acid

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g)\uparrow + H_2O(l)$$

DID YOU KNOW?

When same elements combine in the same proportion it is not necessary that same compound will be formed. Sometimes different compounds under different experimental conditions are formed e.g., combination of carbon, hydrogen and oxygen in the ratio 12 : 3 : 8 may give either ethyl alcohol (C_2H_5OH) or dimethyl ether (CH_3OCH_3) under different experimental conditions.

whatever sample of carbon dioxide is taken. It is found that the ratio of carbon and oxygen is always same which is 12 : 32 or 3 : 8.

Illustration 2 :

Copper oxide was prepared by the following methods:

(a) In one case, 1.75 g of the metal were dissolved in nitric acid and igniting the residual copper nitrate yielded 2.19 g of copper oxide.

(b) In the second case, 1.14 g of metal dissolved in nitric acid were precipitated as copper hydroxide by adding caustic alkali solution. The precipitated copper hydroxide after washing, drying and heating yielded 1.43 g of copper oxide.

(c) In the third case, 1.45 g of copper when strongly heated in a current of air yielded 1.83 g of copper oxide.
Show that the given data illustrate the law of constant composition.

Solution : In the first experiment.

2.19 g of copper oxide contained 1.75 g of Cu.

$\therefore$ 100 g of copper oxide contained = $\frac{1.75}{2.19} \times 100 = 79.91$ g

In the second experiment.

1.43 g of copper oxide contained 1.14 g of copper

$\therefore$ 100 g of copper oxide contained = $\frac{1.14}{1.43} \times 100 = 79.72$ g

In the third experiment.

1.83 g of copper oxide contained 1.46 g of copper

$\therefore$ 100 g of copper oxide contained = $\frac{1.46}{1.83} \times 100 = 79.78$ g

Thus, the percentage of copper in copper oxide derived from all the three experiments is nearly the same.
Hence, the above data illustrate the law of constant composition.

CHECK POINT-1

Is the law of constant composition true for all types of compounds? Explain why or why not.

Sol. No, law of constant composition is not true for all types of compounds. It is true only for the compounds obtained from one isotope. For example, carbon exists in two common isotopes, ^{12}C and ^{14}C. When it forms CO_2 from ^{12}C, the ratio of masses is 12 : 32 = 3 : 8but from ^{14}C, the ratio will be 14 : 32 = 7 : 16 which is not same as in the first case.

DALTON'S ATOMIC THEORY

To describe the structure of matter which could explain the experimental facts known at that time about elements, compounds and mixtures and also the laws of chemical combination, John Dalton in 1808 put forward a theory known as Dalton's atomic theory. The main points of this theory are as follows:

(i) Each element is composed of extremely small particles called atoms.

(ii) All atoms of a given element are identical, the atoms of different elements are different and have different properties (including different masses).

(iii) Atoms are neither created nor destroyed in chemical reactions.

(iv) Compounds are formed when atoms of more than one element combine. a given compound always has the same relative number and kind of atoms.

(v) When atoms combine with one another to form compound, they do so in simple whole number ratios, such as 1 : 1, 2 : 1, 2 : 3 and so on.

(vi) Atoms of two elements may combine in different ratios to form more than one compound. For example, sulphur combines with oxygen to form sulphur dioxide and sulphur trioxide, the combining ratios being 1 : 2 and 1 : 3 respectively.

(vii) An atom is the smallest particle that takes part in a chemical reaction.

Drawbacks of Dalton's Atomic Theory :

Some of the drawbacks of the Dalton's atomic theory of matter are given below :

(i) According to Dalton's atomic theory, atoms were thought to be indivisible. But, it is now known that under circumstances, atoms can be further divided into still smaller particles called electrons, protons and neutrons.

(ii) Dalton's atomic theory said that all the atoms of an element have exactly the same mass. But, it is now known that atoms of the same element can have slightly different masses.

(iii) Dalton's atomic theory said that atoms of different elements have different masses. But, it is now known that even atoms of different elements can have the same mass.

(iv) It could explain the laws of chemical combination by mass but failed to explain the law of gaseous volumes.

(v) Why do atoms of the same or different elements combine at all to form molecules?

(vi) What is the nature of binding force between atoms and molecules which accounts for the existence of matter in three states i.e., solids, liquids and gases?

(vii) It makes no distinction between the ultimate particles of an element or a compound.

CONNECTING TOPIC

In addition to these two laws there are some other laws of chemical combination.

(i) Law of Multiple Proportions :

Whenever two elements form more than one compound, the different masses of one element that combine with the same mass of the other element are in the ratio of small whole numbers. This law was given by Dalton in 1804.

For example, sulfur and oxygen form two different compounds which we call sulfur dioxide and sulfur trioxide. If we decompose a 2.00g sample of sulfur dioxide, we find it contains 1.00g of S and 1.00 g of O. If we decompose a 2.50g sample of sulfur trioxide, we find it also contains 1.00g of S, but this time the mass of O is 1.50g. This is summarized in the following table.

Compound	Sample mass	Mass of sulfur	Mass of oxygen
Sulfur dioxide	2.00 g	1.00 g	1.00 g
Sulfur trioxide	2.50 g	1.00 g	1.50 g

First, notice that sample masses aren't the same, they were chosen so that each has the same mass of sulfur. Second, the ratio of the masses of oxygen in the two samples is one of small whole numbers.

$$\frac{\text{mass of oxygen in sulfur trioxide}}{\text{mass of oxygen in sulfur dioxide}} = \frac{1.50\text{g}}{1.00\text{g}} = \frac{3}{2}$$

DID YOU KNOW?

Similarly, N and O form as many five stable oxides, N_2O, NO, N_2O_3, N_2O_4 and N_2O_5. In these oxides, amount of oxygen which react with 28 g of N_2 are in the ratio 16 : 32 : 48 : 64 : 80 i.e., 1 : 2 : 3 : 4 : 5.

Similar observations are made when we study other elements that form more than one compound with each other, and these observations form the basis of the law of multiple proportions.

Illustration 3 :

The following data were collected for several compounds of nitrogen and oxygen :

Mass of nitrogen that combines with 1g of oxygen

Compound I	**1.750g**
Compound II	**0.8750g**
Compound III	**0.4375g**

Show how these data illustrate the law of multiple proportions.

Solution :

For the law of multiple proportions to hold true, the ratios of the masses of nitrogen combining with 1 gram of oxygen in each pair of compounds should be small whole numbers. We therefore get the ratios as follows :

$$\frac{\text{I}}{\text{II}} = \frac{1.750}{0.875} = \frac{2}{1}$$

$$\frac{\text{II}}{\text{III}} = \frac{0.875}{0.4375} = \frac{2}{1}$$

$$\frac{\text{I}}{\text{III}} = \frac{1.750}{0.4375} = \frac{4}{1}$$

These results support the law of multiple proportions.

(ii) The Law of Reciprocal Proportions or Law of Equivalent Proportions :

This law was given by "Richer" in 1792–94. The law states that the ratio of the masses of two elements A and B which combine separately with a fixed mass of the third element C is either the same or some simple multiple of the ratio of the masses in which A and B combine directly with each other.

Example : $\underset{12}{C} + \underset{32}{O_2} \longrightarrow CO_2$, $\underset{12}{C} + \underset{4}{2H_2} \longrightarrow CH_4$

C combines with O to form CO_2 and with H to form CH_4. In CO_2, 12 gm of C reacts with 32 gm of O, whereas in CH_4, 12 gm of C reacts with 4 gm of H. Therefore when O combines with H, they should combine in the ratio of 32 : 4 (i.e. 8 : 1) or in simple multiple of it. The same ratio is found to be true in H_2O molecule, the ratio of weight of H and O in H_2O is 1 : 8.

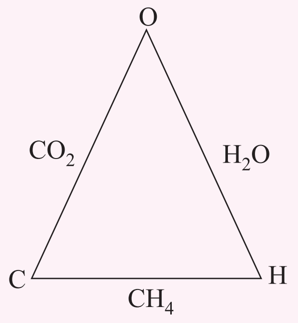
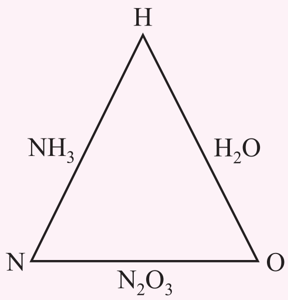


Illustration 4 :

Ammonia contains 82.35% of nitrogen and 17.65% of hydrogen. Water contains 88.90% of oxygen and 11.10% of hydrogen. Nitrogen trioxide contains 63.15% of oxygen and 36.85% of nitrogen. Show that these data illustrate the law of reciprocal proportions.

Solution :



In NH_3, 17.65 g of H combine with N = 82.35 g

$\therefore$ 1 g of H combine with N = $\frac{82.35}{17.65}$ g = 4.67 g

In H_2O, 11.10 g of H combine with O = 88.90 g

$\therefore$ 1 g of H combine with O = $\frac{88.90}{11.10}$ g = 8.01 g

$\therefore$ Ratio of the masses of N and O which combine with fixed mass (=1g) of H = 4.67 : 8.01 = 1 : 1.72

In N_2O_3, ratio of masses of N and O which combine with each other = 36.85 : 63.15 = 1:1.72

Since, the two ratios are the same. Hence, it illustrates the law of reciprocal proportions.

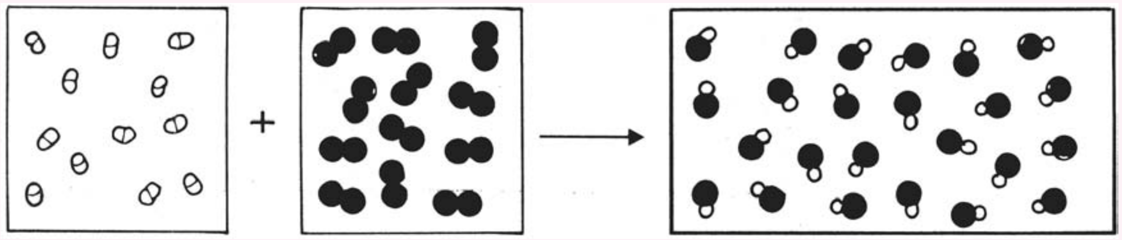
(iii) The Law of Gaseous Volumes :

This law was given by **"Gay Lussac"** in 1808. The law states that "whenever gases react together, the volumes of the reacting gases as well as the products if they are gases, bear a simple whole number ratio, provided all the volumes are measured under similar conditions of temperature & pressure.

Example : $H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$

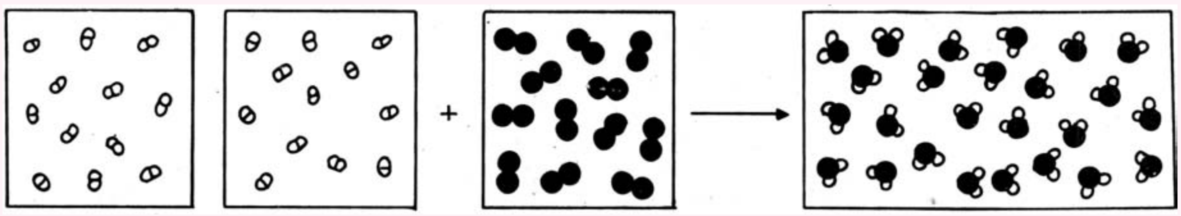
1 unit vol. 1 unit vol. 2 unit vol.

ratio = 1 : 1 : 2



1 volume of hydrogen   1 volume of chlorine   1 volume of hydrogen chloride

Similarly, we observe that 2 volumes of hydrogen combine with 1 volume of oxygen to give 2 volumes of water vapour as



2 volume of hydrogen   1 volume of oxygen   1 volume of water vapour

Thus, the volumes of hydrogen and oxygen which combine to form water is in the ratio : 2 : 1 : 2.

$$\underset{2\,\text{volumes}}{2H_2(g)} + \underset{\text{volumes}}{O_2(g)} \longrightarrow \underset{2\,\text{moles}}{2H_2O(g)}$$

(iv) The Avogadro Law

This law states that "equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules".

Example : $2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$

2 unit vol. 1 unit vol. 2 unit vol.

Ratio of number of molecules = 2 : 1 : 2

Applications of Avogadro's Law

(1) This law helped to remove anomaly between Dalton's atomic theory and Gay Lussac's law of volume by making a clear distinction between atoms and molecules.

(2) It reveals that common elementary gases like hydrogen, nitrogen, oxygen etc. are diatomic.

(3) It provides a method to determine the molecular weights of gaseous elements.

DID YOU KNOW?

It provides a relationship between vapour density and molecular weight of substances.

Vapour density (VD)

$$= \frac{\textit{Mass of same volume of gas at STP}}{\textit{Mass of same volume of } H_2 \textit{ at STP}}$$

or *molecular weight of gas*

= 2 × vapour density

Let's Connect

1. Two samples of lead oxide were separately reduced to metallic lead by heating in a current of hydrogen. The weight of lead from one oxide was half the weight of lead obtained from the other oxide. The data illustrates –

(a) law of reciprocal proportions (b) law of constant proportions

(c) law of multiple proportions (d) law of equivalent proportions

2. In compound A, 1.00g of nitrogen unites with 0.57g of oxygen. In compound B, 2.00g of nitrogen combines with 2.24g of oxygen. In compound C, 3.00g of nitrogen combines with 5.11g of oxygen. These results obey the following law

(a) law of constant proportions (b) law of multiple proportions

(c) law of reciprocal proportions (d) Dalton's law of partial pressure

3. Which one of the following pairs of compounds illustrate the law of multiple proportions ?

(a) H_2O and Na_2O (b) MgO and Na_2O

(c) Na_2O and BaO (d) $SnCl_2$ and $SnCl_4$

4. Among the following pairs of compounds, the one that illustrates the law of multiple proportions is

(a) NH_3 and NCl_3 (b) H_2S and SO_2

(c) CS_2 and $FeSO_4$ (d) CuO and Cu_2O

5. The molecular weight of O_2 and SO_2 are 32 and 64 respectively. At 15°C and 150 mm Hg pressure, one litre of O_2 contains 'N' molecules. The number of molecules in two litres of SO_2 under the same conditions of temperature and pressure will be :

(a) N/2 (b) 1N (c) 2N (d) 4N

Sol. 1. (c) Law of multiple proportions.

2. (b) Law of multiple proportions. As the ratio of oxygen which combines with fix weight of 1 g of nitrogen bears a simple whole number ratio

0.57 : 1.12 : 1.703

1 : 2 : 3

3. (d) $SnCl_2$ $SnCl_4$

$119 : 2 \times 35.5$ $119 : 4 \times 35.5$

Chlorine ratio in the two compounds is $= 2 \times 35.5 : 4 \times 35.5 = 1 : 2$

4. (d) In CuO and Cu_2O the O : Cu is 1 : 1 and 1 : 2 respectively. This is in accordance with the law of multiple proportions.

5. (c) According to Avogadro's law "equal volumes of all gases contain equal number of molecules under similar conditions of temperature and pressure". Thus if 1 L of one gas contains N molecules, 2 L of any other gas under the same conditions of temperature and pressure will contain 2N molecules.

CHECK POINT-2

1 L of a gas at S.T.P. weighs 1.97 g. What is the vapour density of the gas?

Sol. 22.4 L of the gas at S.T.P. will weigh = 1.97 × 22.4 = 44.1 g i.e. Molecular mass = 44.1. Hence vapour density = 44.1 / 2 = 22.05.

ATOM

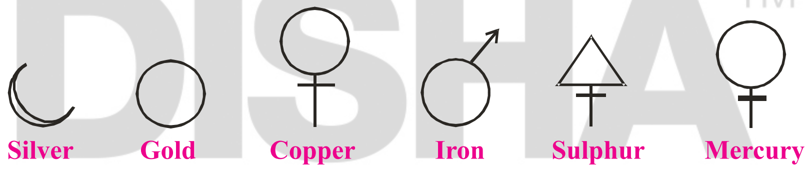
Atoms are the building blocks of all matter. It is the smallest particle of an element that maintains its chemical identity throughout all chemical and physical changes. Atoms of one element are different from those of the other elements. Today, we know that atoms are not truly indivisible, they are themselves made up of particles (protons, neutrons, electrons, etc.). The size of an atom is indicated by its radius which is measured in nanometres (= 10^{-9} m), 1m = 10^9 nm.

DID YOU KNOW?

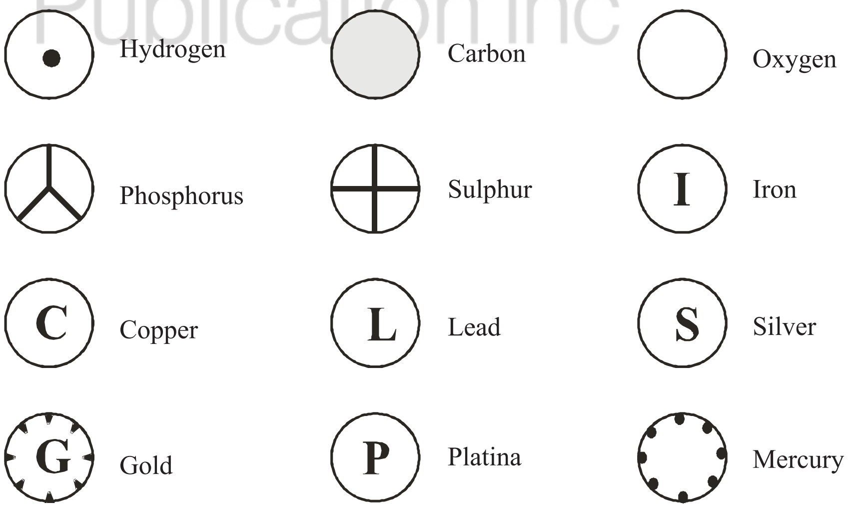
Modern technology has made it possible to take photograph of atoms. The scanning tunneling microscope (STM) is a very sophisticated instrument. It can produce image of the surfaces of the elements which shows the individual atoms. Experiments have demonstrated that the STM probe can be used to move individual atoms or molecules.

Atomic Symbols

For the sake of convenience, the chemists instead of writing the full and lengthy names of elements preferred the use of notations or symbols. In chemistry the history of symbols is very old. Alchemists used the symbol system. They represented gold by the sun, silver by the moon, iron by the planet Mars and so on.



Dalton's used certain symbols for the 36 elements that he believed to exist. Some of Dalton's elements shown below



Berzelius proposed that the symbols of elements be made from one or two letters of the name of the element. Now-a-days, IUPAC (International Union of Pure and Applied Chemistry) approves names of elements. Many of the symbols are the first one or two letters of the element's name in English. The first letter of a symbol is always written as a capital letter (uppercase) and the second letter as a small letter (lowercase).

Sometimes when the names of two elements start with same two letters. We use completely different letter to represent one of them like in case of potassium and polonium where potassium is represented by letter K and polonium as Po.

CHECK POINT-3

What could be possible symbols for elements platinum and palladium?

Sol. Symbol for platinum is 'Pt' and for palladium is 'Pd'.

Elements Named after Places

Scandium (Sc) - found in Scandinavia

Thulium (Tm) - named after Thule, the earlier name of Scandinavia.

Europium (Eu) - after the continent Europe.

Polonium (Po) - named after Curies after their home town.

Named after Celestial Bodies

Selenium (Se) - 'Seles' Greek name for the moon

Plutonium (Pu), Neptunium (Np), Uranium (U) – Named after Planets

Mercury was named after a planet but derives its symbol Hg from the Latin word 'Hydragyrum' meaning liquid silver.

Named after Scientists

Curium (Cm) after Pierre and Marie Curie

Fermium (Fm) after Enrico Fermi

Einsteinium (Es) after Albert Einstein

Mendelevium (Md) after Dimitri Mendeleev

Dalton was the first scientist to use the symbols to represent elements in a very specific sense.

Table : Symbol for some elements

Element	Symbol	Element	Symbol	Element	Symbol
Aluminium	Al	Copper	Cu	Nitrogen	N
Argon	Ar	Fluorine	F	Oxygen	O
Barium	Ba	Gold	Au	Potassium	K
Boron	B	Hydrogen	H	Silicon	Si
Bromine	Br	Iodine	I	Silver	Ag
Calcium	Ca	Iron	Fe	Sodium	Na
Carbon	C	Lead	Pb	Sulphur	S
Chlorine	Cl	Magnesium	Mg	Uranium	U
Cobalt	Co	Neon	Ne	Zinc	Zn

Let's Do Activity

Go through the concept of symbols of elements thoroughly and ask your friends to do the same. Call your friends at your home and divide them in two groups A and B and you also become a part of one group. Now ask your father to conduct a quiz in which he first write symbol of an element at black board. First group A have to answer the name of element then again different symbol should be written for group B. 10 points will be given for right answer and 5 points should be deducted for wrong answer. Let's see which group wins?

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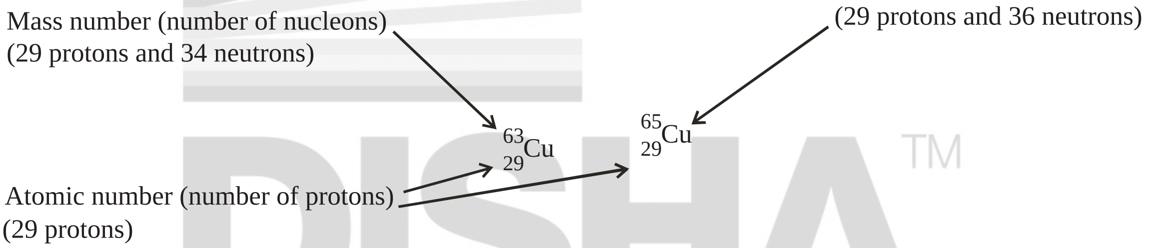
Atomic Mass

In Dalton's original atomic theory, he suggested that all atoms of an element are the same. But if we look at a number of atoms of most elements we find that this statement is not exactly true. For example, consider the atoms of the element copper. A typical atom is composed of a nucleus containing a total of 63 nucleons, of which 29 are protons and 34 are neutrons. The atom also contains 29 electrons that exactly balance the positive charge of the protons. Thus the atom is neutral. The number of protons in the nucleus (which is equal to the total positive charge) is referred to as the atom's atomic number. The total number of nucleons is called the mass number. Therefore, this particular copper atom has an atomic number of 29 and a mass number of 63. There are other copper atoms that are not exactly the same, however. Some atoms of copper have a mass number of 65 rather than 63. This means that these atoms have 36 neutrons and 29 protons. Atoms having the same atomic number but different mass numbers are known as isotopes.

> **DID YOU KNOW?**
> *Most elements that are present in nature exist as a mixture of isotopes. It is the atomic number, however, that distinguishes one element from another. Any atom with an atomic number of 29, regardless of any other consideration, is an atom of copper. If the atomic number is 28, the element is nickel; if it is 30, the element is zinc.*

Specific isotopes are written in a form known as the isotopic notation. In isotopic notation, the mass number is written as a superscript to the left of the element. The atomic number is written as a subscript also on the left. The indication of the atomic number is strictly a convenience, however, since the atomic number determines the identity of the element and vice versa. The isotopic notation for the two isotopes of copper are written as follows.



The convention for verbally naming specific isotopes is to give the element followed by the mass number. For example ^{63}Cu is called "copper-63" and ^{65}Cu is called "copper-65."

Now, another important consideration will be how the mass of one element compares to another. The mass of the electrons is extremely small compared to the masses of the protons and neutrons, so it is not included in the mass of an isotope. Thus the mass number of an isotope is a convenient but rather imprecise measure of its mass. It is imprecise not only because electrons are not included but also because the proton and neutron do not have exactly the same mass. A more precise measure of the mass of one isotope relative to another is known as the *isotopic mass*.

But how to express this mass, Dalton suggested that by considering the weight of hydrogen as standard, the comparative weight of atoms of other elements can be calculated. According to hydrogen standard.

$$\text{Atomic weight of element} = \frac{\text{Weight of one atom of element}}{\text{Weight of one atom of hydrogen}}$$

Later on in comparison to hydrogen atom the sixteenth part of atomic weight of oxygen (16) atom was considered as standard unit because oxygen forms more stable compounds in comparison to hydrogen.

$$\text{Atomic weight of element} = \frac{\text{Weight of one atom of element}}{\text{Weight of } 1/16^{\text{th}} \text{ part of oxygen atom}}$$

These values were based on oxygen-16. But if all the isotopes of oxygen ($_8O^{16}$, $_8O^{17}$, $_8O^{18}$) are included, then the atomic weight of oxygen comes out to be 16.0044. To overcome this problem in 1961, twelfth part of C-12 was considered as standard atomic mass unit (amu) at the International level. Thus according to it, Atomic weight of element = $\dfrac{\text{Mass of an atom of element}}{\frac{1}{12}\times \text{Mass of an atom of C-12}}$

1 amu or u = 1.66056×10^{-24}g.

It has been found even theoretically, that a substance can't be obtained in absolutely pure form and similarly the naturally occuring elements contains more than one isotopes. Isotopes are the atoms of same element having same atomic number but different atomic masses. In such cases the atomic mass reported for an isotopic mixture of an element is actually average atomic mass.

DID YOU KNOW?

It can be calculated as follows :

Average atomic mass

$$= \frac{\Sigma(\%\, abundance \times isotopic\ mass)}{100} \quad or$$

Average atomic mass =

Σ(Isotopic mass × Fractional abundance)

CHECK POINT-4

Why atomic masses are the average values?

Sol. Most of the elements exist in different iostopes i.e. atoms with different masses, e.g., Cl has two isotopes with mass numbers 35 and 37 existing in the ratio 3 : 1. Hence, average value is taken.

Gram Atomic Mass or Gram Atomic Weight

The atomic mass of an element when expressed in grams is known as **gram atomic mass** or simply as **gram atom (g-atom)**. Thus one g-atom of carbon (C-12) weighs 12.0 g. The number of gram-atoms of any element in its certain weight is given by:

Number of gram-atom $= \dfrac{\text{Mass of the element in grams}}{\text{Atomic mass}}$

Thus One g-atom of carbon = 12.000 g carbon = Mass of N_A carbon atoms.

$N_A = 6.022 \times 10^{23}$, the Avogadro's number

Therefore for an element

One gram atom = Atomic mass in gram = 6.022×10^{23} atoms

No. of g-atoms in a given quantity of elements can be calculated by the formula given below :

No. of g-atoms $= \dfrac{\text{Mass in gram}}{\text{Gram atomic mass}}$

and No. of atoms in given mass

= No. of gram atoms × 6.022×10^{23}

DID YOU KNOW?

One gram atomic mass of any element contains the same number of atom of that element as there are carbon atoms in exactly 12 g of carbon-12. This number is Avogadro's number 6.022×10^{23} atoms/g.

DID YOU KNOW?

In accurate determination, atomic mass has been found to be fractional (e.g., Cl = 35.5, etc.,) but here in common practice, its nearby whole number has been used.

Thus, one atomic mass unit is a mass of exactly 1/12 of the mass of ^{12}C. For example, precise measurements show that the mass of ^{10}B is 0.83442 times the mass of ^{12}C, which means it has an isotopic mass of 10.013 amu. From similar calculations we find that the atomic mass of ^{11}B is 11.009 amu.

CHECK POINT-5

Why are the atomic masses of most of the elements fractional?

Sol. This is because atomic masses are the relative masses of atoms as compared with an atom of C-12 isotope taken as 12.

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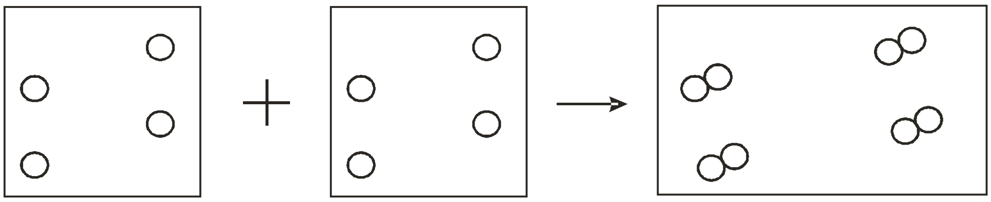
MOLECULE

The smallest particle of an element or compound which can exist independently in nature is termed as molecule. Molecules are formed by the combination of two or more atoms in a constant ratio. If same type of atoms are present in molecule then it is termed as homoatomic molecule.

For example : N_2 (Nitrogen molecule), O_2 (Oxygen molecule), Cl_2 (Chlorine molecule)

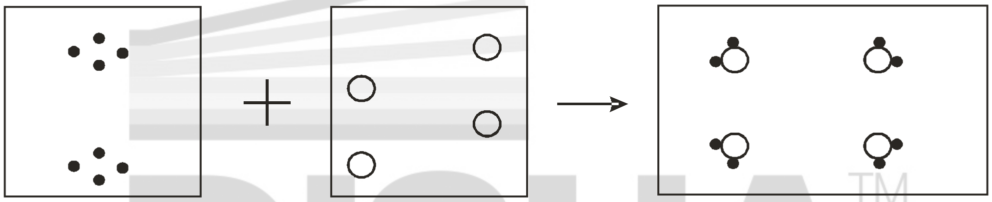


Atoms of nitrogen Atoms of nitrogen Molecules of nitrogen

(Atoms of the same element) (Formation of homodiatomic molecules)

If two or more than two types of atoms are present in any molecule then it is termed as heteroatomic molecule.

For example – H_2O (Water), CO_2 (Carbon dioxide), H_2O_2 (Hydrogen peroxide) etc.



Atoms of hydrogen Atoms of oxygen Molecules of water (H_2O)

(Atoms of different types of elements) (Formation of heteroatomic molecules)

Molecular Mass or Molecular Weight

The molecular mass of a substance (element or compound) is the number or times the molecule of the substance is heavier than 1/12th the mass of an atom of carbon-12 isotope.

Molecular weight is calculated by adding the atomic weights of all the constituent atoms present in a molecule. For example.

Molecular weight of a molecule of hydrogen (H_2) = 2 × atomic weight of hydrogen = 2 × 1 = 2 amu

Similarly, molecular weight of some compounds are:

Water (H_2O) = 2 × 1 + 1 × 16 = 2 + 16 = 18 amu.

Carbon dioxide (CO_2) = 1 × 12 + 2 × 16 = 12 + 32 = 44 amu

Molecular weight is expressed in amu.

Illustration 5 :

Calculate the molecular mass of glucose ($C_6H_{12}O_6$) molecule.

Solution :

Molecular mass of glucose ($C_6H_{12}O_6$)

= 6 (12.011 amu) + 12(1.008 amu) + 6 (16.00 amu)

= 72.066 amu + 12.096 amu + 96.000 amu = 180.162 amu.

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Illustration 6 :

Calculate the molecular masses of the following compounds :

(a) copper sulphate crystals, $CuSO_4.5H_2O$

(b) cane sugar, $C_{12}H_{22}O_{11}$

Solution :

(a) The relative molecular mass of $CuSO_4.5H_2O = 63.5 + 32 + (16 \times 4) + 5\,(2 + 16) = 159.5 + 90 = 249.5$

(b) The relative molecular mass of $C_{12}H_{22}O_{11} = 12 \times 12 + 1 \times 22 + 16 \times 11 = 144 + 22 + 176 = 342$

CHECK POINT-6

Column-II gives molecular mass in amu for substance in column-I Match them correctly.

Column-I	Column-II
(A) H_2O	(p) 58.5
(B) HNO_3	(q) 111
(C) NaCl	(r) 18
(D) $CaCl_2$	(s) 63

Sol. A → (r), B → (s), C → (p), D → (q)

IONS

Imagine that we have a piece of copper metal. Would it be easier to change the number of electrons or number of protons on some of the copper atoms in this piece of metal?

It is much easier to change the number of electrons on an atom than the number of protons in the nucleus. The best evidence for this is the fact that copper can conduct an electric current. Since the number of protons determines the identity of an atom and the copper wire does not change to some other element while conducting electricity, the moving charged particles must be electrons not protons.

The electrically charged particles formed when electrons are added to or removed from a neutral atom are called ions. Neutral atoms are turned into positively charged ions called cations by removing one or more electrons. By removing an electron from a sodium atom-which contains 11 electrons and 11 protons a Na^+ ion is produced that has 10 electrons and 11 protons. Ions with larger positive charges can be produced by removing more electrons. A neutral aluminum atom, for example, has 13 electrons and 13 protons. If we remove three electrons from this atom, we get a positively charged Al^{3+} ion that has 10 electrons and 13 protons, for a net charge of + 3.

Atoms that gain extra electrons become negatively charged ions called anions. A neutral chlorine atom, for example, has 17 protons and 17 electrons. By adding one more electron to this atom, a Cl^- ion is produced that has 18 electrons and 17 protons, for a net charge of –1.

DID YOU KNOW?

The gain or loss of electrons by an atom to form negative or positive ion has an enormous impact on the chemical and physical properties of the atom. Sodium metal, which consists of neutral sodium atoms, bursts into flame when it comes in contact with water. But positively charged Na^+ ions are so unreactive with water that they are essentially inert. Neutral chlorine atoms instantly combine to form Cl_2 molecules, which are so reactive that entire communities are evacuated when train carrying chlorine gas derail. However, chloride ions do not react with one another.

CHECK POINT-7

Is there any difference between property of atom and ion ? Explain with example.

Sol. The identity of an element is not altered when it changeds into ion, however the properties of ion differ from those of atoms e.g., magnesium metal is appropriately reactive but its divalent ion is inert in nature because of stable electronic configuration.

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CASE STUDY-1 : Composition of Matter

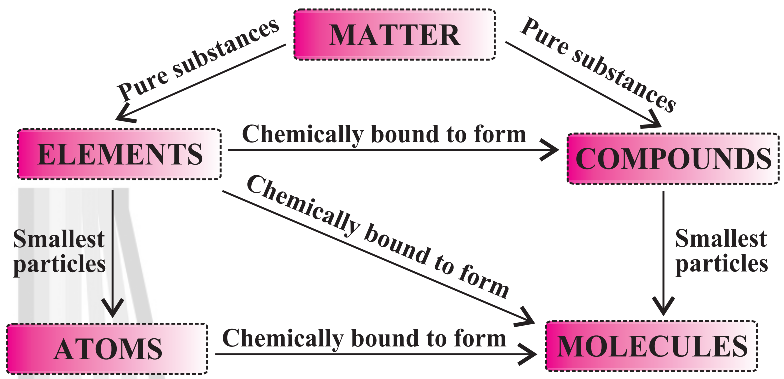
CASE - I :

Atom is the smallest particles of an element that can take part in a chemical reaction and does not exist in free state (except noble gases). Every solid, liquid, gas, and plasma is composed of neutral or ionized atoms. Atoms are extremely small; typical sizes are around 100 picometers. Due to smaller size clear image of an atoms cannot be seen even under the microscope.

An element is a substance that is entirely made up from one type of atom. For example, the element hydrogen is made from hydrogen atoms containing just one proton and one electron.

A molecule is a particle made up of two or more atoms that are chemically bonded together. It is the smallest particles of a substance which have the properties of that substance and can exist in free state. For example, HCl(g) is a molecule made up of one hydrogen atom bonded to one chlorine atom. It is made up of two atoms and is called a diatomic molecule.

Molecule of a compound is made up from two or more atoms of different elements.



CASE - II :

An Element is a substance that is entirely made up of from one kind of atoms. For example element hydrogen, a molecule of hydrogen is made up from two atoms of hydrogen.

CASE - III :

An ion is a positively or negatively charged particle. If an atom gains or losses an electron then it becomes an ion. A positively charged ion is called cation whereas negatively charged ion is called anions. This can happen by chemical change or by receiving energy in the form of heat or radiation. Moreover, ions can be created by chemical means, such as the dissolution of a salt into water, or by physical means, such as passing a direct current through a conducting solution, which will dissolve the anode via ionization. Therefore, all the metal atoms can lose electrons easily so they form cations whereas non-metal atoms can gain electron easily so they form anion. In combined state ions are called ionic compounds, which means those compounds which are made up of cations and anions are known as ionic compounds. For example, Sodium chloride is an ionic compound which is made up of cation, Na^+, and anion, Cl^-.

Think Out of the Box

Q.1 How many atoms are present in the molecule of each of the following compounds?

(i) Aluminium sulphate (ii) Magnesium nitrate (iii) Aluminium bromide (iv) Cesium chloride

Q.2 What is the atomicity of the following elements?

(i) Sulphur (ii) Helium (iii) Nitrogen (iv) Flourine (v) Agron

Q.3 Write the cations and anions are present in the following compounds.

(i) Na_2O (ii) Na_2CO_3 (iii) $CaCl_2$ (iv) K_2SO_4 (v) KNO_3

Polyatomic Ions :

Simple ions, such as Mg^{2+} and N^{3-} ions, are formed by adding or subtracting electrons from neutral atoms. Polyatomic ions are electrically charged substances composed of more than one atom. There are only two polyatomic cations that you will commonly encounter. These are NH_4^+ and H_3O^+. There are many more polyatomic anions, more common anions are –

–1 ions			
HCO_3^-	hydrogen carbonate (bicarbonate)	OH^-	hydroxide
$CH_3CO_2^-$	acetate	ClO_4^-	perchlorate
NO_3^-	nitrate	ClO_3^-	chlorate
NO_2^-	nitrite	ClO_2^-	chlorite
MnO_4^-	permanganate	ClO^-	hypochlorite
CN^-	cyanide		
–2 ions			
CO_3^{2-}	carbonate	O_2^-	peroxide
SO_4^{2-}	sulfate	CrO_4^{2-}	chromate
SO_3^{2-}	sulfite	$Cr_2O_7^{2-}$	dichromate
$S_2O_3^{2-}$	thiosulfate		
–3 ions			
PO_4^{3-}	phosphate	AsO_4^{3-}	arsenate
BO_3^{3-}	borate		

Let's Do Activity

Make a power point presentation on formation of ion from neutral atom. Your presentation should clearly diffrentiate between the ions of opposite charges. Also mention about simple and polyatomic ions in your presentation with some examples.

CHEMICAL FORMULA

The composition of a molecular compound can be represented by a chemical formula. The subscripts in a chemical formula represent the relative numbers of atoms present in the compound. When there is no subscript, as in the case of carbon in CO_2, a value of one is assumed. Thus the formula CO_2 represents a molecule that contains one carbon atom and two oxygen atoms. The same elements can combine in different ratios to give completely different compounds. Carbon dioxide, CO_2, is the gas that results from respiration (we exhale CO_2). Carbon and oxygen can also combine to form carbon monoxide, CO, a very toxic gas. In this compound carbon and oxygen are in a 1 : 1 ratio.

We use coefficients in front of chemical formulas to represent the number of units of a compound present. If we wish to describe three molecules of carbon dioxide, we would write 3 CO_2.

How many atoms are in a molecule of $CaSO_4$ (calcium sulphate)? There is one calcium (Ca), one sulfur (S), and four oxygen (O) for a total of six atoms. What is the formula for ammonium chloride? Since it contains one nitrogen (N), four hydrogen (H), and one chlorine (Cl), the chemical formula is NH_4Cl.

Have you ever thought what to call a chemical compound ? There are a few general rules that can help you. For example, for CaO start with the name of the first atom, Ca., i.e. the element calcium. Then, take the next atom and replace the ending with "ide." Oxygen (O) becomes "oxide". Thus the name for CaO is calcium oxide. As with many rules, though, there are exceptions! In this case, the first part of the name for the last atom, hydrogen (H), is tacked on to the name of the first atom. Thus the correct chemical name for $Ca(OH)_2$ is calcium hydroxide.

Prefixes, the part added to the beginning of a word, can help you figure out names and formulas, too. Mono means one, di means two, So a compound with two of the same molecule (represented by a subscript 2) would have "di" before that molecule name. What would you call CO_2 ? (Carbon dioxide). How many oxygen atoms are in carbon monoxide? (one)

One carbon (C) and four chlorine (Cl) atoms combine to give carbon tetrachloride.

(1) Carbon + (4) Chlorine = Carbon tetrachloride.

HOW TO WRITE CHEMICAL FORMULA OF A COMPOUND?

The formation of a chemical compound (say AB) takes place by the combination of positive radical (say A) and negative radical (say B) to give a neutral molecule. Thus total positive charge due to A will be equal to the total negative charge due to B.

The chemical formula of a compound can be written in following steps –

Step 1. Write the symbols or formulae of the positive and negative radicals side by side. Positive radical should be written to the left and negative radical to the right. If necessary write the formulae of the radicals in the brackets.

Step 2. Write the valency of each radical at the top. If necessary divide the valencies by H. C. F. to get a simple ratio.

Step 3. Now apply the rule of criss-cross by shifting the valencies cross-wise.

The above method of writing a chemical formula can be illustrated by the following examples :

Compound	Symbol with valencies	Criss-cross	Molecular Formula
Sodium chloride	$\overset{1}{Na}$ $\overset{1}{Cl}$	1 1 Na Cl	NaCl
Sodium oxide	$\overset{1}{Na}$ $\overset{2}{O}$	1 2 Na O	Na_2O
Chromium nitrate	$\overset{3}{Cr}$ $\overset{1}{(NO_3)}$	3 1 Cr (NO_3)	$Cr(NO_3)_3$

When the valencies of the radicals are divisible by a common factor they are first reduced to a simple ratio. Then criss-cross is applied. Examples :

Compound	Symbols	Simple ratio	Criss-cross	Formula
Calcium sulphate	$\overset{2}{Ca}$ $\overset{2}{(SO_4)}$	$\overset{1}{Ca}$ $\overset{1}{(SO_4)}$	1 1 Ca (SO_4)	$CaSO_4$
Stannic sulphate	$\overset{4}{Sn}$ $\overset{2}{(SO_4)}$	$\overset{2}{Sn}$ $\overset{1}{(SO_4)}$	2 1 Sn (SO_4)	$Sn(SO_4)_2$

Some more examples are given in the following table.

Compound	Symbol with valencies	Criss-Cross	Formula
1. Phosphorus trichloride	$\overset{3}{P}$ $\overset{1}{Cl}$	3 1 P Cl	PCl_3
2. Ferrous sulphate	$\overset{2}{Fe}$ $\overset{2}{SO_4}$	2 2 Fe SO_4	$FeSO_4$

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3. Ferric sulphate	$\overset{3}{Fe}\quad \overset{2}{SO_4}$	$\overset{3}{Fe} \times \overset{2}{SO_4}$	$Fe_2(SO_4)_3$
4. Ferric chloride	$\overset{3}{Fe}\quad \overset{1}{Cl}$	$\overset{3}{Fe} \times \overset{1}{Cl}$	$FeCl_3$
5. Barium nitrate	$\overset{2}{Ba}\quad \overset{1}{NO_3}$	$\overset{2}{Ba} \times \overset{1}{NO_3}$	$Ba(NO_3)_2$
6. Ammonium phosphate	$\overset{1}{(NH_4)}\quad \overset{3}{(PO_4)}$	$\overset{1}{(NH_4)} \times \overset{3}{(PO_4)}$	$(NH_4)_3PO_4$
7. Bismuth carbonate	$\overset{3}{Bi}\quad \overset{2}{(CO_3)}$	$\overset{3}{Bi} \times \overset{2}{(CO_3)}$	$Bi_2(CO_3)_3$
8. Potassium permanganate	$\overset{1}{K}\quad \overset{1}{(MnO_4)}$	$\overset{1}{K} \times \overset{1}{MnO_4}$	$KMnO_4$
9. Aluminium phosphate	$\overset{3}{Al}\quad \overset{3}{(PO_4)}$	$\overset{3}{Al} \times \overset{3}{(PO_4)}$	$Al_3(PO_4)_3$ $Al\,PO_4$
10. Stannous chloride	$\overset{2}{Sn}\quad \overset{1}{Cl}$	$\overset{2}{Sn} \times \overset{1}{Cl}$	$SnCl_2$

Formulae of some compounds

Negative radical	Chloride	Nitrate	Sulphate	Carbonate	Hydroxide	Phosphate
	Cl^-	NO_3^-	SO_4^{2-}	CO_3^{2-}	OH^-	PO_4^{3-}
Positive radical						
H^+	HCl	HNO_3	H_2SO_4	H_2CO_3	HOH or H_2O	H_3PO_4
Na^+	$NaCl$	$NaNO_3$	Na_2SO_4	Na_2CO_3	$NaOH$	Na_3PO_4
Mg^{2+}	$MgCl_2$	$Mg(NO_3)_2$	$MgSO_4$	Mg_2CO_3	$Mg(OH)_2$	$Mg_3(PO_4)_2$
Al^{3+}	$AlCl_3$	$Al(NO_3)_3$	$Al_2(SO_4)_3$	$Al_2(CO_3)_3$	$Al(OH)_3$	$AlPO_4$
NH_4^+	NH_4Cl	NH_4NO_3	$(NH_4)_2SO_4$	$(NH_4)_2CO_3$	NH_4OH	$(NH_4)_3PO_4$
Fe^{2+} (ous)	$FeCl_2$	$Fe(NO_3)_2$	$FeSO_4$	$FeCO_3$	$Fe(OH)_2$	$Fe_3(PO_4)_2$
Fe^{3+} (ic)	$FeCl_3$	$Fe(NO_3)_3$	$Fe_2(SO_4)_3$	$Fe_2(CO_3)_3$	$Fe(OH)_3$	$FePO_4$
Ba^{2+}	$BaCl_2$	$Ba(NO_3)_2$	$BaSO_4$	$BaCO_3$	$Ba(OH_2)$	$Ba_3(PO_4)_2$

CHECK POINT-8

Calcium carbonate is the major component in chalk, limestone, and marble. It is composed of calcium ions, Ca^{2+} and the polyatomic anion carbonate, CO_3^{2-}. Predict the chemical formula of calcium carbonate.

Sol. The total charge from the positive ions must balance the charge from the negative ions. Since the calcium ion has a charge of +2 and the carbonate ion has a charge of –2, the charge of one calcium ion cancels the charge of one carbonate ion. The ionic formula of calcium carbonate is $CaCO_3$.

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CONNECTING TOPIC

DETERMINATION OF EMPIRICAL AND MOLECULAR FORMULAE

The chemical formula of a substance is symbolic representation indicating the number and kind of atoms of different elements present in one molecule of the substance. For example, chemical formula of water is H_2O and that of ammonia is NH_3. Similarly, chemical formula of sulphur dioxide is SO_2 and that of sodium carbonate is Na_2CO_3. In order to arrive at the chemical formula of any compound, following two analysis have to be carried out.

(i) Elemental analysis of the substance i.e., determination of the types of elements.

(ii) Determination of relative number of each type of elements.

This data is then used to determine first the emperical formula and then molecular formula of the substance.

(a) Empirical Formula :

It expresses the whole number ratio of the atoms of various elements present in one molecule of the compound. For example, the empirical formula of glucose ($C_6H_{12}O_6$) is CH_2O. This shows that C, H and O are present in the simple ratio of 1 : 2 : 1.

Steps Taken to Arrive at the Empirical Formula :

(i) Determine the percentage of each element : Percentage of each element is determined quantitatively. However, the percentage of oxygen is obtained by subtracting the sum of all other percentages from 100 i.e., % of O = [100 – sum of % of all other constituent elements].

(ii) Divide the percentage of each element by its atomic mass to obtain the relative numbers of atoms (or atomic ratio).

$$\text{Atomic ratio} = \frac{\text{Percentage of an element}}{\text{Atomic mass of the same element}}$$

(iii) Divide the atomic ratio by smallest quotient to get the simplest ratio of various elements.

(iv) Convert the simplest ratio to the whole number ratio either by

(1) Rounding them off to whole number, if the value is quite close to whole number e.g., 1.98, 2.99, 3.95 are rounded off as 2, 3 and 4 respectively.

(2) Multiply the figures by suitable integer (2, 3, or 4, etc.)

(v) Write the empirical formula of the compound by writing the symbols of the various elements side by side. Now insert the whole number ratio of each element as the subscripts to the lower right hand corner of each symbol.

(b) Molecular Formula :

It is that formula of the compound which gives the actual number of atoms of various elements in a molecule of that compound. For example, molecular formula of glucose is $C_6H_{12}O_6$. Molecular formula of a compound is related to empirical formula as being a simple whole number multiple of empirical formula. Thus,

Molecular formula = n × Empirical formula [where n = 1, 2, 3, ..., etc.]

The value of n is obtained by the following relation

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

When n = 1, Molecular Formula = Empirical formula

Illustration 7 :

An inorganic salt gave the following percentage composition : Na = 29.11, S = 40.51 and O = 30.38. Calculate the empirical formula of the salt.

Solution :

Calculation of empirical formula

Element	Symbol	Percentage of element	At. mass of element	Moles of the element = Percentage At. mass (Relative no. of moles)	Simplest molar ratio	Simplest whole no. molar ratio
Sodium	Na	29.11	23	$\frac{29.11}{23} = 1.266$	$\frac{1.266}{1.266} = 1$	$1 \times 2 = 2$
Sulphur	S	40.51	32	$\frac{40.51}{32} = 1.266$	$\frac{1.266}{1.266} = 1$	$1 \times 2 = 2$
Oxygen	O	30.38	16	$\frac{30.38}{16} = 1.897$	$\frac{1.89}{1.266} = 1.5$	$1.5 \times 2 = 3$

Thus, the empirical formula is $Na_2S_2O_3$.

Illustration 8 :

2.38 g of uranium was heated strongly in a current of air. The resulting oxide weighed 2.806 g. Determine the empirical formula of the oxide. (At. mass U = 238; O = 16).

Solution :

Step 1. To calculate the percentage of uranium and oxygen in the oxide.

2.806 g of the oxide contain uranium = 2.38 g.

$\therefore$ Percentage of uranium = $\frac{2.38}{2.806} \times 100 = 84.82$

Hence, the percentage of oxygen in the oxide = 100.00 – 84.82 = 15.18.

Step 2. To calculate the empirical formula

Element	Symbol	Percentage of element	At. mass of element	Moles of the element = Percentage At. mass (Relative no. of moles)	Simplest molar ratio	Simplest whole no. molar ratio
Uranium	U	84.82	238	$\frac{84.82}{238} = 0.3562$	$\frac{0.3562}{0.3562} = 1$	$1 \times 3 = 3$
Oxygen	O	15.18	16	$\frac{15.18}{16} = 0.94875$	$\frac{0.94875}{0.3562} = 2.666$	$2.666 \times 3 = 8$

Hence, the empirical formula of the oxide is U_3O_8.

Let's Connect

1. In a compound C, H and N atoms are present in 9 : 1 : 3.5 by weight. Molecular weight of compound is 108. Molecular formula of compound is

(a) $C_2H_6N_2$ (b) C_3H_4N (c) $C_6H_8N_2$ (d) $C_9H_{12}N_3$.

2. An organic compound contains carbon, hydrogen and oxygen. Its elemental analysis gave C, 38.71% and H, 9.67%. The empirical formula of the compound would be:

(a) CH_3O (b) CH_2O (c) CHO (d) CH_4O

3. A gaseous hydrocarbon gives upon combustion 0.72 g of water and 3.08 g. of CO_2. The empirical formula of the hydrocarbon is :

(a) C_2H_4 (b) C_3H_4 (c) C_6H_5 (d) C_7H_8

4. Write the formulae and names of three compounds containing same percentage composition of C, H and O.

5.

Column-I (Empirical formula)	**Column-II (Molecular formula)**
(A) CH	(p) N_2H_4
(B) NH_2	(q) C_2H_4
(C) CH_2O	(r) C_2H_2
(D) CH_2	(s) C_6H_6
	(t) $C_2H_4O_2$

	A	B	C	D
(a)	(r, s)	(p)	(t)	(q)
(b)	(p)	(r, s)	(t)	(q)
(c)	(r, s)	(t)	(p)	(q)
(d)	(q)	(p)	(t)	(r, s)

Sol. 1. (c)

Element	Percentage	R.N.A	Simplest ratio
C	9	$\frac{9}{12}=\frac{3}{4}$	3
H	1	$\frac{1}{1}=1$	4
N	3.5	$\frac{3.5}{14}=\frac{1}{4}$	1

Empirical formula = C_3H_4N

$(C_3H_4N)_n = 108 = (12 \times 3 + 4 \times 1 + 14)_n = 108$

$(54)_n = 108$

$$n = \frac{108}{54} = 2$$

$\therefore$ Molecular formula = $C_6H_8N_2$

2. (a)

Element	Percentage	Atomic weight	Relative number of atoms	Simple ratio
C	38.71	12	$\frac{38.71}{12}=3.23$	$\frac{3.23}{3.23}=1$
H	9.67	1	$\frac{9.67}{1}=9.67$	$\frac{9.67}{3.23}=3$
O	100 – (38.71 + 9.67) = 51.62	16	$\frac{51.62}{16}=3.23$	$\frac{3.23}{3.23}=1$

Hence the empirical formula of the compound is CH_3O

3. **(d)** $\because$ 18 g, H_2O contains = 2 gm H

$\therefore$ 0.72 g H_2O contains = $\frac{2}{18} \times 0.72$ gm = 0.08 gm H

$\because$ 44 g CO_2 contains = 12 gm C

$\therefore$ 3.08 g CO_2 contains = $\frac{12}{44} \times 3.08 = 0.84$ gm C

$\therefore$ C : H = $\frac{0.84}{12} : \frac{0.08}{1}$ = 0.07 : 0.08 = 7 : 8

$\therefore$ Empirical formula = C_7H_8

4. Compounds with the same percentage composition of C, H and O will have the same empirical formula. The compounds with the empirical formula CH_2O can be

HCHO	CH_3COOH	$C_6H_{12}O_6$
Formaldehyde	Acetic acid	Glucose

5. A - (r, s), (B) - (p), (C) - (t), (D) - (q)

Gram Molecular Mass

The molecular mass of a substance when expressed in grams is known as **gram-molecular mass** or simply **gram-mole** (g-**mole**) for the sake of convenience it is expressed simply as **mole**. Thus one **mole** (or **g-mole**) of water weighs 18g (molecular weight of water = 18)

$$\text{Number of moles} = \frac{\text{Mass of substance in grams}}{\text{Molecular weight}}$$

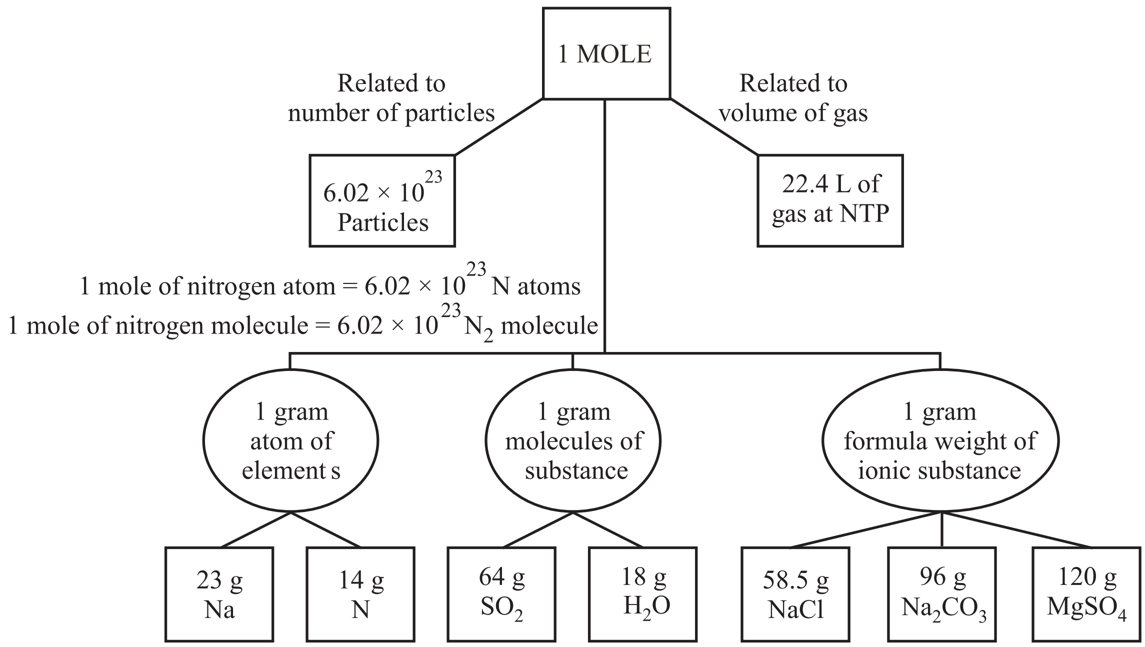
Formula Mass and Gram Formula Mass

Ionic compounds such as NaCl, KNO_3, Na_2CO_3 etc. do not consist of molecules, i.e., single entities but exist as ions. Each ion is surrounded by a number of oppositely charged ions, e.g., in NaCl, each Na^+ ion is surrounded by six Cl^- ions and vice versa. Hence, in such cases, the formula is used to calculate the formula mass instead of molecular mass. Just like the molecular mass, it is found by adding the atomic masses of the atoms present in one formula unit, e.g.,

Formula mass of NaCl = Atomic mass of Na^+ Atomic mass of Cl = 23.0 u + 35.5 u = 58.5 u

Gram formula mass represents the formula mass expressed in grams.

Summing up the above discussed various relationships of mole, a mole may be illustrated as follows :



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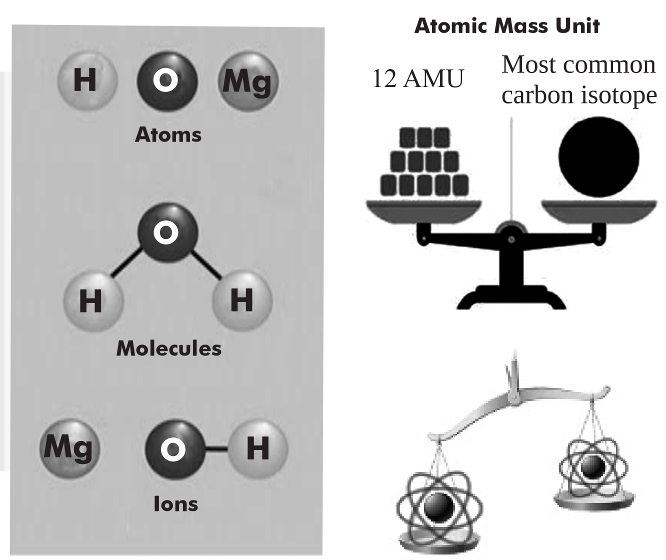
CASE STUDY-2 : Molecular and Formula Unit Mass

CASE - I :
The term molecular formula represents the simplest ratio of the number of atoms of the elements present. Formula unit represents this ratio for ionic compounds.
E.g., H_2O represents a molecular formula while NaCl represents for formula unit.

CASE - II :
The mass corresponding to a particular molecular formula is called the molecular mass.
The mass corresponding to a particular formula unit is called the formula unit mass.
E.g., Molecular mass of $H_2O = (2 \times 1) + (16) = 18$ u
Formula unit mass of NaCl = 23 + 35.5 = 58.5 u.

CASE - III ::
For a given molecular formula and formula unit :
(i) Write the atomic masses of the corresponding elements.
(ii) Multiply the masses with their corresponding number of atoms.
(iii) Calculate the sum of the masses to get the molecular or formula unit mass.
E.g., Mass of $C_6H_{12}O_6 = (6 \times 12) + (12 \times 1) + (6 \times 16) = 180$ u.



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Think Out of the Box

Q.1 Calculate the molecular mass of carbon dioxide.

Q.2 Calculate the formula unit mass of calcium carbonate.

Q.3 A compound has molecular mass 16 u. What could be its possible molecular formula?

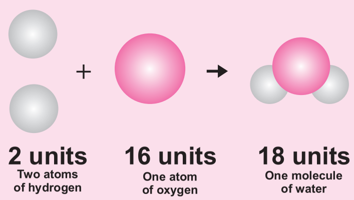
Walk Through the Chapter

Atoms and Molecules

Dalton's Atomic Theory

Postulates
- Each elements are composed of atoms
- Atoms are neither created nor destroyed during chemical reaction
- An atom is the smallest particle

Dalton's Atomic Theory
Example : Formation of Water



Laws of Chemical Combination

Law of Conservation of Mass

In a chemical reaction, the total mass of products is equal to the total mass of the reactants.

Law of Constant Proportion

The element composition of a pure compound is always the same is known as the law of constant composition.

Atoms : Building Blocks of all Matter

Atomic weight of element (u)

$$= \frac{\text{Mass of an atom of element}}{\frac{1}{12} \times \text{mass of an atom of C} - 12}$$

Number of gram-atoms

$$= \frac{\text{Mass of element in grams present}}{\text{Atomic mass in grams}}$$

Ions : Ions are the electrically charged particles formed when electrons are added or removed from neutral atom.

Ex. Na^+, Cl^- F^-, Ca^{2+}, etc.

Molecule

Smallest particle of an element or compound which can exist.

Example :
Element = H_2, He, O_2 etc.
Compounds = H_2O, CO_2, etc.

Chemical Formula :
- Write the symbols
- Write the valency
- Shift the valency cross wise

e.g.,

$$\begin{array}{cc} 1 & 2 \\ Na & O \end{array}$$

$\Rightarrow Na_2O$

Let's Revise Through FIB & T/F

1. During a chemical reaction, the sum of the .... of the reactants and products remains unchanged.
2. In a pure chemical compound, elements are always present in a definite proportion by mass. **(T/F)**
3. Water is an atom. **(T/F)**
4. The abbreviation used for lengthy names of elements are termed as their ....
5. Clusters of atoms that act as an ion are called .... ions.
6. In ionic compounds, the charge on each ion is used to determine the .... of the compound.
7. Those ions which are formed from single atoms are called................. .
8. Ionic componds are formed by the combination between .... and ....
9. Those particles which have more or less electrons than the normal atoms are called ions. **(T/F)**
10. A chemical formula is also known as a ....
11. The valency of an ion is ....to the charge on the ion.
12. Formula for sulphur dioxide is SO_3. **(T/F)**
13. Formula mass of Na_2O is 62 amu. **(T/F)**

EXERCISE-1 Master Board

Multiple Choice Questions (MCQs)

DIRECTIONS : *This section contains multiple choice questions. Each question has four choices (a), (b), (c) and (d) out of which ONLY ONE is correct.*

1. The atomic theory of matter was proposed by :
   (a) John Kennedy (b) Lavoisier
   (c) Proust (d) John Dalton
2. The atoms of which of the following pair of elements are most likely to exist in free state?
   (a) nitrogen (b) neon
   (c) Helium (d) both (b) and (c)
3. In water, the proportion of oxygen and hydrogen by mass is
   (a) 1 : 4 (b) 1 : 8 (c) 4 : 1 (d) 8 : 1
4. The number of electrons in an ion A^{3+} is 10. The atomic number of element A is most likely to be
   (a) 13 (b) 14 (c) 15 (d) 12
5. A particle X has 16 electrons, 15 protons and 18 neutrons. The particle X must be _______.
   (a) Cation (b) anion
   (c) Neutral (d) Compound
6. When an atom loses electrons, it is called a (an) ______ and has a _______ charge.
   (a) anion, positive (b) cation, positive
   (c) anion, negative (d) cation, negative
7. Adding electrons to an atom will result in a (an)
   (a) molecule (b) anion
   (c) cation (d) salt
8. A group of atoms chemically bonded together is a (an)
   (a) molecule (b) ion
   (c) salt (d) element
9. The total number of atoms represented by the compound $CuSO_4.5H_2O$ is
   (a) 27 (b) 21
   (c) 5 (d) 8
10. The percentage of copper and oxygen in samples of CuO obtained by different methods were found to be the same. This illustrates the law of
    (a) constant proportions
    (b) conservation of mass
    (c) multiple proportions
    (d) reciprocal proportions
11. Aspartame, an artificial sweetener, has the molecular formula $C_{14}H_{18}N_2O_5$. What is the mass in grams of one molecule? (Atomic weights: C = 12.01, H = 1.008, N = 14.01, O = 16.00).
    (a) 4.89×10^{-21} (b) 2.24×10^{-21}
    (c) 3.85×10^{-22} (d) 4.89×10^{-22}
12. The law of definite proportions was given by –
    (a) John Dalton (b) Humphry Davy
    (c) Proust (d) Michael Faraday

13. Molecular mass is defined as the
   (a) mass of one atom compared with the mass of one molecule.
   (b) mass of one atom compared with the mass of one atom of hydrogen.
   (c) mass of one molecule of any substance compared with the mass of one atom of C-12.
   (d) None of the above
14. The correct symbol for silver is
   (a) Ag (b) Si
   (c) Ar (d) Al
15. The chemical symbol P stands for
   (a) phosphorus (b) potassium
   (c) polonium (d) promethium
16. The molecular formula P_2O_5 means that
   (a) a molecule contains 2 atoms of P and 5 atoms of O.
   (b) the ratio of the mass of P to the mass of O in the molecule is 2:5.
   (c) there are twice as many P atoms in the molecule as there are O atoms.
   (d) the ratio of the mass of P to the mass of O in the molecule is 5 : 2.
17. The chemical symbol for barium is
   (a) B (b) Ba (c) Be (d) Bi

Assertion & Reason Questions

DIRECTIONS : *Each of these questions contains an assertion followed by reason. Read them carefully and answer the question on the basis of following options. You have to select the one that best describes the two statements.*

(a) If both Assertion and Reason are correct and Reason is the correct explanation of Assertion.
(b) If both Assertion and Reason are correct, but Reason is not the correct explanation of Assertion.
(c) If Assertion is correct but Reason is incorrect.
(d) If Assertion is incorrect but Reason is correct.

1. **Assertion:** An atom is electrically neutral.
   Reason: Atom contains equal number of protons and neutrons.
2. **Assertion:** The sum of protons and neutrons is always different in isobars.
   Reason: Isobars are atoms of different elements having same mass number but different atomic number.
3. **Assertion :** Atoms can neither be created nor destroyed.
   Reason : Atom is made up of electrons, protons and neutron.
4. **Assertion:** The atomicity of ozone is 2.
   Reason: The number of atoms present in one molecule of an element is called its atomicity.
5. **Assertion:** The atomic mass of an element is the relative mass of its atom.
   Reason: The atomic mass of elements are determinal by comparing the mass with the mass of a carbon – 12 atom.

Passage/Case Based Questions

DIRECTIONS : *Study the given passage and answer the related question.*

One of the forms of a naturally occurring solid compound A is usually used for making the floors of houses. On adding a few drops of dilute hydrochloric acid to A, brisk effervescence are produced. When 50 g of reactant A was heated strongly, than 22 g of a gas B and 28 g of a solid C were produced as products. Gas B is the same which produced brisk effervescence on adding dilute HCl to A. Gas B is said to cause global warming whereas solid C is used for white-washing.

1. What is (i) solid A (ii) gas B, and (iii) solid C.
2. What is the total mass of B and C obtained from 50 g of A?
3. How does the total mass of B and C formed compare with the mass of A taken?
4. What conclusion do you get from the comparison of masses of products and reactant?
5. Which law of chemical combination is illustrated by the example given in this problem?

Very Short Answer Questions

1. What do you understand by the term formula unit?
2. Name the following compounds PCl_3 and SO_2.
3. In what smallest whole-number must N and O atoms combine to make dinitrogen tetraxide N_2O_4 ?
4. How many grams of iron are needed to combine with 25.6 g of Oxygen to make Fe_2O_3 ?

Short Answer Questions

1. Convert into moles : 22 g of carbon dioxide
2. What is molecular weight ? Explain with example.
3. Write formulas for ionic compounds formed from
   (a) Na and F (b) Na and O
   (c) Mg and F (d) Al and C
4. Write formulas for
   (a) aluminium sulfide (b) strontium fluoride
   (c) titanium (IV) oxide (d) calcium bromide

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5. Why carbon-12 atom is taken as standard for atom mass calculations ? **(Reasoning)**
6. Why an atom is considered electrically neutral ?
7. Why ionic compounds are neutral in nature ?
8. Why can't we say, 'an atom of water' ?
9. A compound of nitrogen and oxygen has the formula NO. In this compound there are 1.143g of oxygen for each 1.000g of nitrogen. A different compound of nitrogen and oxygen has the formula NO_2. How many grams of oxygen will combine with each 1.000g of nitrogen in NO_2 ?
10. Why different elements show different chemical properties? **(Reasoning)**
11. Why is it not possible to see an atom even with the high resolution microscope ? **(Reasoning)**
12. Why we say that law of conservation of mass is obeyed if a chemical reaction takes place in a closed container? **(Reasoning)**
13. What is the mass of 0.5 mole of water (H_2O). (Atomic masses : H = 1u, O = 16u)
14. Why water collected from different source always has same composition ? **(Reasoning)**
15. What is the percentage composition of the elements in ammonia, NH_3 ? (at. mass: H = 1, N = 14)
16. Calculate the theoretical percentage composition of N_2O_3.
17. 10 grams of $CaCO_3$ on heating gave 4.4 g of CO_2 and 5.6 of CaO. Show that these observations are in agreement with the law of conservation of mass.

Long Answer Questions

1. Calculate the formula mass of each of the following and round your answer to the nearest 0.1 u.
   (a) $NaHCO_3$ (b) $K_2Cr_2O_7$
   (c) $(NH_4)_2CO_3$ (d) $Al_2(SO_4)_3$
   (e) $CuSO_4.5H_2O$
2. (I) Write the chemical formulas for the following compounds: (a) copper (I) oxide, (b) potassium peroxide (c) mercury (I) bromide, (d) iron (III) carbonate, (e) sodium hypobromite.
   (II) Give the name for each of the following acids
   (a) $HBrO_3$, (b) HBr,
   (c) H_3PO_4, (d) HClO
   (e) HIO_3
3. (I) Calculate both the average mass of a single molecule of carbon dioxide and glucose and the molecular weight of these compounds.
   (II) Determine the number of carbon atoms in 0.500 grams of carbon dioxide, CO_2.
4. (i) Distinguish between an atom and a molecule.
   (ii) Differentiate between atom and ion.

HOTS Questions

1. Write formulas for the chlorides and oxides formed by (a) chromium (b) copper.
2. The formula for arsenic acid is H_3AsO_4. What is the name of the salt Na_3AsO_4 ?
3. Sucrose (table sugar) has the formula $C_{12}H_{22}O_{11}$. In this compound, what is the
   (a) atom ratio of C to H ?
   (b) atom ratio of H to O ?
   (c) mole ratio of C to O ?
   (d) mole ratio of H to O ?
4. A sample of ascorbic acid (vitamin C) is synthesized in the laboratory. It contains 1.50g of carbon and 2.00g of oxygen. Another sample of ascorbic acid isolated from citrus fruits contains 6.35g of carbon. How many grams of oxygen does it contain ? Which law are you assuming in answering this question ?
5. Calculate the weight of (i) one atom of oxygen, and (ii) one molecule of oxygen.
   G.A.W. of Oxygen = 16.0g
6. How does Dalton's atomic theory account for the fact that when 1.000g of water is decomposed into its elements, 0.111g of hydrogen and 0.889g of oxygen are obtained regardless of the source of the water ?
7. An atom is 200 times heavier than $\frac{1}{12}$ of the mass of an atom of carbon (C-12). What is its mass in amu (or u)?
8. Find the percentage of nitrogen in urea (H_2NCONH_2).
9. 1.375 g of pure cupric oxide was reduced by heating in a current of pure dry hydrogen and the mass of copper that remained 1.0980 g. In another experiment, 1.179 g of pure copper was dissolved in pure HNO_3 and the resulting copper nitrate converted into cupric oxide by ignition. The mass of copper oxide formed was 1.476 g. Show that the results illustrate the law of constant composition with in the limits of experimental error.
10. Write the chemical formula for each substance mentioned in the following word descriptions (a) Zinc carbonate can be heated to form zinc oxide and carbon dioxide. (b) On treatment with hydrofluoric acid, silicon dioxide forms silicon tetrafluoride and water. (c) Sulfur dioxide reacts with water to form sulfurous acid. (d) The substance hydrogen phosphide, commonly called phosphine, is a toxic gas. (e) Perchloric acid reacts with cadmium to form cadmium (II) perchlorate.

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EXERCISE-2 NCERT Questions

Intext-Questions

1. In a reaction 5.3g of sodium carbonate reacted with 6g of ethanoic acid. The products were 2.2g of carbon dioxide, 0.9g of water and 8.2g of sodium ethanoate. Show that these observations are in agreement with the law of conservation of mass.

   Sodium carbonate + Ethanoic acid $\longrightarrow$
   (5.3g) (6.0g)
   Sodium ethanoate + Carbon dioxide + Water
   (8.2g) (2.2g) (0.9g)

2. Hydrogen and oxygen combine in the ratio of 1: 8 by mass to form water. What mass of oxygen gas would be required to react completely with 3g of hydrogen gas?
3. Which postulate of Dalton's atomic theory is the result of the law of conservation of mass?
4. Which postulate of Dalton's atomic theory can explain the law of definite proportions?
5. Define atomic mass unit.
6. Why is it not possible to see an atom with naked eye?
7. Write down the formula of
   (i) Sodium oxide
   (ii) Aluminium chloride
   (iii) Sodium sulphide
   (iv) Magnesium hydroxide.
8. Write down the names of the following compounds.
   (i) $Al_2(SO_4)_3$ (ii) $CaCl_2$
   (iii) K_2SO_4 (iv) KNO_3
   (v) $CaCO_3$
9. What is meant by the term chemical formula?
10. How many atoms are present in a
    (i) H_2S molecule
    (ii) PO_4^{3-} ion
11. Calculate the molecular masses of H_2, O_2, Cl_2, CO_2, CH_4, C_2H_6, C_2H_4, NH_3, CH_3OH.
12. Calculate the formula unit mass of ZnO, Na_2O, K_2CO_3. Given atomic masses of Zn = 65u, Na = 23u, K = 39u, C = 12u and O = 16u.

Text Book Exercise

1. 0.24 g of sample of a compound of oxygen and boron was found by analysis to contain 0.096 g of boron and 0.144 g of oxygen. Calculate the % age composition of each of the element in the compound by weight.
2. When 3.0 g of carbon is burnt in 8.0 g of oxygen, 11.0 g of carbon dioxide is formed. What mass of carbon dioxide will be formed, when 3.0 g of carbon is burnt in 50.0 g of oxygen? Which law of chemical composition will govern your answer?
3. What are polyatomic ions? Give examples.
4. Write the chemical formula of the following :
   (a) Magnesium chloride (b) Calcium oxide
   (c) Copper nitrate (d) Aluminium chloride
   (e) Calcium carbonate (f) Zinc sulphate
5. Give the names of the elements present in the following compounds :
   (a) Quick lime (b) Hydrogen bromide
   (c) Baking powder (d) Potassium sulphate

Exemplar Questions

1. Which of the following correctly represents 360 g of water?
   (i) 2 moles of H_2O
   (ii) 20 moles of water
   (iii) 6.022×10^{23} molecules of water
   (iv) 1.2044×10^{25} molecules of water
   (a) only (i) (b) (i) and (iv)
   (c) (ii) and (iii) (d) (ii) and (iv)
2. Which of the following statements is not true about an atom?
   (a) Atoms are not able to exist independently.
   (b) Atoms are the basic units from which molecules and ions are formed.
   (c) Atoms are always neutral in nature.
   (d) Atoms aggregate in large numbers to form the matter that we can see, feel or touch.
3. The chemical symbol for nitrogen gas is
   (a) Ni (b) N_2
   (c) N^+ (d) N
4. The chemical symbol for sodium is
   (a) So (b) Sd
   (c) NA (d) Na
5. Which of the following would weight the highest?
   (a) 0.2 mole of sucrose ($C_{12}H_{22}O_{11}$)
   (b) 2 moles of CO_2
   (c) 2 moles of $CaCO_3$
   (d) 10 moles of H_2O

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6. Which of the following has maximum number of atoms?
(a) 18g of H_2O (b) 18g of O_2
(c) 18g of CO_2 (d) 18 g of CH_4

7. Which of the following contains maximum number of molecules?
(a) 1 g CO_2 (b) 1 g N_2
(c) 1 g H_2 (d) 1 g CH_4

8. Mass of one atom of oxygen is
(a) $\frac{16}{6.023\times10^{23}}$ g (b) $\frac{32}{6.023\times10^{23}}$ g
(c) $\frac{1}{6.023\times10^{23}}$ g (d) 8u

9. 3.42 g of sucrose are dissolved in 18g of water in a beaker. The number of oxygen atoms in the solution are
(a) 6.68×10^{23} (b) 6.09×10^{22}
(c) 6.022×10^{23} (d) 6.022×10^{21}

10. A change in the physical state can be brought about
(a) only when energy is given to the system
(b) only when energy is taken out from the system
(c) when energy is either given to, or taken out from the system
(d) without any energy change.

11. Which of the following represents a correct chemical formula? Name it.
(a) CaCl (b) $BiPO_4$
(c) $NaSO_4$ (d) NaS

12. Write the molecular formulae of all the compounds that can be formed by the combination of following ions.
Cu^{2+}, Na^+, Fe^{3+}, Cl^-, $SO^{2-}{}_4$, $PO^{3-}{}_4$

13. Calculate the number of moles of magnesium present in a magnesium ribbon weighing 12g. Molar atomic mass of magnesium is 24g mol^{-1}.

14. State the number of atoms present in each of the following chemical species
(a) CO_3^{2-} (b) PO_4^{3-}
(c) P_2O_5 (d) CO

15. Calcium chloride when dissolved in water dissociates into its ions according to the following equation.
$CaCl_2\,(aq) \rightarrow Ca^{2+}\,(aq) + 2Cl^-\,(aq)$
Calculate the number of ions obtained from $CaCl_2$ when 222 g of it is dissolved in water.

16. The mass of one steel screw is 4.11g. Find the mass of one mole of these steel screws. Compare this value with the mass of the Earth (5.98×10^{24} kg). Which one of the two is heavier and by how many times?

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EXERCISE-3 Foundation Builder

Multiple Choice Questions (MCQs)

Directions : *This section contains multiple choice questions. Each question has four choices (a), (b), (c) and (d) out of which ONLY ONE is correct.*

1. The formation of SO_2 and SO_3 explain
(a) the law of conservation of mass
(b) the law of multiple proportions
(c) the law of definite properties
(d) Boyle's law

2. The formula of a chloride of a metal M is MCl3, the formula of the phosphate of metal M will be
(a) MPO_4 (b) M_2PO_4
(c) M_3PO_4 (d) $M_2(PO_4)_3$

3. One mole of a gas occupies a volume of 22.4 L. This is derived from
(a) Berzelius' hypothesis (b) Gay-Lussac's law
(c) Avogadro's law (d) Dalton's law

4. ^{35}Cl and ^{37}Cl are the two isotopes of chlorine, in the ratio 3 : 1 respectively. If the isotope ratio is reversed, the average atomic mass of chlorine will be –
(a) 35.0 u (b) 35.5 u
(c) 36.0 u (d) 36.5 u

5. The metal (M) forms an oxide, M_2O_3. The formula of its nitride will be
(a) M_2N_3 (b) MN (c) M_2N (d) M_3N_2

6. A mixture of gases contains H_2 and O_2 gases in the ratio of 1 : 4 (w/w). What is the molar ratio of the two gases in the mixture ? **[JSTSE]**
(a) 4 : 1 (b) 16 : 1
(c) 2 : 1 (d) 1 : 4

7. The number of water molecules is maximum in : **[JSTSE]**
(a) 18 molecules of water
(b) 1.8 gram of water
(c) 18 gram of water
(d) 18 moles of water

8. If Avogadro number N_A, is changed from $6.022 \times 10^{23}\ mol^{-1}$ to $6.022 \times 10^{20}\ mol^{-1}$ this would change : **[JSTSE]**

(a) the definition of mass in units of grams
(b) the mass of one mole of carbon
(c) the ratio of chemical species to each other in a balanced equation.
(d) the ratio of elements to each other in a compound

9. An unknown chlorohydrocarbon has 3.55% of chlorine. If each molecule of the hydrocarbon has one chlorine atom only, chlorine atoms present in 1g of chlorohydrocarbon are: **[NTSE]**

(Atomic wt. of Cl = 35.5u; Avogadro constant = $6.0^{23} \times 10^{23}\ mol^{-1}$)

(a) 6.023×10^9 (b) 6.023×10^{23}
(c) 6.023×10^{21} (d) 6.023×10^{20}

10. In which case is number of molecules of water maximum? **[JSTSE]**

(a) 18 mL of water
(b) 0.18 g of water
(c) 10^{-3} mol of water
(d) 0.00224 L of water vapours at 1 atm and 273 K

11. The percentage composition of carbon by mole in methane is : **[NTSE]**

(a) 75% (b) 80% (c) 25% (d) 20%

12. Which of the following contains the same number of oxygen atoms? **[Olympiad]**

A. 1 g of O atoms
B. 1 g of O_2
C. 1 g of Ozone O_3

(a) A and B only (b) B and C only
(c) C and A only (d) A, B and C

13. Arrange the following atoms in the order of decreasing number of moles. **[Olympiad]**

I. 28 g of He
II. 46 g of Na
III. 60 g of Ca

(a) I > II > III (b) III > II > I
(c) I > III > II (d) III > I > II

14. Which of the following contains the greatest number of atoms? **[Olympiad]**

(a) 0.3 mol of nitrogen gas
(b) 0.5 mol of oxygen gas
(c) 0.4 mol of ozone gas
(d) 0.2 mol of carbon dioxide gas

15. Which of the following solutions has the highest mass by volume percentage? **[Olympiad]**

(a) 10 g of sugar in 50 mL solution
(b) 25 g of potassium chloride in 100 mL solution
(c) 30 g of magnesium sulphate in 50 mL solution
(d) 60 g of sodium chloride in 200 mL solution

16. Consider the following statements : **[Olympiad]**

A. Formula for sulphur dioxide is SO_2
B. Water is an atom.
C. In a pure chemical compound, elements are always present in a definite proportion by mass.

Which of these statement(s) is/are correct?

(a) A and B
(b) B and C
(c) A and C
(d) All are correct

17. Which of the following statements are correct? **[Olympiad]**

I Carbonate ion is a polyatomic ion that carries a charge of –2.
II. Sulphide ion is trivalent and positive.
III. Argon and oxygen are monoatomic and diatomic gases respectively.
IV. Ammonium ion is negatively charged and carries a single negative charge only.

(a) II and IV only (b) I and III only
(c) I and II only (d) III and IV only

18. Consider the following statements :

A. 22gm. of CO_2 consist of 1 mole.
B. Number of molecules in 4 gm of oxygen is 10^{22}.
C. Mass of 1 mole of a substance is called its formula mass. **[Olympiad]**

Which of these statement(s) is/are correct?

(a) A and B (b) B and C
(c) A and C (d) All are incorrect

19. Molecular formula of chloride of a metal 'M' is MCl_2. Molecular formula of oxide of 'M' will be **[NTSE]**

(a) MO (b) M_2O
(c) MO_2 (d) M_2O_2

20. If the aluminium salt of anion 'X' is Al_2X_3 the formula of magnesium salt of 'X' will be: **[NTSE]**

(a) Mg_2X (b) MgX_2
(c) MgX (d) Mg_2X_3

21. Formula of aluminium carbonate is: **[NTSE]**

(a) $Al_2(CO_3)_3$ (b) Al_2CO_3
(c) Al_2HCO_3 (d) $AlCO_3$

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22. Number of molecules present in 14 g of N_2 molecules is [NTSE]
(a) 6.022×10^{23} (b) 3.011×10^{23}
(c) 1.51×10^{23} (d) 3.011×10^{22}

23. A mixture of non- reacting gasses contains hydrogen and oxygen gases in the mass ratio of 1 : 4 respectively. What will be the molar ratio of the above two gases in the mixture? [NTSE]
(a) 16 : 1 (b) 1 : 4 (c) 4 : 1 (d) 1 : 6

More than One Option Correct

DIRECTIONS : *This section contains multiple choice questions. Each question has four choices (a), (b), (c) and (d) out of which ONE OR MORE may be correct.*

1. Which of the following has same mass?
(a) 4g of He
(b) 6.023×10^{23} atoms of He
(c) 1 atom of He
(d) 1 mole atoms of He

2. Which of the following is/are the best example of law of conservation of mass?
(a) 12g of carbon combines with 32g of oxygen to form 44g of CO_2.
(b) When 12g of carbon is heated in a vacuum there is no change in mass.
(c) A sample of air increases in volume when heated at constant pressure but its mass remains unaltered.
(d) 2 g of hydrogen combines with 16 g of oxygen to form 18 g of water.

3. Which one of the following pairs of gases contains the same number of molecules?
(a) 16 g of O_2 and 14 g of N_2
(b) 8 g of O_2 and 22 g of CO_2
(c) 28 g of N_2 and 22 g of CO_2
(d) 8 g of O_2 and 7 g of N_2

4. Which of the following represents a polyatomic ion?
(a) Sulphite (b) Chloride
(c) Sulphate (d) Phosphate

5. Which of the following statements is /are correct?
(a) An atom is the smallest particle of matter according to Dalton's theory.
(b) An atom is the smallest particle of an element.
(c) An atom is the smallest indivisible particle of an element that can take part in a chemical change.
(d) An atom is the radioactive emission.

6. Which of the following symbols does not represent an element?
(a) CO (b) Ar
(c) K (d) NO

7. Which of the following element(s) has a symbol having two letters?
(a) Tin (b) Uranium
(c) Carbon (d) Aluminium

8. Which of the following is / are not a correct symbol for an element(s)?
(a) Ng (b) Fi
(c) Bk (d) Zc

9. Which of the following set of compounds illustrates the law of reciprocal proportions?
(a) HCl, HBr, HI
(b) CO_2, CH_4, H_2O
(c) NH_3, N_2O_3, H_2O
(d) NH_3, NCl_3, N_2O_3

10. Which of the following contains the same number of molecules?
(a) 1g of O_2, 2 g of SO_2
(b) 1g of CO_2, 1g of N_2O
(c) 112 mL of O_2 at STP, 224 mL of He at 0.5 atm and 273K
(d) 1g of oxygen, 1g of ozone

11. SO_2 gas is slowly passed through an aqueous suspension containing 12 g $CaSO_3$ till the milkiness just disappears. What amount of SO_2 would be required ?
(a) 12.8 g (b) 6.4 g
(c) 0.1 mo (d) 0.2 mo

12. 8g of O_2 has the same number of molecules as
(a) 7g CO (b) 14g N_2
(c) 11g CO_2 (d) 16g SO_2

13. A vessel contains 4.4 g of CO_2. It means that it contains
(a) 0.1 mole of CO_2
(b) 6.02×10^{22} molecules of CO_2
(c) 8.8g atoms of oxygen
(d) 1120 mL of CO_2 S.T.P.

14. Which of the following pairs of substances illustrate the law of multiple proportions?
(a) CO and CO_2
(b) H_2O and D_2O
(c) SO_2 and SO_3
(d) MgO and $Mg(OH)_2$

15. Which of the following pairs of compound illustrate the law of multiple proportions.
    (a) CO and CO_2
    (b) NH_3 and NCl_3
    (c) H_2O and H_2O_2
    (d) CO_2 and CH_4

Assertion & Reason Questions

DIRECTIONS : *Each of these questions contains an Assertion followed by Reason. Read them carefully and answer the question on the basis of following options. You have to select the one that best describes the two statements.*

(a) If both Assertion and Reason are correct and Reason is the correct explanation of Assertion.
(b) If both Assertion and Reason are correct, but Reason is not the correct explanation of Assertion.
(c) If Assertion is correct but Reason is incorrect.
(d) If Assertion is incorrect but Reason is correct.

1. **Assertion :** Molecular weight of oxygen is 16.
   Reason : Atomic weight of oxygen is 16.
2. **Assertion :** Pure water obtained from different sources such as, river, well, spring, sea etc. always contains hydrogen and oxygen combined in the ratio 1 : 8 by mass.
   Reason : A chemical compound always contains elements combined together in same proportion by mass, it was discovered by French chemist, Joseph Proust (1799).
3. **Assertion :** 1 amu equals to 1.66×10^{-24} g
   Reason : 1.66×10^{-24} g is equal to $\frac{1}{12}$ the of mass of a C–12 atom.
4. **Assertion :** Volume of a gas is inversely proportional to the number of moles of gas.
   Reason : The ratio by volume of gaseous reactants and products is in agreement with their mole ratio.
5. **Assertion:** The empirical mass of ethene is half of its molecular mass.
   Reason: The empirical formula represents the simplest whole number ratio of various atoms present in a compound.

Passage/Case Based Questions

DIRECTIONS : *Study the given paragraph and answer the following questions.*

An empirical formula gives the smallest whole-number ratio of atoms of each element in a compound. NO_2 is an empirical formula showing that nitrogen and oxygen atoms are present in a 1 : 2 ratio. N_2O_4 is not an empirical formula. It shows a 2 : 4 ratio of nitrogen atoms to oxygen atoms. The 2 : 4 ratio has a common divisor, 2, and can therefore be reduced to 1 : 2, just as you reduce a fraction to lowest terms. Therefore, the empirical formula of N_2O_4 is NO_2.

Not all empirical formulas represent real compounds. For example, the molecular formula for hydrogen peroxide is H_2O_2, which gives HO as its empirical formula. There is no known stable substance with the formula HO.

The distinction between empirical and molecular formulas may be summarised.

If the subscripts of a formula can be divided evenly by a whole number, it is a molecular formula. If the subscripts of a formula cannot be reduced further, it is an empirical formula. It may also be a molecular formula.

1. What is the empirical formula of the compound whose molecular formula is P_4O_{10}?
   (a) PO (b) PO_2
   (c) P_2O_5 (d) P_8O_{20}
2. A compound is 86% carbon and 14% hydrogen by mass. What is the empirical formula for this compound?
   (a) CH (b) CH_2
   (c) CH_3 (d) CH_4
3. What is the empirical formula of a compound that contains 85% Ag and 15% F by mass?
   (a) AgF (b) Ag_2F
   (c) AgF_2 (d) Ag_2F_2

Integer/Numerical Value Type Questions

DIRECTIONS : *Following are integer/numerical value type questions. Each question, when worked out will result in one integer/numerical value.*

1. Calculate the number of moles of oxygen molecules in 6.02×10^{24} CO molecules.
2. If the mass (kg) of a molecule of water is $x \times 10^{-26}$. Find the value of x.
3. A compound was found to contain 55.2 percent xenon and 44.8 percent chlorine. If the the empirical formula of the compound is $XeCl_n$. Calculate the value of n.
4. A 10g sample of a compound contains 4.00g C, 0.667g H, and 5.33g O. if the the ratio between empirical and molecular formulas is n:1, find the value of n. The MW is 180 amu.

SOLUTIONS

Brief Explanations of Selected Questions

Let's Revise Through FIB & T/F

1. masses	**2.** True
3. False	**4.** symbol
5. polyatomic	**6.** chemical formula
7. simple ions	**8.** metal and non-metals
9. True	**10.** molecular formula
11. equal	**12.** False
13. True	

EXERCISE-1 **Master Board**

Multiple Choice Questions (MCQs)

1. **(d)** The atomic theory was proposed by John Dalton.

2. **(d)** Neon and helium are inert gases.

3. **(d)** 8 : 1

4. **(a)** The number of electrons in A^{3+} = 10

The number of electrons in A = 10 + 3 = 13

5. **(b)** As the number of electrons are more than the number of protons. Therefore X must be anion.

6. **(b)** **7.** **(b)** **8.** **(a)**

9. **(b)** 1 atom of Cu + 1 atom of sulphur + 9 atoms of oxygen + 10 atoms of hydrogen. Total number of atoms in compound is 21.

10. **(a)** Constant proportions. According to this a pure chemical compound always contains same elements combined together in the same definite proportion of weight, whatever may be its source.

11. **(d)** Molar mass of aspartame

$=14 \times 12 + 18 \times 1 + 14 \times 2 + 16 \times 5 = 294$

Q Mass of 6.023×10^{23} molecules = 294 g

$\therefore$ Mass of 1 molecule $= \dfrac{294}{6.023\times10^{23}}$

$= 4.89 \times 10^{-22}$ g

12. **(c)** **13.** **(c)** **14.** **(a)** **15.** **(a)**

16. **(a)** **17.** **(b)**

Assertion & Reason Questions

1. **(c)** An atom is electrically neutral.

Atoms necessarily contain equal number of protons and electrons, but not neutrons.

2. **(c)** Isobars have the same atomic mass (sum of protons and neutrons) but different atomic number.

3. **(b)** Atoms can be created or destroyed. Atom is made up of subatomic particles.

4. **(d)** The atomicity of ozone is 3.

5. **(a)** Atomic mass unit

$= \dfrac{1}{12}$ mass of a carbon – 12 atom.

Passage/Case Based Questions

1. (i) Calcium carbonate ($CaCO_3$) in the form of marble (ii) Carbon dioxide (CO_2) (iii) Calcium oxide (CaO)

2. 50

3. Total mass of B and C (50 g) is equal to the mass of A taken (50 g)

4. The mass of products is equal to the mass of reactant

5. Law of conservation of mass.

Very Short Answer Questions

1. It is the smallest repeating formula in the structure of ionic compound.

2. Phosphorus trichloride and sulfur dioxide.

3. 1 : 2

4. 2 mol of Fe combine with 3 mol of oxygen atom. Therefore 1 mol of Fe combine with $\dfrac{3}{2}$ mol of oxygen atom.

1 mol of Fe = 56 g

$\dfrac{3}{2}$ mol of O atom = 24 g of O atom

24 g of O atom = 56 g of Fe

25.6 g of O atom $= \dfrac{56}{24}\times25.6$ g of Fe = 59.4 g of Fe

Short Answer Questions

1. 44 g CO_2 = 1 mole

$\therefore$ 22 g CO_2 = 0.5 mol

2. The molecular mass of a substance (element or compound) is the number of times the molecule of the substance is heavier than 1/12th the mass of an atom of carbon-12 isotope.

Molecular weight is calculated by adding the atomic weights of all the constituent atoms present in a molecule. For example :

Molecular weight of a molecule of hydrogen (H_2)

= 2 × atomic weight of hydrogen = 2 × 1 = 2 amu

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3. (a) NaF (b) Na_2O
(c) MgF_2 (d) Al_4C_3

4. (a) Al_2S_3 (b) SrF_2
(c) TiO_2 (d) $CaBr_2$

5. Carbon-12 atom has 6 protons and 6 neutrons in its nucleus so that its mass number is 12, carbon-12 atom has been assigned an atomic mass of exactly 12 atomic mass units. Also carbon-12 has highest isotopic abundance, which make it ideal for relative atomic mass calculations.

6. An atom consists of electrons, protons and neutrons. Electrons are negatively charged, protons are positively charged while neutrons are neutral. Every atom contains equal number of electrons and protons which balances the charges in the atom and makes an atom electrically neutral.

7. Ionic compounds are made up of ions i.e., positively charged ions and negatively charged ions. They are held together by strong electrostatics force of attraction. Ionic compounds consists of an equal number of positive ions and negative ions, so the overall, charge on an ionic compound is zero.

8. Water molecule consists of three interconnected particles or atoms. Each water molecule can decompose into two hydrogen atoms and one oxygen atom. Molecules of a substance can be destroyed by a chemical reaction, whereas atoms remain unchanged. For this reason, we cannot say 'an atom of water' but we say 'a molecule of water'.

9. The amount of oxygen per gram of nitrogen in NO_2 should be exactly twice as that of NO, as required by the formulas of the two substances. Therefore, 2.286 g oxygen would combine with 1.000 g nitrogen.

10. Elements are made up of atoms. All the atoms of a given elements are identical in mass, size therefore they exhibit same chemical properties. We can say that different elements show specific chemical properties.

11. The size of an atom is very very small. The size of an atom is indicated by its radius called as atomic radius. Atomic radius is measured in nano metres. Due to very small size of an atom its cannot be seen even with the high resolution microscope.

12. For a chemical reaction taking place in a closed container, the total mass of products is equal to the total mass of reactants. There is no change in mass during the reaction. Therefore law of conservation of mass is said to be satisfied.

13. In order to solve this problem, we should know the mass of 1 mole of water. This can be obtained by using the given values of the atomic masses of hydrogen and oxygen as follows :

1 mole of water (H_2O) = Molecular mass of H_2O in grams

= Mass of 2H atoms + Mass of O atom

= 2 × 1 + 16 = 2 + 16 = 18 grams

Thus, the mass of 1 mole of water = 18 g

So, Mass of 0.5 mole of water = 18 × 0.5 g = 9 g

14. According to law of constant proportions, a chemical compound always consists of the same elements. Combined together in the same proportion by mass. Water is a compound, it always consists of the same two elements, hydrogen and oxygen combined in 1 : 8 proportion by mass.

15. Molar mass of $NH_3 = 14 \times 1 + 1 \times 3 = 17$ g/mol

Percentage of nitrogen

$$= \frac{\text{Molar mass of N in } NH_3}{\text{Molar mass of } NH_3} \times 100$$

$$= \frac{14}{17} \times 100 = 82.35\%$$

Percentage of hydrogen $= \frac{3}{17} \times 100 = 17.65\%$

16. We first determine the number of grams of each element that are present in one mole of sample :

2 mol of N × 14.01 g/mol = 28.02 g of N

3 mol of O × 16.00 g/mol = 48.00 g of O

The percentages by mass are then obtained by using the formula mass of the compound (76.02 g) :

$$\% \text{ N} = \frac{28.02}{76.02} \times 100 = 36.86\% \text{ of N}$$

$$\% \text{ O} = \frac{48.00}{76.02} \times 100 = 63.14\% \text{ of O}$$

17. Mass of the reactants = 10 g

Mass of the products = 4.4 + 5.6 g = 10 g

Since the mass of the reactants is equal to the mass of the products, the observations are in agreement with the law of conservation of mass.

Long Answer Questions

1. (a) $NaHCO_3$ = 1 Na + 1H + 1C + 3 O
= (22.99) + (1.01) + (12.01) + (3 × 16.00)
= 84.0 g/mole

(b) $K_2Cr_2O_7$ = 2K + 2Cr + 7O
= (2 × 39.10) + (2 × 52.00) + (7 × 16.00)
= 294.2 g/mol

(c) $(NH_4)_2CO_3$ = 2N + 8H + C + 3O
= (2×14.01) + (8×1.01) + (12.01) + (3×16.00)
= 96.1 g/mol

(d) $Al_2(SO_4)_3 = 2Al + 3S + 12O$
$= (2 \times 26.98) + (3 \times 32.07) + (12 \times 16.00)$
$= 342.2$ g/mol

(e) $CuSO_4.5H_2O = 1Cu + 1S + 9O + 10H$
$= 63.55 + 32.07 + (9 \times 16.00) + (10 \times 1.01)$
$= 249.7$ g/mol

2. (I) (a) Cu_2O (b) K_2O_2 (c) Hg_2Br_2
(d) $Fe_2(CO_3)_3$ (e) NaBrO

(II) (a) Bromic acid (b) Hydrobromic acid
(c) Phosphoric acid (d) Hypochlorous acid
(e) Iodic acid

3. (I) The average mass of a molecule of carbon dioxide would be equal to the sum of the atomic weights of the three atoms in a CO_2 molecule.

Mass of a single CO_2 molecule :
1 C atom = 1 (12.011 amu) = 12.011 amu
2 O atom = 2 (15.999 amu) = 31.998 amu
44.009 amu

The mass of a mole of carbon dioxide would be 44.009 grams.

The average mass of a molecule of glucose is equal to the sum of the atomic weights of the 24 atoms in a $C_6H_{12}O_6$ molecule.

Mass of a single $C_6H_{12}O_6$ molecule :
6 C atom = 6 (12.011 amu) = 72.066 amu
12 H atoms = 12 (1.0079 amu) = 12.095 amu
6 O atoms = 6 (15.999 amu) = 95.994 amu
180.155 amu

The molecular weight of the compound is therefore 180.155 g/mol.

(II) 1 mol of CO_2 = 44.009 g
$\therefore$ Number of moles in 0.5 g of CO_2

$$= 0.500\text{g of } CO_2 \times \frac{1 \text{ mol of } CO_2}{44.009\text{g of } CO_2}$$

$= 1.14 \times 10^{-2}$ moles of CO_2
1 mole of $CO_2 = 6.023 \times 10^{23}$ molecules
1.14×10^{-2} mol of CO_2
$= 6.023 \times 10^{23} \times 1.14 \times 10^{-2}$ CO_2 molecules
$= 6.86 \times 10^{21}$ CO_2 molecules

We can now use the chemical formula for carbon dioxide to determine the number of carbon atoms in the sample. The formula suggests that there is a single carbon atom for each CO_2 molecule.

$\therefore$ 6.86×10^{21} CO_2 molecules
$= 6.86 \times 10^{21}$ C atoms

4. (i)

	Atom		Molecule
1.	It is the smallest unit of an element	1.	It is the smallest unit of an element or a compound.
2.	It may or may not be capable of free existence but it is the smallest unit that takes part in a chemical reaction.	2.	It is capable of free existence.

(ii)

	Atom		Ion
1.	It is neutral and carries no charge.	1.	It is a charged particle and carries a positive or negative charge.
2.	The number of electrons in an atom is equal to its atomic number.	2.	In cation (ion with positive charge) the number of electrons are less than that present in corresponding atom. In an anion (ion with negative charge) the number of electrons are more than that present in corresponding atom.
3.	It is quite reactive.	3.	It is quite stable.

HOTS Questions

1. (a) $CrCl_3$ and $CrCl_2$, Cr_2O_3 and CrO
(b) CuCl, $CuCl_2$, Cu_2O and CuO

2. Sodium arsenate

3. (a) 6 : 11 (b) 2 : 1
(c) 12 : 11 (d) 2 : 1

4. In first sample of ascorbic acid
1.50 g of carbon combines with 2.00 g of oxygen.

1 g of carbon combines with $\frac{2}{1.5}$g of oxygen.

In second sample

6.35 g of carbon needs $\frac{2 \times 6.35}{1.5}$g = 8.47 gof oxygen

Sample I	Sample II
2 : 1.5	8.47 : 6.35
	2 : 1.5

Thus, the above calculations supports the law of constant composition.

5. (i) Number of oxygen atoms in 16.0g of oxygen
$= 6.023 \times 10^{23}$ atoms
$\therefore$ Weight of one atom of oxygen

$$= \frac{16.0}{6.023 \times 10^{23}} = 2.657 \times 10^{-23} \text{ g}$$

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(ii) G.M.W. of oxygen = 32.0g

Number of O_2 molecules in 32.0 g of O_2

= 6.023×10^{23} molecules

∴ Weight of one molecule of oxygen

$$= \frac{32.0}{6.023 \times 10^{23}} = 5.314 \times 10^{-23} g$$

6. The atomic theory states that the relative number and kinds of atoms in a compound are constant, regardless of the source. Therefore, 1.0g of pure water should always contain the same relative amounts of hydrogen and oxygen, no matter from where or how the sample is obtained.

7. Its mass is 200 amu (or u).

8. Molecular mass of urea (H_2NCONH_2)

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

= 2 + 14 + 12 + 16 + 14 + 2 = 60u

Mass of nitrogen in 1 molecule of urea

= 1 × 14 + 1 × 14 = 14 + 14 = 28u

$$\therefore \text{ \% age of nitrogen in urea } = \frac{28}{60} \times 100 = 46.6\%$$

9. 1.375 g of pure cupric oxide gave 1.098 g of Cu hence,

$$\text{Percentage of Cu in the oxide} = \frac{1.098}{1.375} \times 100 = 79.85\%$$

In another experiment, 1.179 gm of pure copper gave 1.476 g of the oxide hence,

$$\text{Percentage of Cu in the oxide} = \frac{1.179}{1.476} \times 100 = 79.87\%$$

Since, both the oxides have almost the same percentage of Cu and oxygen, therefore result is obeying the law of constant composition.

10. (a) $ZnCO_3$, ZnO, CO_2

(b) HF, SiO_2, SiF_4, H_2O

(c) SO_2, H_2O, H_2SO_3

(d) H_3P (or PH_3)

(e) $HClO_4$, Cd, Cd $(ClO_4)_2$

EXERCISE-2 **NCERT Questions**

Intext-Questions

1. Here, the mass of reactants = (5.3 + 6.0) g = 11.3 g

Mass of products = (8.2 + 2.2 + 0.9) g = 11.3 g

Since, the mass of reactants is equal to mass of products. So, these observations are in agreement with the law of conservation of mass.

2. Hydrogen + Oxygen = Water

The ratio in which hydrogen and oxygen combine is 1 : 8 i.e., 1 g of hydrogen requires 8 g of oxygen for complete reaction.

Thus, mass of oxygen required by 1 g of hydrogen = 8 g

∴ Mass of oxygen required by 3 g of hydrogen

$$= \frac{8}{1} \times 3 \text{ g} = 24 \text{ g}$$

3. The law of conservation of mass is the result of the following postulate of Dalton's atomic theory.

Atoms are indivisible and they can neither be created nor be destroyed.

4. The following postulates of Dalton's atomic theory can explain the law of definite proportions.

Atoms of various elements combine in a simple whole number but fixed ratio to form compound atoms (molecules).

5. One atomic mass unit (a.m.u. or 'u') is exactly equal to $\left(\frac{1}{12}\text{th}\right)$ of the mass of one atom of carbon – 12

$$\boxed{1 \text{ u} = \frac{1}{12}\text{th of mass of 1 atom of C-12}}$$

6. We cannot see an atom with naked eye because of its very small size. The radius of an atom is in the order of 10^{-10} m and such a small object cannot be seen with a naked eye.

7. (i) Sodium Oxide

Na^+ O^{2-}

Na_2O_1 or Na_2O (neglect the subscript 1)

(ii) Aluminium Chloride

Al^{3+} Cl^-

Al_1Cl_3 or $AlCl_3$

(iii) Sodium Sulphide

Na^+ S^{2-}

Na_2S_1 or Na_2S

(iv) Magnesium Hydroxide

Mg^{2+} $(OH)^{-1}$

$Mg_1(OH)_2$ or Mg $(OH)_2$

8. (i) Aluminium sulphate

(ii) Calcium chloride

(iii) Potassium sulphate

(iv) Potassium nitrate

(v) Calcium carbonate

9. The chemical formula of a compound represents the composition of a molecule of the compound in terms of symbols of the elements present in it.

10. (i) H_2S molecule consists of 2 atom of hydrogen and one atom of sulphur. So the total number of atoms present in a molecule of H_2S is 3 (2 + 1 = 3). Thus, the atomicity of H_2S is 3.

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(ii) PO_4^{3-} ion contains 1 atom of phosphorus and 4 atoms of oxygen. So, the total number of atoms in it is 5 (1 + 4 = 5).

11. Molecular mass of $H_2 = 2 \times 1 = 2u$ (or 2 amu)

Molecular mass of $O_2 = 2 \times 16 = 32u$

Molecular mass of $Cl_2 = 2 \times 35.5 = 71u$

Molecular mass of $CO_2 = 12 + 2 \times 16 = 12 + 32 = 44u$

Molecular mass of $CH_4 = 1 \times 12 + 4 \times 1 = 12 + 4 = 16u$

Molecular mass of $C_2H_6 = 2 \times 12 + 6 \times 1 = 24 + 6 = 30u$

Molecular mass of $C_2H_4 = 2 \times 12 + 4 \times 1 = 24 + 4 = 28u$

Molecular mass of $NH_3 = 1 \times 14 + 3 \times 1 = 14 + 3 = 17u$

Molecular mass of CH_3OH

$= 1 \times 12 + 3 \times 1 + 1 \times 16 + 1 \times 1$

$= 12 + 3 + 16 + 1 = 32$ u

12. Formula unit mass of ZnO = 1 × 65 + 1 × 16

= 65 + 16 = 81u

Formula unit mass of $Na_2O = 2 \times 23 + 1 \times 16$

= 46 + 16 = 62u

Formula unit mass of K_2CO_3

= 2 × 39 + 1 × 12 + 3 × 16

= 78 + 12 + 48 = 138u

Text Book Exercise

1. Mass of the compound = 0.24 g

Mass of boron in the compound = 0.096 g

Mass of oxygen in the compound = 0.144 g

Percentage of born (by weight)

$$= \frac{\text{Mass of boron}}{\text{Mass of compound}} \times 100$$

$$= \frac{0.096}{0.24} \times 100 = 40$$

Percentage of oxygen (by weight)

$$= \frac{\text{Mass of oxygen}}{\text{Mass of compound}} \times 100 \quad = \frac{0.144}{0.24} \times 100 = 60$$

2. The reaction between carbon and oxygen to form carbon dioxide can be represented by the following chemical equation.

	C	+	O_2 (2×16)	$\longrightarrow$	CO_2
	12g		32g		12 + 32 = 44g
or	3g		8g		11g

From the above data we find that 3g of carbon combines with 8g of oxygen to form 11 g of carbon dioxide.

This is also the data given in first case.

In second case when 3.0 g of carbon is burnt in 50.0g of oxygen, only 8.0g of oxygen will combine with 3.0 g of carbon to produce 11.0g of CO_2 and the remaining oxygen will remain as such (i.e., unreacted)

Thus, mass of carbon dioxide formed will the 11g

In second case, mass of unreacted oxygen

= (50 – 8.0) = 42.0g

[In this case, carbon is present in smaller amount (limiting reagent) is completely consumed. The answer of this problem is based on laws of constant proportions.

3. Polyatomic ions are the group of atoms that carry either some positive charge (cation) or some negative charge (anion) *e.g.*

$NH_4^+, CO_3^{2-}, NO_3^-, PO_4^{3-}$

4. (a) Magnesium choride

Mg^{+2} $\times$ $C1^{-1}$ $\quad Mg_1Cl_2$ or $MgCl_2$

(b) Calcium oxide

Ca^{2+} $\times$ O^{-2} $\quad Ca_2O_2$ or CaO

(c) Copper nitrate

Cu^{2+} $\times$ NO_3^{-1} $\quad Cu_1(NO_3)_2$ or $Cu(NO_3)_2$

(d) Ammonium chloride

NH_4^{+1} $\quad Cl^{-1}$ $\quad NH_4Cl$

(e) Calcium carbonate

Ca^{2+} $\quad CO_3^{2-}$ $\quad CaCO_3$

(f) Zinc sulphate

Zn^{2+} $\quad SO_4^{2-}$ $\quad ZnSO_4$

5. (a) Quick lime is the commercial name of calcium oxide (CaO). The elements present in it are calcium and oxygen.

(b) Hydrogen bromide has the chemical formula HBr and the elements present in it are hydrogen and bromine.

(c) Baking powder is the commercial name of sodium hydrogen carbonate ($NaHCO_3$). The elements present in it are sodium, hydrogen, carbon and oxygen.

(d) Potassium sulphate has the formula K_2SO_4. The elements present in it are potassium, sulphur and oxygen.

Exemplar Questions

1. **(d)** (ii) and (iv) points correctly represent 360 g of water.

(ii) From point,

$\because$ 1 mole of water = molar mass of water = 18g

$\therefore$ 20 moles of water = 18 g × 20 = 360 g

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(iv) From point,

$\because$ 6.022×10^{23} molecules of water = 1 mole = 18g of water

$\therefore$ 1.2044×10^{25} molecules of water

$$= \frac{18g \times 1.2044 \times 10^{25}}{6.022 \times 10^{23}} = 360g$$

Therefore, points (ii) and (iv) represent 360 g of water.

Students can check other points also, as

(i) 1 mole of water = molar mass of water = 18g of water.

$\therefore$ 2 mole of water = $18 \times 2 = 36$ g

(iii) 6.022×10^{23} molecules of water = 18g water.

2. (d) This statement is not be true.

The correct statement is as the molecules and ions aggregate together in large number to form the matter. We cannot see the individual molecules/ ions with our eyes, only we can see the various substances which are a big collection of molecules / ions. So, option (d) is incorrect.

Atoms of most of the elements are chemically very reactive and do not exist in the free state. The atoms of only noble gases are chemically unreactive and exist in free state.

Atoms usually exist in the two forms

(i) molecules and (ii) ions

Therefore, atoms are the basic unit from which molecules and ions are formed. When atoms form molecules or ions, they become stable because they acquire the stable noble gas electron arrangement. Therefore, they are neutral in nature.

3. (b) Nitrogen molecule is diatomic molecule, therefore, it exists as N_2 molecules.

4. (d) The chemical symbol for sodium is derived from its Latin name 'Natrium'. In a 'two letter' symbol, the first letter is the 'capital letter' but the second letter is the 'small letter'. Therefore, its symbol is 'Na'.

5. (c) 2 moles of $CaCO_3$ would weigh the highest.

(a) $\therefore$ Mass of 1 mole of sucrose ($C_{12}H_{22}O_{11}$) = $(12 \times 12) + (1 \times 22) + (16 \times 11) = 342$ g

$\therefore$ Mass of 0.2 mole of sucrose = $342 \times 0.2 = 68.4$g

(b) $\therefore$ Mass of 1 mole of CO_2 = $12 + (16 \times 2) = 44$ g

$\therefore$ Mass of 2 moles of CO_2 = $44 \times 2 = 88$g

(c) $\therefore$ Mass of 1 mole of $CaCO_3$ = $40 + 12 + (16 \times 3) = 100$g

$\therefore$ Mass of 2 moles of $CaCO_3$ = $2 \times 100 = 200$ g

(d) $\therefore$ Mass of 1 mole of H_2O = $H_2O = 2 + 16 = 18$g

$\therefore$ Mass of 10 moles of H_2O = $18 \times 10 = 180$ g.

Therefore, mass of 2 moles of $CaCO_3$ is the highest, i.e., 200g.

6. (d) Option (d) is the correct answer, i.e., 18g of CH_4 has maximum number of atoms.

As,

(a) Number of atoms in 18 g of H_2O

$$= \frac{18}{18} \times 6.022 \times 10^{23} \times 3$$

$= 18.066 \times 10^{23} = 1.8066 \times 10^{24}$

(b) Number of atoms in 18g of O_2

$$= \frac{18}{32} \times 6.022 \times 10^{23} \times 2$$

$= 3.387 \times 10^{23} \times 2 = 6.774 \times 10^{23}$

(c) Number of atoms in 18 g of CO_2

$$= \frac{18}{44} \times 6.022 \times 10^{23} \times 3 = 7.390 \times 10^{23}$$

(d) Number of atoms in 18g of CH_4

$$= \frac{18}{16} \times 6.022 \times 10^{23} \times 5 = 3.387 \times 10^{24}$$

Thus, 18g of CH_4 contains the maximum number of atoms.

7. (c) Option (c) the correct answer, i.e., 1g H_2 contains maximum number of molecules.

(a) Number of molecules in 44g CO_2 = 6.022×10^{23} (molar mass of CO_2 = 44g)

$\therefore$ Number of molecules in 1g CO_2

$$= \frac{6.022 \times 10^{23}}{44} = 1.37 \times 10^{22}$$

(b) Number of molecules in 28g N_2 = 6.022×10^{23} (molar mass of N_2 = 28 g)

$\therefore$ Number of molecules in 1 g N_2

$$= \frac{6.022 \times 10^{23}}{28} = 2.15 \times 10^{22}$$

(c) Number of molecules in 2 gH_2 = 6.022×10^{23} (molar mass of H_2 = 2g)

$\therefore$ Number of molecules in 1 g H_2

$$= \frac{6.022 \times 10^{23}}{2} = 3.011 \times 10^{23}$$

(d) Number of molecules in 16g CH_4 = 6.022×10^{23} (molar mass of CH_4 = 16g)

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$\therefore$ Number of molecules in 1 g CH_4

$$= \frac{6.022 \times 10^{23}}{16} = 3.76 \times 10^{22}$$

Thus, 1g H_2 contains the maximum number of molecules.

8. **(a)** $\because$ Mass of 6.023×10^{23} atoms of oxygen = gram atomic mass of oxygen

Mass of 6.023×10^{23} atoms of oxygen = 16g

$\therefore$ Mass of 1 atom of oxygen $\frac{16}{6.023 \times 10^{23}}$ g

9. **(a)** **Step 1 :** Molar mass of sucrose, $C_{12}H_{22}O_{11}$

$= 12 \times 12 + 1 \times 22 + 16 \times 11 = 342g$

or 342 g = 1 mole of sucrose

342 g = 0.01 mole of sucrose

$\because$ 1 mole of sucrose ($C_{12}H_{22}O_{11}$) contain oxygen-atoms = $11 \times 6.022 \times 10^{23}$ atoms

$\therefore$ 0.01 mole of sucrose will contains oxygen-atoms

$= 0.01 \times 11 \times 6.022 \times 10^{23}$ atoms $= 6.6242 \times 10^{22}$

Step 2 : 18g of water(H_2O) = 1 mole of water

1 mole of water (H_2O) contains oxygen-atoms $= 6.022 \times 10^{23}$ atoms

Step 3 : By adding the number of oxygen-atoms present in 3.42g of sucrose and 18 g of water, we get

$6.022 \times 10^{23} + 6.6242 \times 10^{22}$

$= 10^{22} (60.22 + 6.6242) = 66.844 \times 10^{22}$

$= 6.68 \times 10^{23}$ atoms

10. **(c)** A change in the physical state can be brought about when energy is either given to, or taken out from the system. It is because, energy change helps in changing the magnitude of attraction forces between the particles, thus helps in changing the physical states (i.e., solid, liquid, gas) of matter.

11. **(b)** $BiPO_4$ — Both ions are trivalent

Bismuth phosphate

12. $CuCl_2$ / $CuSO_4$ / $Cu_3(PO_4)_2$

$NaCl$ / Na_2SO_4 / Na_3PO_4

$FeCl_3$ / $Fe_2(SO_4)_3$ / $FePO_4$

13. Number of moles $= \frac{12}{24} = 0.5$ mol

14. (a) 4 (b) 5 (c) 7 (d) 2

15. 1 mole of calcium chloride = 111g

$\therefore$ 222g of $CaCl_2$ is equivalent to 2 moles of $CaCl_2$

Since 1 formula unit $CaCl_2$ gives 3 ions, therefore, 1 mol of $CaCl_2$ will give 3 moles of ions

2 moles of $CaCl_2$ would give $3 \times 2 = 6$ moles of ions.

No. of ions = No. of moles of ions × Avogadro number

$= 6 \times 6.022 \times 10^{23} = 36.132 \times 10^{23}$

$= 3.632 \times 10^{24}$ ions

16. One mole of screws weigh $= 2.475 \times 10^{24}$ g

$= 2.475 \times 10^{21}$ kg

$$\frac{\text{Mass of the Earth}}{\text{Mass of 1 mole of screws}} = \frac{5.98 \times 10^{24} \text{kg}}{2.475 \times 10^{21} \text{kg}} = 2.4 \times 10^3$$

Mass of earth is 2.4×10^3 times the mass of screws

The earth is 2400 times heavier than one mole of screws.

EXERCISE-3 Foundation Builder

Multiple Choice Questions (MCQs)

1. **(b)**

2. **(a)** $MCl_3 \rightarrow M^{3+} + 3Cl^-$

$M^{3+} + PO_4^{3-} \rightarrow MPO_4$

3. **(c)**

4. **(d)** The ratio of isotopes ^{35}Cl and ^{37}Cl = 1 : 3

Average atomic mass $= \frac{35 \times 1 + 37 \times 3}{4} = \frac{35 + 111}{4}$

$= \frac{146}{4} = 36.5$ u

5. **(b)** The formula of metal nitride is MN

6. **(a)** Ratio of weight of gases $= w_{H_2} : w_{O_2} = 1 : 4$

Ratio of moles of gases $= n_{H_2} : n_{O_2} = \frac{1}{2} : \frac{4}{32}$

$\therefore$ Molar Ratio $= \frac{1}{2} \times \frac{32}{4} = 4 : 1$

7. **(d)** No. of moles of water

In 1.8 g of H_2O = 0.1 moles

In 18 g of H_2O = 1 moles

1 mole contain 6.022×10^{23} molecules of water therefore maximum number of molecules is in 18 moles of water.

8. **(b)** If 6.022×10^{23} changes to 6.022×10^{20} mol than this would change mass of one mole of carbon.

9. **(d)** Given percentage of chlorine in an hydrocarbon = 3.55%

i.e.,

100 g of chlorohydrocarbon has 3.55 g of chlorine.

1 g of chlorohydrocarbon will have

$\frac{3.55}{100} = 0.0355$ g of chlorine.

Atomic wt. of Cl = 35.5 g/mol

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Number of moles of Cl $= \dfrac{0.0355 \text{ g}}{35.5 \text{ g/mol}} = 0.001 \text{ mol}$

Number of atoms of

Cl = $0.001 \text{ mol} \times 6.0^{23} \times 10^{23} \text{ mol}^{-1}$

$= 6.023 \times 10^{20}$

10. (a)

(a) Mass of water = 18 × 1 = 18 g

Molecules of water = mole × $N_A = \dfrac{18}{18} N_A = N_A$

(b) Molecules of water = mole × N_A

$= \dfrac{0.18}{18} N_A = 10^{-2} N_A$

(c) Molecules of water = mole × $N_A = 10^{-3} N_A$

(d) Moles of water = $\dfrac{0.00224}{22.4} = 10^{-4}$

Molecules of water = mole × $N_A = 10^{-4} N_A$

11. (d) CH_4 has one atom of carbon among 5 atoms (1C + 4H)

$\therefore$ Mole % of C = $\dfrac{1}{5} \times 100 = 20$ %

12. (d) Number of moles 1/16×1 in O, 1/32×2 in O_2 and 1/48×3 in O_3 of oxygen atom = $\dfrac{1}{16}$.

13. (a) Number of moles = $\dfrac{\text{Given mass (W)}}{\text{Molar mass M}}$

(I) Number of moles of He

$= \dfrac{28}{4} = 7 \text{ moles } (\because M_{He} = 4$

(II) Number of moles of Na

$= \dfrac{46}{23} = 2 \text{ moles } (\because M_{Na} = 23)$

(III) Number of moles of Ca

$= \dfrac{60}{40} = 1.5 \text{ moles } (\because MC_a = 40$

Hence, decreasing order of no. of moles is

(I) > (II) > (III).

14. (c) Number of particles = Number of moles × Avogadro No.

(a) 0.3 mol of N_2 gas contains

$= 0.3 \times 2 \times 6.023 \times 10^{23}$

i.e., 3.61×10^{23} atoms of N.

(b) 0.5 mol of O_2 gas contains

$= 0.5 \times 2 \times 6.023 \times 10^{23}$

i.e., 6.023×10^{23} atoms of O.

(c) 0.4 mol of O_3 gas contains

$= 0.4 \times 3 \times 6.023 \times 10^{23}$

i.e., 7.23×10^{23} atoms of O.

(d) 0.2 mol of CO_2 gas contains

$= 0.2 \times 3 \times 6.023 \times 10^{23}$

i.e., 3.61×10^{23} atoms of C and O collectively.

Hence, 0.4 mol of ozone gas possesses highest number of atoms.

15. (c) The Mass by volume percentage

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

For the given solutions,

(a) $(m/V)\% = \dfrac{10}{50} \times 100 = 20\%$

(b) $(m/V)\% = \dfrac{25}{100} \times 100 = 25\%$

(c) $(m/V)\% = \dfrac{30}{50} \times 100 = 60\%$

(d) $(m/V)\% = \dfrac{60}{200} \times 100 = 30\%$

Therefore, the solution given in option (c) has the highest mass by volume percentage.

16. (c) Water is a molecule containing hydrogen and oxygen atom. Its formula is H_2O.

17. (a) Sulphide ion is divalent and negative, i.e., S^{2-} Ammonium ion $\left(NH_4^+\right)$ is positively charged and carries a single positive charge only.

Therefore statement (ii) and (iv) are incorrect.

18. (d) 1 mole of CO_2 = 44g

Mass of 1 mole of a substance is called molar mass.

19. (a) Molecular formula of chloride = MCl_2

$\therefore$ Valency of metal M = 2

Molecular formula of its oxide will be

M O

Valency 2 2

Formula M_2O = MO

20. (c) Valency of X = 2, So correct formula is MgX.

21. (a)

22. (b) 14 g N_2 contains 14/28 = 0.5 moles (if molecular weight of N_2 is = 28)

Number of molecules in 0.5 moles of

$N_2 = 0.502 \times 6.022 \times 10^{23} = 3.011 \times 10^{23}$

23. (c) Hydrogen : Oxygen ⇒ 1 : 4 =

$$\frac{n_{H_2}}{n_{O_2}} = \frac{W_{H_2} \times M_{O_2}}{M_{H_2} \times W_{O_2}} = \frac{1 \times 32}{4 \times 2} = \frac{4}{1} \Rightarrow 4:1$$

More than One Option Correct

1. (a, b, d)

2. (a, d)

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3. **(a, d)** 16 g O_2 has no. of moles = $\frac{16}{32} = \frac{1}{2}$ mol

14 g N_2 has no. of moles = $\frac{14}{28} = \frac{1}{2}$ mol

No. of moles are same, so no. of molecules are same.

8 g of O_2 has no. of moles = $\frac{8}{32} = \frac{1}{4}$ mol

7 g of N_2 has no. of moles = $\frac{7}{28} = \frac{1}{4}$ mol

4. **(a, c, d)** Sulphite (SO_3^{2-}), Sulphate (SO_4^{2-}) and Phosphate (PO_4^{3-}) ions are polyatomic as they contain more than one ion.

5. **(a, b, c)**

6. **(a, d)**

7. **(a, d)** Symbol of tin and aluminium are Sn and Al respectively.

8. **(a, b, d)** Bk is Berkelium.

9. **(b, c)**

10. **(a, b)**

11. **(b, c)** $CaSO_3(s) + H_2O(l) + SO_2(g) \rightarrow \underset{\text{(soluble)}}{Ca(HSO_3)_2}$

mol of SO_2 required = mol of $CaSO_3$

mol of $CaSO_3$ = $\frac{12}{120} = 0.1$;

mass of SO_2 = 0.1 × 64 = 6.4 g

12. **(a, c, d)** 8g O_2 = $\frac{8}{32}$ mol = 0.25 mol

(a) 7g CO = $\frac{7}{28}$ = 0.25 mol

(b) 14 g N_2 = $\frac{14}{28}$ = 0.5 mol

(c) 11g CO_2 = $\frac{11}{44}$ = 0.25 mol

(d) 16g SO_2 = $\frac{16}{64}$ = 0.25 mol

Equal moles contain equal number of molecules.

13. **(a, b)** 4.4g of Co_2 = $\frac{4.4}{44}$ mol = 0.1 mol

= 0.1 × 6.02 × 10^{23} molecules

= 6.02 × 10^{22} molecules

14. **(a,c)** Here the amount of oxygen which combines with fixed amount of C and S in their oxides will be in a simple whole number ratio.

15. **(a,c)** Ratio of masses of oxygen in CO and C O_2 = 16 : 32 = 1 : 2

Ratio of masses of oxygen in H_2O and H_2O_2 = 16 : 32 = 1 : 2

Assertion & Reason Questions

1. **(d)** Molecular weight of oxygen is 32.

2. **(a)**

3. **(a)** 6.023 × 10^{23} atom of C–12 = 12g

1 atom of C–12 = $\frac{12}{6.023 \times 10^{23}}$ g

$\frac{1}{12}$ of one atom of C–12 = $\frac{1}{6.023 \times 10^{23}}$ g

[1 amu = $\frac{1}{12^{th}}$ of one atom of C–12]

= 0.166 × 10^{-23} g = 1.66 × 10^{-24} g

4. **(d)** We know that from the reaction $H_2 + Cl_2 \rightarrow 2HCl$ that the ratio of the volume of gaseous reactants and products is in agreement with their molar ratio. The ratio of $H_2 : Cl_2$: HCl volume is 1: 1: 2 which is the same as their molar ratio. Thus volume of gas is directly related to the number of moles. Therefore, the assertion is false but reason is true.

5. **(a)**

Passage/Case Based Questions

1. **(c)** To find its empirical formula reduce P_4O_{10} to the smallest whole number ratio of atoms. P_4O_{10} reduced by 2 equals P_2O_5.

2. **(b)** Moles of C = 86/12 = 7 mol of C; Moles of H = 14 / 1 = 14 mol of H.

Hence the ratio of moles of C and H to each other: 14 / 7 is 2.

In other words there are 2 H moles for every C mole or CH_2. This is the empirical formula.

3. **(a)** The molecular mass of Ag is 108/mol; molecular mass of F is 19/mol.

Moles of Ag = 85/108 = 0.79 mol

Moles of F = 15/19 = 0.79 mol

Or the whole number ratio is 0.79/0.79 = 1

Hence the empirical formulae is AgF

Integer/Numerical Value Type Questions

1. **(5)** 6.02 × 10^{23} molecules of CO =1mole of CO

6.02 × 10^{24} CO molecules = 10 moles CO

= 10 mol atoms of O = 5 mol molecules of O_2

2. **(3)** Mass of one molecule of water

$= \frac{18}{6.023 \times 10^{23}} = 3 \times 10^{-23} g = 3 \times 10^{-26} Kg$

∴ $x = 3$

3. **(3)** moles of Xe = $\frac{55.2\text{ g}}{131\text{ g/mol}} = 42$ mol

moles of Cl = $\frac{44.8\text{g}}{35.5\text{ g/mole}} = 1.26$ mol

$\frac{\text{mol Cl}}{\text{mol Xe}} = \frac{1.26}{.42} = 3$

$\therefore$ Empirical formula = $XeCl_3$

Hence the value of *n* is 3.

4. **(6)** moles of C = $\frac{4\text{g}}{12\text{g/mole}} = 0.333$ moles

moles of H = $\frac{.667\text{ g}}{1\text{g/mole}} = 0.667$ moles

moles of O = $\frac{5.33}{16\text{g/mole}} = 0.333$ moles

$\frac{\text{moles H}}{\text{moles C}} = \frac{0.667}{0.333} = 2$

$\frac{\text{moles C}}{\text{moles O}} = \frac{0.333}{0.333} = 1$

$\frac{\text{moles O}}{\text{moles O}} = \frac{0.333}{0.333} = 1$

$\therefore$ Empirical formula = CH_2O

EFW = 12 + 2(1) + 16 = 30 grams

$\frac{\text{MW}}{\text{EFW}} = \frac{180}{30} = 6$

$\therefore$ Molecular formula = $6(CH_2O) = C_6H_{12}O_6$

Hence the value of *n* is 6.

Think out of the Box

Case Study-1

1.

	Compound's name	Compound's formulae	No. of atoms
(i)	Aluminium sulphate	$Al_2(SO_4)_3$	Al = 2, S = 3, O = 12 , total = 17
(ii)	Magnesium nitrate	$Mg(NO_3)_2$	Mg = 1, N = 2, O = 6, total = 9
(iii)	Aluminium bromide	$AlBr_3$	Al = 1, Br = 3, total = 4
(iv)	Cesium chloride	CsCl	Cs = 1, Cl = 1, total = 2

2.

S. No.	Element	Atomicity
(i)	Sulphur	Polyatomic
(ii)	Helium	Monoatomic
(iii)	Nitrogen	Diatomic
(iv)	Flourine	Diatomic
(v)	Agron	Monoatomic

3.

	Compound's formulae	Cations	Anions
(i)	Na_2O	Na^+	O^{2-}
(ii)	Na_2CO_3	Na^+	CO_3^{2-}
(iii)	$CaCl_2$	Ca^{2+}	Cl^-
(iv)	K_2SO_4	K^+	SO_4^{2-}
(v)	KNO_3	K^+	NO_3^-

Case Study-2

1. Carbon dioxide = CO_2, atomic mass of C = 12 u, O = 16 u.
Therefore, molecular mass of CO_2 = 12 + (2 × 16) = 44 u.

2. Calcium carbonate = $CaCO_3$
Atomic mass of Ca = 40 u, C =12 u, O = 16 u.
Therefore, formula unit mass of $CaCO_3$
= 40 + 12 + (3 × 16) = 100 u.

3. Molecular mass of a chemical compound is 16 u.
Since more than one elements will be involved, x + y = 16
If x = 12, y = 4, x can represent one atom of carbon and y can represent four atoms of hydrogen.
Thus, the molecular formula can be CH_4.

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