IX CBSE CHAPTR - ATOMS AND MOLECULES

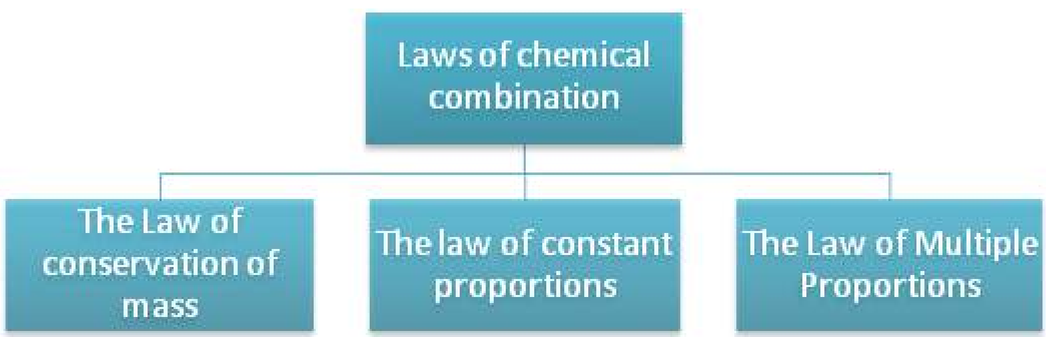
The matter and its composition were always of great interest for scientist and philosopher.

Ancient ideas about the composition of matter:

An Indian philosopher Maharishi Kanad, postulated that if we go on dividing matter (*padarth*), we shall get smaller and smaller particles. Ultimately, a time will come when we shall come across the smallest particles beyond which further division will not be possible. He named these particles *Parmanu*.

Ancient Greek philosophers – Democritus and Leucippus suggested that if we go on dividing matter, a stage will come when particles obtained cannot be divided further. Democritus called these indivisible particles atoms (meaning indivisible). The word atom is derived from the Greek word **"Atomos"** which means indivisible

For the explanation of the question "How and why elements combine and what happens when they combine" Antoine L. Lavoisier along with Proust laid the foundation of chemical sciences by establishing two important laws of chemical combination.



(a) The Law of conservation of mass was stated by Antoine L. Lavoisier in 1785 as" Mass can neither be created nor destroyed in a chemical reaction" [The Law of conservation of mass is the 2^{nd} postulate of Dalton's atomic theory. It states that Atoms are indivisible particles, which cannot be created or destroyed in a chemical reaction.]

Example:

- $C + O_2 \rightarrow CO_2$ Mass of reactant = 12 + 16x2 = 44gm ; Mass of product = 12 + 16x2 = 44gm
  Here Total mass of reactant = Total Mass of product

- Water forms by the union of hydrogen and oxygen. If we weigh the reactants (hydrogen and oxygen) before the chemical reaction, we find the weight of the product (water) equal to the combined weight of reactants.
- The weight of iron increases on rusting. The increase in weight is equal to the weight of oxygen added to iron.
- Carbon combines with Sulphur to form Carbon disulphide. The mass of reactants i.e. carbon and sulphur is same mass of products (carbon disulphide).

Carbon + Sulphur -----➜ Carbon Disulphide

$C + S \longrightarrow C_2S$

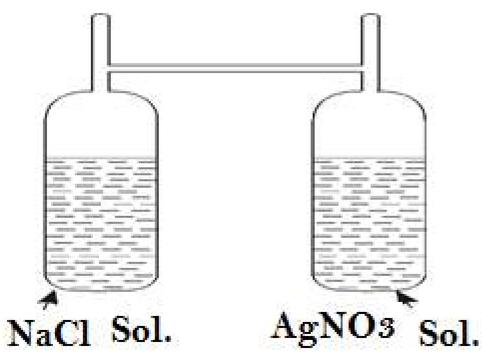
1g + 5. 34g = 6.34 g

LHS = RHS

The verification of the Law of conservation of mass by Landolt Experiment

H. Landolt was German Chemist. He proved the law of conservation of mass by using an H-shaped glass tube. He filled silver nitrate in limb A and hydrochloric acid in limb B. The tube was sealed and weighed before the chemical reaction. The reactants were mixed by inverting and shaking the tube. A white precipitate of silver chloride was formed along with Sodium nitrate. The tube was weighed again. He found that there was no change in weight during the following chemical reaction.

$AgNO_3 + NaCl \rightarrow AgCl_2$ (white precipitate) $+ NaNO_3$



Limitations [Is there any exception to law of conservation of mass?]

In all the chemical reactions, energy is evolved or absorbed which would be at the expense of change in mass. In ordinary chemical reactions, this change in mass is so small that it cannot be registered on the most sensitive balance. This suggests that some matter of the reaction mixture gets converted into energy such as light, heat etc. Thus mass and energy are interconvertible. The mass is converted to energy by Einstein’s relation $E = mc^2$.

OR, Later after atoms were discovered, it was found during nuclear reactions this law does not hold good. In a nuclear reaction, some of the mass gets converted into energy, as given by famous Einstein's mass-energy relationship ($E = mC^2$). The law was given a new name to law of conservation of mass as **Law of conservation of mass and energy.**

Q. Give an example to show Law of conservation of mass applies to physical change also.

Answer: When ice melts into water, is a physical change. Take a piece of ice in small flask, cork it and weight it (say W_{ice} gm). Heat the flask gently and ice (solid) slowly melts into water (liquid). Weigh the flask again (W_{water} gm). It is found there is no change in the weight i.e. $W_{ice} = W_{water}$.

$$\text{Ice} \xrightarrow{\text{Heat}(\Delta)} \text{Water}$$

This shows law of conservation of mass holds true for physical changes.

(b) The law of constant proportions which is also known as the law of definite proportions was stated by Proust in 1799 as "*In a chemical substance the elements are always present in definite proportions by mass*".

[The Law of constant proportions is the 6th postulate of Dalton's atomic theory. The relative number and kinds of atoms are constant in a given compound.]

E.g. In a compound such as water, the ratio of the mass of hydrogen to the mass of oxygen is always 1:8, whatever the source of water. Thus, if 9 g of water is decomposed, 1 g of hydrogen and 8 g of oxygen are always obtained.

if the element 'A' and 'B' combine chemically to form the compound AB, then in whatever manner AB is formed, it is always composed of same two elements 'A' and 'B' combined together in the same fixed ratio or proportion by mass.

For example: Sulphur dioxide can be obtained b following sources:

(i). Sulphur is burnt in air,

$$S + O_2 \longrightarrow SO_2$$

(ii). Copper is heated with conc. sulphuric acid

$$Cu + 2H_2SO_4 \longrightarrow CuSO_4 + 2H_2O + SO_2$$

(iii). Dilute hydrochloric acid is added to sodium bisulphate

$$NaHSO_3 + HCl \longrightarrow NaCl + H_2O + SO_2$$

In each case, the ratio of sulphur and oxygen in the sulphur dioxide obtained is of 32: 32 or 1: 1 by mass.

<u>The Law of Multiple Proportions</u>

when two elements A and B combine to form more than one compounds, then the weight of one is constant and the other has a simple ratio. [The Law of Multiple Proportions is the third postulate of Dalton's atomic theory. It states that the masses of one element which combine with a fixed mass of the second element are in a ratio of whole numbers.]

E.g. Two different compounds are formed by the elements carbon and oxygen.

$$C\ (12gm) + \tfrac{1}{2} O_2\ (16gm) \longrightarrow CO\ (28gm)$$

C (12gm) + O_2 (32gm) ---------→ CO_2 (44gm)

Here, 12 gm of carbon combine with 16g and 32gm of Oxygen to form Carbon monoxide and Carbon dioxide respectively. The ratio of oxygen combining with 12 gm of Carbon is 16: 32 or, 1:2 which is in a simple ratio

Dalton's explanation for the law of conservation of mass and the law of definite proportions

In 1803, A British school teacher John Dalton provided the basic theory about the nature of matter which provides explanation for the law of conservation of mass and the law of definite proportions.

According to Dalton's atomic theory, all matter, whether an element, a compound or a mixture is composed of small particles called atoms. The postulates of this theory may be stated as follows:

(i) All matter is made of very tiny particles called atoms.

(ii) Atoms are indivisible particles, which cannot be created or destroyed in a chemical reaction.

(iii) Atoms of a given element are identical in mass and chemical properties.

(iv) Atoms of different elements have different masses and chemical properties.

(v) Atoms combine in the ratio of small whole numbers to form compounds.

(vi) The relative number and kinds of atoms are constant in a given compound.

Drawbacks of Dalton's Atomic Theory:

- The atom is further subdivided into protons, neutrons and electrons.
- The atoms of same elements are not similar in all respect. They may vary in mass and density. These are known as isotopes. For example: chlorine has two isotopes having mass numbers 35 a.m.u and 37 a.m.u.
- Atoms of different elements are not different in all respects. Atoms of different elements that have the same atomic mass are called isobar.
- According to Dalton atoms of different elements combine in simple whole number ratio to form compounds. This is not seen in complex organic compounds like sugar $C_{12}H_{22}O_{11}$.
- The theory fails to explain the existence of allotropes like Diamond and Graphite which having different properties even these are made up of same kind ot atom namely Carbon.

The introduction of matter wave concept by de Broglie, the principle of uncertainty by Heisenberg etc., paved the way for **modern atomic theory** [Modification in Dalton's atomic theory]

Modifications in Dalton's atomic theory - Modern atomic theory are as follows.

- Atom is considered to be a divisible particle.
- Atoms of the same element may not be similar in all respects. eg: Isotopes ($^{17}Cl_{35}$,$^{17}Cl_{37}$)
- Atoms of different elements may be similar in some respects eg. Isobars ($^{18}Ar_{40}$, ^{20}Ca 40)
- Atom is the smallest particle which takes part in chemical reactions.

- The ratio of atoms in a molecule may be fi xed and integral but may not be simple e.g., $C_{12}H_{22}O_{11}$ is not a simple ratio (Sucrose)
- Atoms of one element can be changed into atoms of other element by transmutation.
- The mass of an atom can be converted into energy. This is in accordance with Einstein's equation $E = mc^2$

Atom: It is the smallest particle of an element which may or may not have independent existence. The atoms of certain elements such as hydrogen, oxygen, nitrogen, etc .do not have independent existence whereas atoms of helium, neon, argon, etc. do have independent existence. Thus we can say that all elements are composed of atoms.

Q. How do we know the presence of atoms if they do not exist independently for most of the elements?

Answer: Atom join in different way to form matter (neutral molecules or ion) that we are able to touch, feel and see.

How big are atoms?

Atoms are extremely small. They are so small, that you cannot see them with most microscopes. Now, Scan tunneling Microscope (STM) is the modern instrument that made it possible to take photograph of atom. The size of an isolated atom can 't be measured because we can 't determine the location of the electrons that surround the nucleus. We can estimate the size of an atom, however, by assuming that the radius of an atom is half the distance between adjacent atoms in a solid. This technique is best suited to elements that are metals, which form solids composed of extended planes of atoms of that element. The results of these measurements are therefore often known as **metallic radii**.

Q. What is the unit of measurement of atomic radius?

Ans: Picometers (pm) or Angstroms (Å)

Q.The size of sodium atom is bigger than that of hydrogen atom. Why?

Answer: Size of atom is the distance between the nucleus and outermost shell (valence shell) of an atom. The atomic number of sodium is greater than that of hydrogen. So, it needs more number of shells to fill electrons and hence will have more number of shells than hydrogen. Hence, atomic size of sodium is bigger than that of sodium.

Naming of an element

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

Q. Why are Dalton's symbols not used in chemistry?

Answer: Dalton was the first scientist to use the symbol for the name of the elements a specific sense but it was difficult to memorize and in uses so Dalton's symbol are not used in chemistry

Berzelius suggested that the symbols of elements be made from one or two letters of the name of the element.

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)

For example: (i) hydrogen, H (ii) aluminum, Al and not AL (iii) cobalt, Co and not CO.

Symbols of some elements are formed from the first letter of the name and a letter, appearing later in the name. Examples are: (i) chlorine, Cl, (ii) zinc, Zn etc.

Other symbols have been taken from the names of elements in Latin, German or Greek. For example, the symbol of iron is Fe from its Latin name ferrum, sodium is Na from natrium, and potassium is K from kalium. Therefore, each element has a name and a unique chemical symbol.

Molecule: A molecule is the smallest or the simplest structural unit of an element (or) a compound which contains one (or) more atoms. It retains the characteristics of an element. A molecule can exist freely and it is a combined form of bonded units whereas an atom is a singular smallest form of non bonded unit.

Molecules are of two types, namely homo atomic molecules and hetero atomic molecules.

Homo atomic molecules: These are the molecules which are made up of atoms of the same element. For example hydrogen gas consists of two atoms of hydrogen (H_2).Similarly oxygen gas consists of two atoms of oxygen (O_2).

HETERO ATOMIC MOLECULES: The hetero atomic molecules are made up of atoms of different elements. They are also classified as diatomic, triatomic, or polyatomic molecules depending upon the number of atoms present. H_2O, NH_3, CH_4, etc., are the examples for hetero atomic molecules.

Atomicity: The number of atoms present in one molecule of an element is called the atomicity of an element. Depending upon the number of atoms in one molecule of an element, molecules are classified into monoatomic, diatomic, triatomic or poly atomic molecules containing one, two, three, or more than three atoms respectively.

Mono atomic molecules: Helium (He) Neon (Ne) Metals ; Di atomic molecules: Hydrogen H_2 Chlorine Cl_2

Tri atomic molecules: Ozone (O_3) ; Poly atomic molecules: phosphorous P_4 Sulphur S_8

Atomicity = Molecular Mass/Atomic mass

MORE TO KNOW:

Isotopes ⇒ These are the atoms of same element with same atomic number (Z) but different mass number (A). Example ($^{17}Cl_{35}$,$^{17}Cl_{37}$)

Isobars ⇒ These are the Atoms of the different element with same mass number but different atomic number. Example ($^{18}Ar_{40}$, $^{20}Ca_{40}$)

Isotones ⇒ These are the atoms of different elements with same number of neutrons Example ($^{6}C_{13}$, $^{7}N_{14}$)

AVOGADRO'S HYPOTHESIS: Amedeo Avogadro put forward hypothesis and is based on the relation between number of molecules and volume of gases that is "volume of a gas at a given temperature and pressure is proportional to the number of particles".

Avogadro's Law: Equal volumes of all gases under the same conditions of temperature and pressure. contain the equal number of molecules.

TEST YOUR UNDERSTANDING SKILL

(a) Find the atomicity of chlorine if its atomic mass is 35.5 and its molecular mass is 71

(b) Find the atomicity of ozone if its atomic mass is 16 and its molecular mass is 48

WHAT IS AN ION?

An ion is a charged particle and can be negatively or positively charged.

A negatively charged ion is called an 'anion' and the positively charged ion, a 'cation'. For example, sodium chloride (NaCl). Its constituent particles are positively charged sodium ions (Na+) and negatively charged chloride ions (Cl–).

Ions may consist of a single charged atom or a group of atoms that have a net charge on them. A group of atoms carrying a charge is known as a polyatomic ion e.g. Calcium oxide (Ca^{+2} O^{-2})

Atomic mass and Relative Atomic mass (RAM):

Q. Each element had a characteristic atomic mass even then we are using Relative Atomic mass. Give reason?

Answer: Since determining the mass of an individual atom was a relatively difficult task due to extremely smaller size, relative atomic masses were determined using the laws of chemical combinations and the compounds formed.

Relative Atomic mass (RAM): In 1961 IUPAC selected an isotope of carbon (^{12}C) as a slandered for comparing atomic and molecular mass of element and compound.

Relative atomic mass of an element is the ratio of mass of one atom of element to the **1/12**th part of mass of one atom of carbon. Relative atomic mass is a pure ratio and has no unit. If the atomic mass of an element is expressed in grams, it is known as **gram atomic mass. e.g.** Gram atomic mass of hydrogen = 1g where as gram atomic mass of carbon = 12g

Atomic mass is expressed in atomic mass unit (amu). One atomic mass unit is defined as 1/12th part of the mass of one atom of carbon.

Q. The atomic mass of an element is in fraction .What does it mean?

Ans If the atomic mass of an element is in fraction, this mean that it exists in the form of isotopes. The atomic mass is the average atomic mass and is generally fractional.

Chemical Formulae*:* The chemical formula is a symbolic representation of a compound of its composition.

For writing Chemical Formulae the name or symbol of the metal is written first then non-metals with their valencies. Then we must crossover the valencies of the combining atoms. For example:

(a) Formula for aluminium oxide: $Al^{3+}\ O^{-2} \Rightarrow Al_2O_3$ (b) calcium hydroxide : $Ca^{+2}\ OH^{-1} \Rightarrow Ca(OH)_2$

Valency: The combining power (or capacity) of an element is known as its valency. Valency can be used to find out how the atoms of an element will combine with the atom(s) of another element to form a chemical compound.

RELATIVE MOLECULAR MASS (RMM) : The relative molecular mass of an element or a compound is the ratio of mass of one molecule of the element or a compound to the mass of 1/12 th part of mass of one atom of carbon. Relative Molecular mass is a pure ratio and has no unit. If the molecular mass of a given substance is expressed in gram, it is known as **gram molecular mass** of that substance.

Molecular mass is the sum of the masses of all the atoms present in one molecule of the compound or an element.

Key Concepts

- Relative molecular mass is also known as molecular weight, relative molar mass, molar weight, formula mass and formula weight.
- Relative molecular mass is usually given the symbol M_r
  Other symbols commonly used are MM, MW, FM, FW*.
- Relative molecular mass of a compound is defined as the mass of a formula unit of the compound relative to the mass of a carbon atom taken as exactly 12.
- In practice, the relative molecular mass, M_r, of a compound is the sum of the relative atomic masses (atomic weights) of the atomic species as given in the chemical formula.
- Relative molecular mass is a dimensionless quantity, it has no units #.

Relative Molecular Mass Calculations

Example 1

Calculate the relative molecular mass (M_r) of the compound carbon monoxide, CO

a. Determine the number of atoms of each element present in the chemical formula:
   CO is composed of one atom of C (carbon) and one atom of O (oxygen)
b. Use the Periodic Table to find the relative atomic mass (atomic weight) for each element present:
   relative atomic mass of C (carbon) = 12.0
   relative atomic mass of of O (oxygen) = 16.0
c. Calculate the relative molecular mass
   M_r = sum of the atomic masses present in the formula
   $M_{r(CO)}$ = (1 x relative atomic mass of carbon) + (1 x relative atomic mass of oxygen)
   $M_{r(CO)} = (1 \times 12.0) + (1 \times 16.0) = 28.0$

Test your numerical skill:

Problem: Find the gram molecular mass of water (H2O)

Solution: $\Rightarrow$ 2(H) = 2 x 1 = 2 and 1(O) = 1 x 16 = 16 ; Gram molecular mass of H2O = 2 + 16 = 18g

Problem: Find the gram molecular mass of carbon dioxide

Solution: $\Rightarrow$ (CO2) 1(C) = 1 x 12 = 12 and 2(O) = 2 x 16 = 32; Gram molecular mass of CO2 = 12 + 32 = 44 g

Empirical formula and molecular formula:

Empirical formula: The empirical formula is the simplest formula for a compound in which atoms of different elements are present in simple ratio. It shows the relative number of atoms of each element. For example CH_2O is the empirical formula of Glucose $C_6H_{12}O_6$

Calculating Empirical Formula from Percentage Composition

Assume 100g of sample

Convert all percentages to a mass in grams, eg, 21% = 21g, 9% = 9g

Find the relative atomic mass (r.a.m) of each element present using the Periodic Table

Calculate the moles of each element present: n = mass ÷ r.a.m

Divide the moles of each element by the smallest of these to get a mole ratio

If the numbers in the mole ratio are all whole numbers (integers) convert this to an empirical formula

If the numbers in the mole ratio are NOT whole numbers, you will need to further manipulate these until the mole ratio is a ratio of whole numbers (integers)

A compound is found to contain 47.25% copper and 52.75% chlorine. Find the empirical formula for this compound.

element	Cu	Cl
mass in grams	47.25	52.75
r.a.m	63.6	35.5
moles = mass ÷ r.a.m	47.25 ÷ 63.6 = 0.74	52.75 ÷ 35.5 = 1.49
divide throughout by lowest number	0.74 ÷ 0.74 = 1	1.49 ÷ 0.74 = 2.01 = 2

Empirical formula for this compound is $CuCl_2$

Percent Composition (Percentage Composition)

Key Concepts

The percent composition (percentage composition) of a compound is a relative measure of the mass of each different element present in the compound.

To calculate the percent composition (percentage composition) of a compound:

- Calculate the relative molecular mass (molecular weight, formula mass, formula weight), M_r, of the compound
- Calculate the total mass of each element present in the formula of the compound
- Calculate the percentage compositon : % by weight (mass) of element
  = (total mass of element present ÷ molecular mass) x 100

Examples

Example 1.Calculate the percent by mass (weight) of sodium (Na) and chlorine (Cl) in sodium chloride (NaCl)

a. Calculate the relative molecular mass (M_r): M_r = 22.99 + 35.45 = 58.44
b. Calculate the total mass of Na present: 1 Na is present in the formula, mass = 22.99
c. Calculate the percent by mass (weight) of Na in NaCl:

   %Na = (mass Na ÷ M_r) x 100 = (22.99 ÷ 58.44) x 100 = 39.34%
d. Calculate the total mass of Cl present:
   1 Cl is present in the formula, mass = 35.45
e. Calculate the percent by mass (weight) of Cl in NaCl:
   %Cl = (mass Cl ÷ M_r) x 100 = (35.45 ÷ 58.44) x 100 = 60.66%

The answers above are probably correct if %Na + %Cl = 100, that is, 39.34 + 60.66 = 100.

Example 2

Calculate the percent by mass (weight) of each element present in sodium sulfate (Na_2SO_4).

a. Calculate the relative molecular mass (M_r):
   M_r = (2 x 22.99) + 32.06 + (4 x 16.00) = 142.04
b. Calculate the total mass of Na present:
   2 Na are present in the formula, mass = 2 x 22.99 = 45.98
c. Calculate the percent by mass (weight) of Na in Na_2SO_4:
   %Na = (mass Na ÷ M_r) x 100 = (45.98 ÷ 142.04) x 100 = 32.37%
d. Calculate the total mass of S present in Na_2SO_4:
   1 S is present in the formula, mass = 32.06

e. Calculate the percent by mass (weight) of S present:
   %S = (mass S ÷ M_r) x 100 = (32.06 ÷ 142.04) x 100 = 22.57%

f. Calculate the total mass of O present in Na_2SO_4:
   4 O are present in the formula, mass = 4 x 16.00 = 64.00

g. Calculate the percent by mass (weight) of O in Na_2SO_4:
   %O = (mass O ÷ M_r) x 100 = (64.00 ÷ 142.04) x 100 = 45.06%

The answers above are probably correct if %Na + %S + %O = 100, that is,
32.37 + 22.57 + 45.06 = 100

<u>Molecular formula</u>: It is the formula in which the actual number of atoms of different element are present. For example, if the empirical formula of benzene is CH where as molecular formula is C_6H_6, etc.

An empirical formula is often calculated from elemental composition data. The weight percentage of each of the elements present in the compound is given by this elemental composition.

Let's determine the empirical formula for a compound with the following elemental composition:40.00% C, 6.66% H, 53.34% O.

Element	percentage	Atomic mass	Relative number of Atoms	Dividing by least number	Simple ratio
C	40	12	40/12 = 3.33	3.33/3.33	1
H	6.66	1	6.66/1 =6.66	6.66/3.33	2
O	53.34	16	53.34/16 =3.33	3.33/3.33	1

Empirical formula =$C_1H_2O_1$; Empirical formula mass = 12 + 2x1 +16 = 30 a.m.u

Given relative molecular mass = 180

Divide the relative molecular mass by the Empirical formula mass to find a multiple: 180/30 = 6

The molecular formula is a multiple of 6 times the empirical formula: $(C_1H_2O_1)$x6 = $C_6H_{12}O_6$

Q. A compound with a molecular mass of 34.0g/mol is known to contain 5.88% hydrogen and 94.12% oxygen. Find the molecular formula for this compound.

First, find the empirical formula of the compound.

element	H	O
mass in grams	5.88	94.12
r.a.m	1.0	16.0
moles = mass ÷ r.a.m	5.88 ÷ 1.0 = 5.88	94.12 ÷ 16.0 = 5.88
divide throughout by the smallest number	5.88 ÷ 5.88 = 1	5.88 ÷ 5.88 = 1

Empirical formula is HO

Calculate the empirical formula mass: 1.0 + 16.0 = 17.0 g/mol

Molecular Mass = n x empirical formula mass

34.0 = n x 17.0

n = 34.0 ÷ 17.0 = 2

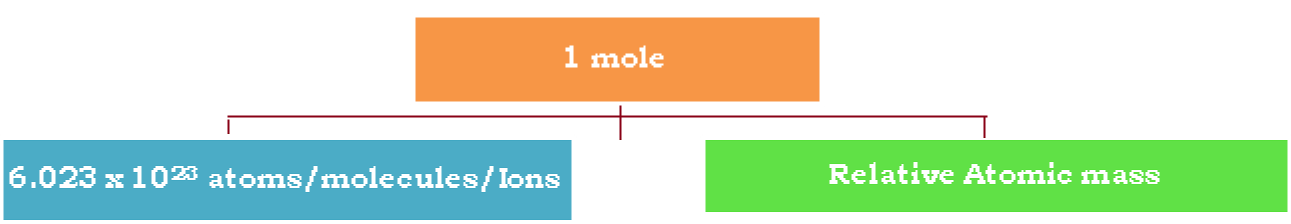
Molecular Formula is 2 x (HO) which is H_2O_2

MOLE CONCEPT

While performing a reaction, to know the number. of atoms (or) molecules involved, the **concept of mole** was introduced. The quantity of a substance is expressed in terms of mole.

Definition of mole: Mole is defined as the amount of substance that contains as many specified elementary particles as the number of atoms in 12g of carbon-12 isotope.

One mole is also defined as the amount of substance which contains Avogadro number (6.023×10^{23}) of particles.



MORE TO KNOW

Avogadro number: Number of atoms or molecules or ions present in one mole of a substance is called Avogadro number. Its value is 6.023×10^{23}.

Therefore, one mole of any substance = 6.023×10^{23} particles may be atoms, molecules, ions

For example: One mole of oxygen atoms represents 6.023×10^{23} atoms of oxygen and 5 moles of oxygen atoms contain $5 \times 6.023 \times 10^{23}$ atoms of oxygen.

Questions based on mole concept:

1. When the mass of the substance is given: Use this formula: Number of moles = given mass/ atomic mass

(a). Calculate the number of moles in (i) 81g of aluminium ii) 4.6g sodium (iii) 5.1g of Ammonia (iv) 90g of water (v) 2g of NaOH
Solution: (i) Number of moles of aluminum = given mass of aluminium / atomic mass of aluminium = 81/27 = 3 moles of aluminum [Rest Question do yourself]

(b) Calculate the mass of 0.5 mole of iron

Solution: mass = atomic mass x number of moles = 55.9 x 0.5 = 27.95 g

Do yourself: Find the mass of 2.5 mole of oxygen atoms [Mass = molecular mass x number of moles]

2. Calculation of number of particles when the mass of the substance is given:

Number of particles =(Avogadro number x given mass)/gram molecular mass

Problem: Calculate the number. of molecules in 11g of CO_2

Solution: gram molecular mass of CO_2 = 44g

Number of molecules = $(6.023 \times 10^{23} \times 11) / 44 = 1.51 \times 10^{23}$ molecules

Do yourself: Calculate the number of molecules in 360g of glucose

3. Calculation of mass when number of particles of a substance is given:

Mass of a substance = (gram molecular mass x number of particles)/6.023×10^{23}

Problem: Calculate the mass of 18.069×10^{23} molecules of SO_2

Solution: Gram molecular mass SO_2 = 64gm

The mass of 18.069×10^{23} molecules of SO_2 = $(64 \times 18.069 \times 10^{23}) / (6.023 \times 10^{23})$ = 192 g

Do yourself: (a) Calculate the mass of glucose in 2 x 10^{24} molecules (b) Calculate the mass of 12.046 x 1023 molecules in CaO

4. Calculation of number of moles when you are given number of molecules:

Problem: Calculate the number moles for a substance containing 3.0115 x 10^{23} molecules in it.

Solution: Number of moles = [Number of molecules/(6.023 x 10^{23})] = (3.0115 x 10^{23})/(3.0115 x 10^{23}) =0.5 moles

Do yourself: (a) Calculate number of moles in 12.046x 1022 atoms of copper (b) Calculate the number of moles in 24.092 x 1022 molecules of water.

Problem: Calculate the number of aluminum ions present in 0.051 g of aluminum oxide. (Hint: The mass of an ion is the same as that of an atom of the same element. Atomic mass of Al=27 u)

Solution: Mass of the 1 mole of Al_2O_3 = 2x27 + 3x16 = 102gm

The number of ions present in 102 gm of aluminum oxide = 6.023 x 10^{23} ion

The number of ions present in 0.051g of aluminum oxide= (6.023 x 10^{23} ion x 0.051g)/ 102 gm

= 6.023 x 10^{23} ion x0.0005 = 3.0115 x 10^{20} ions

In Al_2O_3, Aluminium and oxygen are in ratio 2:3

So, The number of aluminum ions present(Al^{3+}) in 0.051g of aluminum oxide = 2 x 3.0115 x 10^{20} ions =6.023 x 10^{20} ion

MORE TO KNOW

Volume occupied by one mole of any gas at STP is called molar volume. Its value is 22.4 litres 22.4 litres of any gas contains 6.023 x 1023 molecules.

Problem: Calculate the volume occupied at STP by :- (a) 64 gram of oxygen gas (b) 6.02 x10^{22} molecules of CH_4 (c) 5 moles of nitrogen gas

Solution: (a) One mole of a gas occupies 22.4 L volume at STP

Now, number of moles in 64 g oxygen gas = 64/32 = 2

Therefore, volume occupied by 2 moles(64 g) of oxygen gas = 2 x 22.4 L = 44.8 L

(b) 1 mole = 6.02 x 10^{23} molecules

Therefore, 1 mole (6.02 x 10^{23} molecules) of CH_4 gas occupies 22.4 L volume at STP.

(c) One mole of a gas occupies 22.4 L volume at STP

Therefore, volume occupied by 5 moles of nitrogen gas = 5 x 22.4 L = 112 L

Problem: Calculate the volume at STP occupied by 10^{21} molecules of Oxygen?

Solution: The molar volume that is the volume occupied by one mole of gas is 22.4 L. We know there are 6.022 X 10^{23} particles of a substance in one mole of that substance. Thus

volume occupied by 6.022 X 10^{23} molecules of oxygen = 22.4 L

volume occupied by 10^{21} molecule of oxygen = = [(22.4 X 10^{21}) / (6.022 X 10^{23})] = 3.72 X 10^{-2} L = 37.2 ml

Q. Calculate the mass and volume of oxygen at STP, which will be evolved on electrolysis of 1 mole of(18 g) of water.

Answer: Electrolysis of water will break it down in its component as Hydrogen and Oxygen

Balanced chemical reaction:

$2H_2O \rightarrow 2H_2 + O_2$

From above equation, 2 mole of water will evolve 1 mole of oxygen gas upon electrolysis. Therefore 1 mole of water will produce 1/2 mole of Oxygen.

Mass of Oxygen evolved = number of moles of Oxygen evolved × Molecular wt. of Oxygen =1/2 × 32 = **16 g**

At STP 1 mole of any gas will occupy 22.4l volume.

Volume of Oxygen evolved = No. of moles × 22.4l = 1/2 × 22.4l = **11.2l**.

Q. How many molecules are present in 1 ml of water?

Answer: we know that density of water is 1gm/ml.

Hence, 1 gm water will = 1 ml water.

Now, we have molecular mass of water H_2O = 1x2 + 16 = 18 gm

18 gm of water contain 6.022 x 10^{23} molecules

1 gm of water will contain = $(6.022 \times 10^{23})/18$ molecules = 0.33×10^{23} molecules

So, the no. of molecules of water in 1ml of water = 3.3×10^{22}

Definitions of a mole

Key Concepts

- Mole is abbreviated to mol and given the symbol n
- 1 mole contains the same number of particles as there are in 12g of carbon-12 atoms by definition.
  This number is called the Avogadro number or Avogadro constant (N_A or L) and is equal to 6.022×10^{23} particles.
- 1 mole of a pure substance has a mass in grams equal to its relative molecular mass(M_r) (also known as molecular weight or formula mass or formula weight).
  This is known as the molar mass and is given the symbol M
  The units for molar mass are g mol^{-1} *
- 1 mole of a gas occupies a specific volume at a particular temperature and pressure.
  This is known as the molar volume and given the symbol V_m
  Standard Temperature and Pressure (STP) is defined as a temperature of 0^oC (273.15 K) and a pressure of 100 kPa (0.987 atm)#
  At STP 1 mole of an ideal gas has a volume of 22.71 L (V_m = 22.71 L mol^{-1})

At 25^oC (298.15 K) and 100 kPa (0.987 atm) 1 mole of an ideal gas has a volume of 24.79 L (V_m = 24.79 L mol^{-1})**

Examples

Avogadro Number (N_A)

1 mole of particles contains the Avogadro Number, N_A, of particles.
$N_A = 6.022 \times 10^{23}$ particles per mole

- 1 mole of helium atoms contains 6.022×10^{23} helium atoms.
- 1 mole of carbon monoxide molecules contains 6.022×10^{23} carbon monoxide

molecules.

- 1 mole of H_2O contains 6.022 x 10^{23} H_2O molecules.
- 1 mole of NaCl contains 6.022 x 10^{23} NaCl units.

Molar Mass (M)

1 mole of a pure substance has a mass in grams equal to its relative molecular mass (M_r).

- 1 mole of helium gas (a monatomic gas with the formula He)
  chemical formula of helium gas is He
  relative molecular mass He = 4.003
  1 mole of He has a mass equal to its relative atomic mass in grams = 4.003 g
  Molar mass of helium = 4.003 g mol^{-1}
  M(He) = 4.003 g mol^{-1}

- 1 mole of carbon monoxide gas
  chemical formula of carbon monoxide is CO
  relative molecular mass of CO = 12.01 + 16.00 = 28.01
  1 mole of CO has a mass equal to its relative atomic mass in grams = 28.01 g
  Molar mass of CO = 28.01 g mol^{-1}
  M(CO) = 28.01 g mol^{-1}

- 1 mole of H_2O
  relative molecular mass H_2O = (2 x 1.008) + 16.00 = 18.016
  1 mole of H_2O has a mass equal to its relative atomic mass in grams = 18.016 g
  Molar mass of H_2O = 18.016 g mol^{-1}
  $M(H_2O)$ = 18.016 g mol^{-1}

- 1 mole of NaCl
  relative molecular mass NaCl = 22.99 + 35.45 = 58.44
  1 mole of NaCl has a mass equal to its relative atomic mass in grams = 58.44 g
  Molar mass of NaCl = 58.44 g mol^{-1}
  M(NaCl) = 58.44 g mol^{-1}

Ideal Gas Volumes

At standard temperature and pressure (STP), defined as a temperature of 0^oC (273.15 K) and a pressure of 100 kPa (0.987 atm), an ideal gas has a volume of 22.71 L.

- 1 mole of helium gas at STP has a volume of 22.71 L
  Molar volume of helium gas is 22.71 L
  $V_m(He_{(g)}) = 22.71$ L

- 1 mole of carbon monoxide gas at STP has a volume of 22.71 L
  Molar volume of carbon monoxide gas is 22.71 L
  $V_m(CO_{(g)}) = 22.71$ L

- 1 mole of $H_2O_{(g)}$ at STP has a volume of 22.71 L
  Molar volume of $H_2O_{(g)}$ is 22.71 L
  $V_m(H_2O_{(g)}) = 22.71$ L

Mole-Number of Particles Calculations

Key Concepts

- 1 mole of any substance contains 6.022×10^{23} particles.
- 6.022×10^{23} is known as the Avogadro Number or Avogadro Constant and is given the symbol N_A or L
- To find the number of particles in a mole of substance, multiply the moles by the Avogadro number:

1 mole contains 6.022×10^{23} particles

2 moles contains $2 \times 6.022 \times 10^{23} = 1.204 \times 10^{24}$

10 moles contains $10 \times 6.022 \times 10^{23} = 6.022 \times 10^{24}$

0.5 moles contains $0.5 \times 6.022 \times 10^{23} = 3.011 \times 10^{23}$

- This leads to the equation: $N = n \times N_A$
  where N = number of particles in the substance
  and n = moles of substance

and N_A = Avogardro Number = 6.022×10^{23} particles mol^{-1}

- To find the number of particles, N, in a substance: $N = n \times N_A$
- To find the moles, n, of substance, $n = N/N_A$

Examples

Find the number of particles

1. Find the number of ammonia, NH_3, molecules in 3.5 moles of ammonia.

a. Extract the data from the question:

moles of ammonia, $n(NH_3) = 3.5$ mol

number of ammonia molecules, $N(NH_3)$ = ?

b. Write the equation: $N(NH_3) = n(NH_3) \times N_A$

N_A is the Avogadro constant = 6.022×10^{23} molecules mol^{-1}

c. Substitute in the vales and solve:

$N(NH_3) = 3.5 \times 6.022 \times 10^{23} = 2.1 \times 10^{24}$ ammonia molecules

2. Find the number of hydrogen atoms in 1.5 moles of water, H_2O.

a. Extract the data from the question:

moles of water, $n(H_2O) = 1.5$ mol

number of H atoms, N(H) = ?

b. Calculate the moles of H atoms present in 1.5 mol H_2O:

From the chemical formula we see that 1 molecule of water is made up of 2 atoms of hydrogen and 1 atom of oxygen.

So, 1 mole of water molecules contains 2 x 1 = 2 moles of hydrogen atoms.

Therefore 1.5 moles of water molecules contains 1.5 x 2 = 3.0 moles of hydrogen atoms.

n(H) = 3.0 mol

c. Calculate the number of hydrogen atoms using the equation $N(H) = n(H) \times N_A$

$N(H) = 3.0 \times 6.022 \times 10^{23} = 1.8 \times 10^{24}$ hydrogen atoms.

Find the moles of substance

1. A sample of gas contains 4.4×10^{24} carbon dioxide molecules.

How many moles of carbon dioxide are present in the sample?

a. Extract the data from the question:

Let $N(CO_2)$ be the number of carbon dioxide molecules

$N(CO_2) = 4.4 \times 10^{24}$ molecules

moles of CO_2, $n(CO_2) = ?$

b. Write the equation: $N(CO_2) = n(CO_2) \times N_A$

where N_A, the Avogadro constant, $= 6.022 \times 10^{23}$ molecules mol^{-1}

c. Rearrange the equation to find $n(CO_2)$:

$n(CO_2) = N(CO_2)/N_A$

d. Substitute the values into the equation and solve:

$n(CO_2) = 4.4 \times 10^{24} / 6.022 \times 10^{23} = 7.3$ mol

2. A sample contains 2.4×10^{22} molecules of oxygen gas (O_2).

How many moles of oxygen atoms are present in the sample?

a. Extract the data from the question:

Let $N(O_2)$ be the number of oxygen gas molecules

$N(O_2) = 2.4 \times 10^{22}$ molecules

moles of O atoms, $n(O) = ?$

b. Calculate the number of oxygen atoms present:

1 molecule of O_2 contains 2 molecules of O atoms

2.4×10^{22} O_2 molecules contains $2 \times 2.4 \times 10^{22}$ O atoms

$N(O) = 2 \times 2.4 \times 10^{22} = 4.8 \times 10^{22}$ O atoms

c. Write the equation: $N(O) = n(O) \times N_A$

where N_A, the Avogadro constant, $= 6.022 \times 10^{23}$ molecules mol^{-1}

d. Rearrange the equation to find n(O):

$n(O) = N(O)/N_A$

e. Substitute the values into the equation and solve:

$n(O) = 4.8 \times 10^{22} / 6.022 \times 10^{23} = 0.08$ mol

Molar Gas Volume

Key Concepts

- 1 mole of a gas occupies a specific volume at a particular temperature and pressure.
- This is known as the molar volume and given the symbol V_m
- The units most commonly used for molar volume, V_m, are litres per mole, L mol^{-1}
- Examples of molar gas volume (V_m) for ideal gases:

Temperature		Pressure		Molar Volume
^oC	(K)	kPa	(atm)	(V_m) / L mol^{-1}
0^oC	(273.15K)	100kPa	(0.987 atm)	22.71
25^oC	(298.15 K)	100 kPa	(0.987 atm)	24.79

- 0^oC (273.15K) and 100kPa (0.987 atm) is known as Standard Temperature and Pressure and is often abbreviated to STP *
  25^oC (298.15 K) and 100 kPa (0.987 atm) is sometimes referred to as Standard Ambient Temperature and Pressure, SATP, or even as Standard Laboratory Conditions, SLC.**
- Calculations involving molar gas volumes:
  n(gas) = moles of gas
  V(gas) = volume of gas
  V_m = molar gas volume (at some specified temperature and pressure)

$$n(gas) = V(gas) / V_m$$

or

$$V(gas) = n(gas) \times V_m$$

Example : Calculating Moles of Gas

1. A sample of pure helium gas occupies a volume of 6.8 L at 0^oC and 100 kPa.
How many moles of helium gas are persent in the sample?

a. Extract the data from the question:
   V(He) = 6.8 L
   V_m = 22.71 L mol^{-1} (at STP 1 mole of gas occupies 22.71 L)

$n(He) = ?$ mol

b. Check for consistency in units, are all the volumes in the same units?

$V(He)$ is given in L

V_m is given in L (mol^{-1})

Both volumes are in the same units, L, so no conversion is necessary.

c. Write the equation:

$n(He) = V(He) / V_m$

d. Substitute the values into the equation and solve:

$n(He) = 6.8 / 22.71 = 0.3$ L

2 A sample of nitrogen gas, $N_{2(g)}$, has a volume of 956 mL at 273.15K and 100kPa. How many moles of nitrogen gas are present in the sample?

a. Extract the data from the question:

$V(N_2) = 956$ mL

$V_m = 22.71$ L mol^{-1} (at STP 1 mole of gas occupies 22.71 L)

$n(N_2) = ?$ mol

b. Check for consistency in units, are all the volumes in the same units?

$V(N_2)$ is given in mL

V_m is given in L (mol^{-1})

Convert the gas volume, $V(N_2)$, from a volume in millilitres, mL, to a volume in litres, L.

$V(N_2) = 956$ mL $= 956 \times 10^{-3}$ L

c. Write the equation:

$n(N_2) = V(N_2) / V_m$

d. Substitute the values into the equation and solve:

$n(N_2) = 956 \times 10^{-3} / 22.71 = 0.04$ L

Example : Calculating Volume of Gas

1. A balloon contains 0.5 moles of pure helium gas at standard temperature and pressure. What is the volume of the balloon?

a. Extract the data from the question:

$n(He) = 0.5$ mol

$V_m = 22.71$ L mol^{-1} (at STP 1 mole of gas occupies 22.71 L)

$V(He) = ?$ L

b. Write the equation:

$V(He) = n(He) \times V_m$

c. Substitute in the values and solve:

$V(He) = 0.5 \times 22.71 = 11.4$ L

2. What is the volume occupied by 3.7 moles of $N_{2(g)}$ at STP?

a. Extract the data from the question:

$n(N_2) = 3.7$ mol

$V_m = 22.71$ L mol^{-1} (at STP 1 mole of gas occupies 22.71 L)

$V(N_2) = ?$ L

b. Write the equation:

$V(N_2) = n(N_2) \times V_m$

c. Substitute in the values and solve:

$V(N_2) = 3.7 \times 22.71 = 84.0$ L