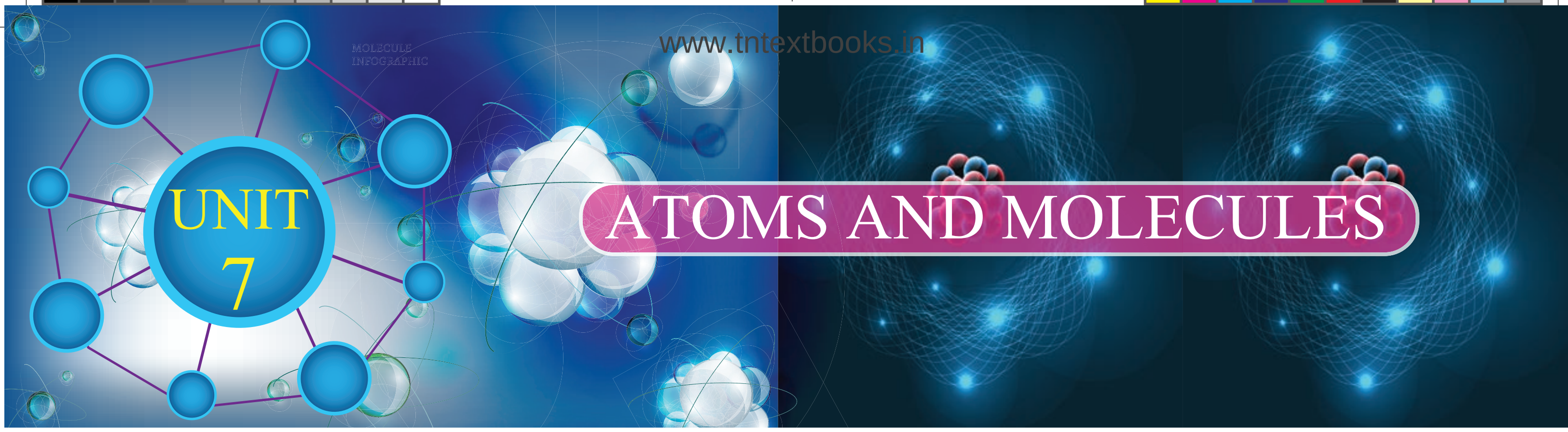

UNIT 7

ATOMS AND MOLECULES

Learning Objectives

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At the end of the lesson the students will be able to:

- acquire the ability to learn about the atoms and molecules.
- comprehend atomic mass and molecular mass.
- have information about gram atomic mass and gram molecular mass.
- perceive the intended meaning of Avogadro's hypothesis of gases.
- interpret the application of Avogadro's hypothesis.
- determine the atomicity of a molecule.
- interpret the relation between vapour density and relative molecular mass.
- have the facts about the relationship between the volume of a gas and the number of molecules present in it.
- grasp the idea of mole concept and solve many problems using it.
- calculate the percentage of composition of a compund.

INTRODUCTION

You have learnt, in your lower classes that matter is around us everywhere. Matter is made of atoms. Curiously the idea of atom was first proposed by the Greek philosophers in the fifth century BC (BCE). But, their theory was more philosophical than scientific.

The first scientific theory of the atom was proposed by John Dalton. Few of the postulates of Dalton's theory about an atom were found incorrect by the later on studies made by J.J. Thomson, Rutherford, Neils Bohr and Schrodinger. In the light of the result of the researches most of the limitations of the Dalton's theory were removed and a new theory known as the modern atomic theory was put forward. **'The main postulates of modern atomic theory'** are as follows:

- **An atom is no longer indivisible** (after the discovery of the electron, proton, and neutron).
- Atoms of the same element may have different atomic mass. (discovery of **isotopes** $_{17}Cl^{35}$, $_{17}Cl^{37}$).
- Atoms of different elements may have same atomic masses (discovery of **Isobars** $_{18}Ar^{40}$, $_{20}Ca^{40}$).
- Atoms of one element can be transmuted into atoms of other elements. In other words, atom is no longer indestructible (discovery of **artificial transmutation**).

- Atoms may not always combine in a simple whole number ratio (E.g. Glucose $C_6H_{12}O_6$ C:H:O = 6:12:6 or 1:2:1 and Sucrose $C_{12}H_{22}O_{11}$ C:H:O = 12:22:11).
- Atom is the **smallest particle that takes part in a chemical reaction.**
- The mass of an atom can be converted into energy ($E = mc^2$).

The modern atomic theory is the basis for all the studies of chemical and physical processes that involve atoms. You have studied the most fundamental ideas about an atom in your lower classes. Let us discuss some more concepts about atoms in this lesson.

7.1 ATOM AND ATOMIC MASS

As you know, anything that has mass and occupies space is called matter. Atoms are the building blocks of matter. Since matter has mass, it must be due to its atoms. According to the modern atomic theory, an atom contains subatomic particles such as protons, neutrons and electrons. **Protons and neutrons have considerable mass, but electrons don't have such a considerable mass**. Thus, the mass of an atom is mainly contributed by its protons and neutrons and hence **the sum of the number of protons and neutrons of an atom is called its mass number**.

Individual atoms are very small and it is difficult to measure their masses. You can measure the mass of macroscopic materials in gram or kilogram. The mass of an atom is measured in atomic mass unit (amu).

Atomic mass unit is one-twelfth of the mass of a carbon-12 atom; an isotope of carbon, which contains 6 protons and 6 neutrons.

(**Note:** The symbol 'amu' is no longer used in the modern system and instead, it uses the symbol 'U' to denote unified atomic mass. The mass of a proton or neutron is approximately 1 amu).

7.1.1 Relative Atomic Mass (RAM)

As an atom is very small, its absolute mass cannot be determined directly. The early pioneers of chemistry used to measure the atomic mass of an atom relative to an atom of another element. They measured the masses of equal number of atoms of two or more elements at a time, to determine their relative masses. They established one element as a standard, gave it an arbitrary value of atomic mass and using this value they measured the relative mass of other elements. The mass obtained by this way is called relative atomic mass. In the beginning, the mass of hydrogen atom was chosen as a standard and masses of other atoms were compared with it, because of the existence of isotopic character of hydrogen (${}_1H^1$, ${}_1H^2$, ${}_1H^3$). Later hydrogen atom was replaced by oxygen atom as the standard. Now, the stable isotope of carbon (C-12) with atomic mass 12 is used as the standard for measuring the relative atomic mass of an element.

Relative atomic mass of an element is the ratio between the average mass of its isotopes to $\frac{1}{12^{th}}$ part of the mass of a carbon-12 atom. It is denoted as A_r. It is otherwise called "Standard Atomic Weight".

Relative Atomic Mass

$$(A_r) = \frac{\text{Average mass of the isotopes of the element}}{\frac{1}{12^{th}} \text{ of the mass of one Carbon-12 atom}}$$

Modern methods of determination of atomic mass by Mass Spectrometry uses C-12 as standard. For most of the elements, the relative atomic mass is very closer to a whole number and it is rounded off to a whole number, to make calculations easier. Table 7.1 lists some of the elements of periodic table and their A_r values.

Table 7.1 Relative atomic mass of elements (C-12 Scale)

Element	Symbol	A_r
Hydrogen	H	1
Carbon	C	12
Nitrogen	N	14
Oxygen	O	16
Sodium	Na	23
Magnesium	Mg	24
Sulphur	S	32

DO YOU KNOW?

Relative Atomic Mass is only a ratio, so it has no unit. If the atomic mass of an element is expressed in grams, it is called as **Gram Atomic Mass**

Gram Atomic Mass of hydrogen = 1 g
Gram Atomic Mass of carbon = 12 g
Gram Atomic Mass of nitrogen = 14 g
Gram Atomic Mass of oxygen = 16 g

7.1.2 Average Atomic Mass (AAM)

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How can one measure the atomic mass of an element? It is somewhat more complicated because most of the naturally occuring elements exist as a mixture of isotopes, each of which has its own mass. Thus, it is essential to consider this isotopic mixture while calculating the atomic mass of an element.

The average atomic mass of an element is the weighted average of the masses of its naturally occurring isotopes.

But, the abundance of isotopes of each element may differ. So, the abundancy of all these isotopes are taken into consideration while calculating the atomic mass. Then, what do we mean by a weighted average? Let us consider an element which exists as a mixture of 50% of an isotope having a mass of 9 amu, and 50% of another isotope having a mass of 10 amu. Then, its average atomic mass is calculated by the following equation:

Average atomic mass

= (Mass of 1st isotope × % abundance of 1st isotope) + (Mass of 2nd isotope × % abundance of 2nd isotope)

Thus, for the given element the average

$$\text{atomic mass} = \left(9 \times \frac{50}{100}\right) + \left(10 \times \frac{50}{100}\right)$$

$$= 4.5 + 5 = 9.5 \text{ amu}$$

(Note: In the calculations involving percentages, you need to convert percentage abundance into fractional abundance. For example, 50 percent is converted into 50/100 or 0.50 as shown in the a foresaid calculation.)

The atomic masses of elements, given in the periodic table, are average atomic masses. Sometimes, the term atomic weight is used to mean average atomic mass. It is observed, from the periodic table that atomic masses of most of the elements are not whole numbers. For instance, the atomic mass of carbon given in the periodic table is 12.01 amu, not 12.00 amu. The reason is that while calculating the atomic mass of carbon, both of its natural isotopes such as carbon-12. and carbon-13 are considered. The natural abundance of C-12 and C-13 are 98.90 % and 1.10 % respectively. The average of the atomic mass of carbon is calculated as follows:

Average atomic mass of carbon

$$= \left(12 \times \frac{98.9}{100}\right) + \left(13 \times \frac{1.1}{100}\right)$$

$$= (12 \times 0.989) + (13 \times 0.011)$$

$$= 11.868 + 0.143 = 12.011 \text{ amu}$$

So it is important to understand that if it is mentioned that the atomic mass of carbon is 12 amu, it refers to the average atomic mass of the carbon isotopes, not the mass of the individual atoms of carbon.

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Table 7.2 Atomic mass of some elements

Atomic Number	Name	Symbol	Atomic Mass (amu)
1	Hydrogen	H	1.008
2	Helium	He	4.003
3	Lithium	Li	6.941
4	Beryllium	Be	9.012
5	Boron	B	10.811

Calculation of average atomic mass – Solved Examples

Example 1: Oxygen is the most abundant element in both the Earth's crust and the human body. It exists as a mixture of three stable isotopes in nature as shown in Table 7.3:

Table 7.3 Isotopes of oxygen

Isotope	Mass (amu)	% abundance
${}_8O^{16}$	15.9949	99.757
${}_8O^{17}$	16.9991	0.038
${}_8O^{18}$	17.9992	0.205

The atomic mass of

$$\text{oxygen} = (15.9949 \times 0.99757) + (16.9991 \times 0.00038) + (17.9992 \times 0.00205)$$
$$= 15.999 \text{ amu.}$$

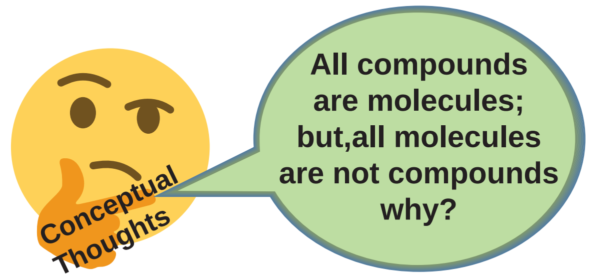
Example 2: Boron naturally occurs as a mixture of boron-10 (5 protons + 5 neutrons) and boron-11 (5 protons + 6 neutrons) isotopes. The percentage abundance of B-10 is 20 and that of B-11 is 80. Then, the atomic mass of boron is calculated as follows:

Atomic mass of

$$\text{boron} = \left(10 \times \frac{20}{100}\right) + \left(11 \times \frac{80}{100}\right)$$
$$= (10 \times 0.20) + (11 \times 0.80)$$
$$= 2 + 8.8$$
$$= 10.8 \text{ amu}$$

7.2 MOLECULE AND MOLECULAR MASS

Except noble gases, atoms of most of the elements are found in the combined form with itself or atoms of other elements. It is called as a molecule. **A molecule is a combination of two or more atoms held together by strong chemical forces of attraction, i.e. chemical bonds.**



7.2.1 Classification of molecules

A molecule may contain atoms of the same element or may contain atoms of two or more elements joined in a fixed ratio, in accordance with the law of definite proportions. Thus, a molecule may be an **element or a compound**. If the molecule is made of similar kind of atoms, then it is called **homoatomic molecule**.

The molecule that consist of atoms of different elements is called **heteroatomic molecule.** A compound is a heteroatomic molecule. The number of atoms present in the molecule is called its '**atomicity**'.

Table 7.4 Classification of molecules

Atomicity	No. of atoms present	Name
1	1	Monoatomic
2	2	Diatomic
3	3	Triatomic
>3	>3	Polyatomic

Let us consider oxygen. Oxygen gas exists in two allotropic forms: Oxygen (O_2) and Ozone (O_3). In oxygen molecule, there are two oxygen atoms. So its atomicity is two. Since both the atoms are similar, oxygen (O_2) is a homodiatomic molecule. Other elements

Activity 7.1

Complete the following table by filling the appropriate values /terms

Element	No. of Protons	No. of Neutrons	Mass Number	Stable Isotopes (abundance)	Atomic Mass (amu)
	7			N-14 (99.6 %)	
		8		N-15 (0.4 %)	
Silicon	**14**		**28**	**Si-28 (92.2 %)**	
	14			**Si-29 (4.7 %)**	
		16		**Si-30 (3.1 %)**	
	17			Cl-35 (75 %)	
	17			Cl-37 (25 %)	

that exist as diatomic molecules are hydrogen (H_2), nitrogen (N_2) and halogens: fluorine (F_2), chlorine (Cl_2), bromine (Br_2) and iodine (I_2).

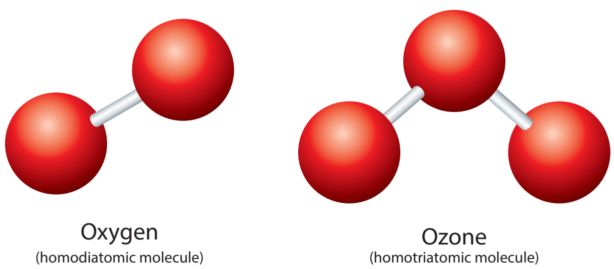


Figure 7.1 Homoatomic molecules

Ozone (O_3) contains three oxygen atoms and hence it is called homotriatomic molecule. If a molecule contains more than three atoms, then it is called **polyatomic molecule**.

Conisder hydrogen chloride. It consists of two atoms, but of different elements, i.e. hydrogen and chlorine. So, its atomicity is two. It is a heterodiatomic molecule. Similarly, the water molecule contains two hydrogen atoms and one oxygen atom. So its atomicity is three. It is a heterotriatomic molecule.

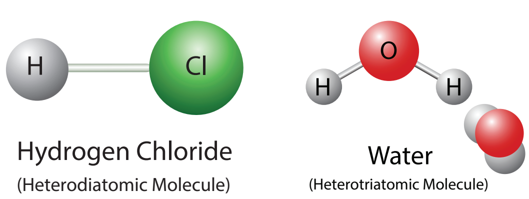


Figure 7.2 Heteroatomic molecules

Activity 7.2

Classify the following molecules based on their atomicity and fill in the table:

Fluorine (F_2), Carbon dioxide (CO_2), Phosphorous (P_4), Sulphur (S_8), Ammonia (NH_3), Hydrogen iodide (HI), Sulphuric Acid (H_2SO_4), Methane (CH_4), Glucose ($C_6H_{12}O_6$), Carbon monoxide (CO)

Molecule	Di - atomic	Tri - atomic	Poly - atomic
Homo			
Hetero			

7.2.2 Relative Molecular Mass (RMM)

As the molecules are made of atoms, they also have their own mass. The mass of the molecule of an element or compound is measured in the C-12 scale and hence called relative molecular mass.

The Relative Molecular Mass of a molecule is the ratio between the mass of one molecule of the substance to $\frac{1}{12^{th}}$ mass of an atom of Carbon -12.

DO YOU KNOW?

Relative Molecular Mass is only a ratio. So, it has no unit. If the molecular mass of a compound is expressed in grams, it is called ***Gram Molecular Mass.***

Gram Molecular Mass of water	=	18 g
Gram Molecular Mass of carbon dioxide	=	44 g
Gram Molecular Mass of ammonia	=	17 g
Gram Molecular Mass of HCl	=	36.5 g

The relative molecular mass is obtained by adding together the relative atomic masses of all the atoms present in a molecule.

Calculation of relative molecular mass – Solved examples:

Example 1: Relative molecular mass of sulphuric acid (H_2SO_4) is calculated as follows: Sulphuric acid conatins 2 atoms of hydrogen, 1 atom of sulphur and 4 atoms of oxygen.

Therefore, Relative molecular mass of sulphuric acid = (2 × mass of hydrogen) + (1 × mass of sulphur) + (4 × mass of oxygen)

$$= (2 \times 1) + (1 \times 32) + (4 \times 16)$$

$$= 98$$

i.e., one molecule of H_2SO_4 is 98 times as heavy as $\frac{1}{12^{th}}$ of the mass of a carbon –12.

Example 2: Relative molecular mass of water (H_2O) is calculated as follows: A water molecule is made of 2 atoms of hydrogen and one atom of oxygen.

So, the relative molecular mass of water

= (2 × mass of hydrogen) + (1 × mass of oxygen)

$$= (2 \times 1) + (1 \times 16)$$

$$= 18$$

i.e., one molecule of H_2O is 18 times as heavy as $\frac{1}{12^{th}}$ of the mass of a carbon –12.

7.3 DIFFERENCE BETWEEN ATOMS AND MOLECULES

Even though atoms are the basic components of molecules, they differ in many aspects when compared to the molecules. Table 7.5 consolidates the major difference between atoms and molecules.

Table 7.5 Difference between atoms and molecules

Atom	Molecule
An atom is the smallest particle of an element	**A molecule is the smallest particle of an element or compound.**
Atom does not exist in free state except in noble gas	**Molecule exists in a free a state**
Except some of noble gas, other atoms are highly reactive	**Molecules are less reactive**
Atom does not have a chemical bond	**Atoms in a molecule are held by chemical bonds**

7.4 MOLE CONCEPT

So far we discussed about matters in terms of individual atoms and molecules. Atomic mass units provide a relative scale for the masses of the elements. Since the atoms have such small masses, no usable scale can be devised to weigh them in the calibrated units of atomic mass units. In any real situation, we deal with macroscopic samples containing enormous number of atoms. Therefore, it is convenient to have a special unit to describe a very large number of atoms. The idea of a 'unit' to denote a particular number of objects is not new. For example, **the pair (2 items) and the dozen (12 items)**, are all familiar units. Chemists measure atoms and molecules in 'moles'. So, you can now understand that 'mole' denotes a number of particles.

In the SI system, the ***mole (mol)* is *the amount of a substance that contains as many elementary entities (atoms, molecules, or other particles) as there are atoms in exactly 12 g (or 0.012 kg) of the carbon-12 isotope.*** The actual number of atoms in 12 g of carbon-12 is determined experimentally. This is called **Avogadro's Number (N_A),** named after an Italian scientist **Amedeo Avogadro** who proposed its significance. Its value is **6.023×10^{23}**. So one mole of a substance contains **6.023×10^{23}** entities. Thus, 5 moles of oxygen molecules contain $5 \times 6.023 \times 10^{23}$ molecules.

Mole Concept: The study of the collection of particles by using mole as the counting unit, in order to express the mass and volume of such unit particles in a bulk of matter is known as **mole concept**.

The number of moles of a substance can be calculated by various means depending on the data available, as follows:

- Number of moles of molecules.
- Number of moles of atoms.
- Number of moles of a gas (Standard molar volume at STP = 22.4 litre).
- Number of moles of ions.

Note:

STP-Standard Temperature and Pressure(273.15 K,1.00 atm)

Mole of atoms:

One mole of an element contains 6.023×10^{23} atoms and it is equal to its gram atomic mass.

i.e., one mole of oxygen atom contains 6.023×10^{23} atoms of oxygen and its gram atomic mass is 16 g.

Mole of molecules:

One mole of matter contains 6.023×10^{23} molecules and it is equal to its gram molecular mass.

i.e., one mole of oxygen molecule contains 6.023×10^{23} molecules of oxygen and its gram molecular mass is 32 g.

Molar volume:

One mole of any gas occupies 22.4 litre or 22400 ml at S.T.P. This volume is called as molar volume.

Calculation of number of moles by Different modes

Number of moles

= Mass / Atomic Mass

= Mass / Molecular mass

= Number of Atoms / 6.023×10^{23}

= Number of Molecules / 6.023×10^{23}

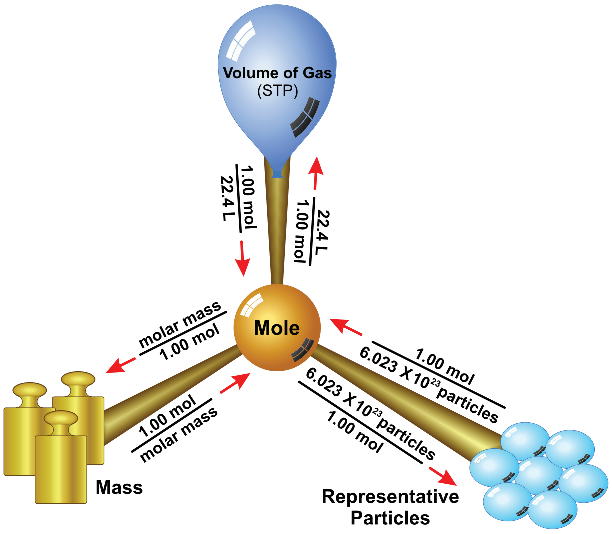


Figure 7.3 Mole concept

7.5 PERCENT COMPOSITION

So for, we were dealing with the number of entities present in a given substance. But many times, the information regarding the percentage of a particular element present in a compound is required.

The percentage composition of a compound represents the mass of each element present in 100 g of the compound.

Let us understand the percentage composition of oxygen and hydrogen by taking the example of H_2O. It can be calculated using the formula

Mass % of an element

$$= \frac{\textit{mass of that element in the compound}}{\textit{molecular mass of the compound}} \times 100$$

Now,

molecular mass of $H_2O = 2(1) + 16$

$= 18$ g

$$\text{Mass \% of hydrogen} = \frac{2}{18} \times 100$$

$= 11.11$ %

$$\text{Mass \% of oxygen} = \frac{16}{18} \times 100$$

$= 88.89$ %

This percentage composition is useful to determine the empirical formula and molecular formula.

Example 1: Find the mass percentage composition of methane (CH_4).

molecular mass of $CH_4 = 12 + 4$

$= 16$ g

$$\text{Mass \% of carbon} = \frac{12}{16} \times 100$$

$= 75$ %

$$\text{Mass \% of hydrogen} = \frac{4}{16} \times 100$$

$= 25$ %

7.6 AVOGADRO HYPOTHESIS

In 1811 Avogadro framed a hypothesis based on the relationship between the number of molecules present in equal volumes of gases in different conditions.

The Avogadro's law states that ***"equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules"***

It follows that the volume of any given gas must be proportional to the number of molecules in it. If 'V' is the volume and 'n' is the number of molecules of a gas, then Avogadro law is represented, mathematically, as follows:

$V \propto n$

$V = \text{constant} \times n$

Thus, one litre (1 dm^3) of hydrogen contains the same number of molecules as in one litre of oxygen, i.e. the volume of the gas is directly proportional to the number of molecules of the gas.

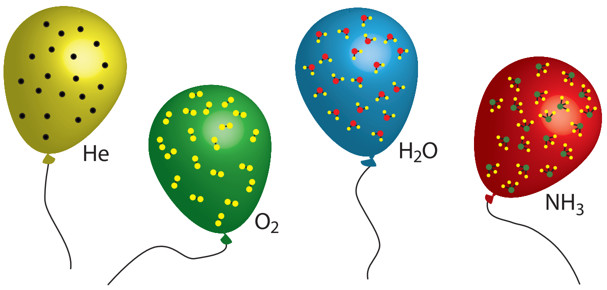


Figure 7.4 Avogadro Hypothesis

Explanation

Let us consider the reaction between hydrogen and chlorine to form hydrogen chloride gas

$H_{2(g)} + Cl_{2(g)} \rightarrow 2\,HCl_{(g)}$

1 vol + 1 vol → 2 volumes

According to Avogadro's law 1 volume of any gas is occupied by "n" number of molecules.

n molecules + n molecules → 2n molecules

if n = 1 then

1molecule + 1 molecule → 2 molecules.

½ molecule + ½ molecule → 1 molecule

1 molecule of hydrogen chloride gas is made up of ½ molecule of hydrogen and ½ molecule of chlorine. Hence, the molecules can be subdivided. This law is in agreement with Dalton's atomic theory.

Activity 7.3

Under same conditions of temperature and pressure if you collect 3 litre of O_2, 5 litre of Cl_2 and 6 litre of H_2,

i. Which has the highest number of molecules?

ii. Which has the lowest number of molecules?

7.7 APPLICATIONS OF AVOGADRO'S LAW

i. It explains Gay-Lussac's law.

ii. It helps in the determination of atomicity of gases.

iii. Molecular formula of gases can be derived using Avogadro's law

iv. It determines the relation between molecular mass and vapour density.

v. It helps to determine gram molar volume of all gases (i.e, 22.4 litre at S.T.P)

7.8 RELATIONSHIP BETWEEN VAPOUR DENSITY AND RELATIVE MOLECULAR MASS

i. Relative molecular mass: (Hydrogen scale)

The Relative Molecular Mass of a gas or vapour is the ratio between the mass of one molecule of the gas or vapour to mass of one atom of Hydrogen.

ii. Vapour Density:

Vapour density is the ratio of the mass of a certain volume of a gas or vapour, to the mass of an equal volume of hydrogen, measured under the same conditions of temperature and pressure.

Vapour Density (V.D.)

$$= \frac{\text{Mass of a given volume of gas or vapour at S.T.P.}}{\text{Mass of the same volume of hydrogen}}$$

According to Avogadro's law, equal volumes of all gases contain equal number of molecules.

Thus, let the number of molecules in one volume = n, then

V.D. at S.T.P

$$= \frac{\text{Mass of 'n' molecules of a gas or vapour at S.T.P.}}{\text{Mass of 'n' molecules of hydrogen}}$$

Cancelling 'n' which is common, you get

V.D.

$$= \frac{\text{Mass of 1 molecule of a gas or vapour at S.T.P.}}{\text{Mass of 1 molecules of hydrogen}}$$

However, since hydrogen is diatomic

V.D.

$$= \frac{\text{Mass of 1 molecule of a gas or vapour at S.T.P.}}{\text{Mass of 2 atoms of hydrogen}}$$

When you compare the formula of vapour density with relative molecular mass, they can be represented as

V.D.

$$= \frac{\text{Mass of 1 molecule of a gas or vapour at S.T.P.}}{2 \times \text{Mass of 1 atom of hydrogen}}$$

(Eqn 7.1)

Relative molecular mass (hydrogen scale) =

$$\frac{\text{Mass of 1 molecule of a gas or vapour at STP}}{\text{Mass of 1 atom of hydrogen}}$$

(Eqn 7.2)

You can therefore substitute the above equation to an Eqn 7.1 and arrive at the following formula

$$\text{V.D.} = \frac{\text{Relative molecular mass}}{2}$$

Now on cross multiplication, you have

2 × vapour density = Relative molecular mass of a gas

(Or)

Relative molecular mass = 2 × Vapour density

7.9 SOLVED PROBLEMS

I. Calculation of molecular mass

Calculate the gram molecular mass of the following.

1) H_2O 2) CO_2 3) $Ca_3(PO_4)_2$

Solution:

1) H_2O

Atomic masses of H = 1, O = 16

Gram molecular mass of H_2O

$= (1 \times 2) + (16 \times 1)$

$= 2 + 16$

Gram molecular mass of H_2O = 18 g

2) CO_2

Atomic masses of C = 12, O = 16

Gram molecular mass of CO_2

$= (12 \times 1) + (16 \times 2)$

$= 12 + 32$

Gram molecular mass of CO_2 = 44 g

3) $Ca_3(PO_4)_2$

Atomic masses of Ca = 40, P = 30, O = 16.

Gram molecular mass of $Ca_3(PO_4)_2$

$= (40 \times 3) + [30 + (16 \times 4)] \times 2$

$= 120 + (94 \times 2)$

$= 120 + 188$

Gram molecular mass of $Ca_3(PO_4)_2$ = 308 g

II. Calculation based on number of moles from mass and volume

1) Calculate the number of moles in 46 g of sodium?

$$\text{Number of moles} = \frac{\text{Mass of the element}}{\text{Atomic mass of the element}}$$

$= 46 / 23$

$= 2$ moles of sodium

2) 5.6 litre of oxygen at S.T.P

$$\text{Number moles} = \frac{\text{Given volume of } O_2 \text{ at S.T.P}}{\text{Molar volume at S.T.P}}$$

$$\text{Number of moles of oxygen} = \frac{5.6}{22.4}$$

$= 0.25$ mole of oxygen

3) Calculate the number of moles of a sample that contains 12.046×10^{23} atoms of iron ?

$$\text{Number of moles} = \frac{\text{Number of atoms of iron}}{\text{Avogadro's number}}$$

$= 12.046 \times 10^{23} / 6.023 \times 10^{23}$

$= 2$ moles of iron

III. Calculation of mass from mole

Calculate the mass of the following

1) 0.3 mole of aluminium (Atomic mass of Al = 27)

$$\text{Number of moles} = \frac{\text{Mass of Al}}{\text{Atomic mass of Al}}$$

Mass = No. of moles × atomic mass

So, mass of Al = 0.3×27

$= 8.1$ g

2) 2.24 litre of SO_2 gas at S.T.P

Molecular mass of $SO_2 = 32 + (16 \times 2)$

$= 32 + 32 = 64$

$$\text{Number of moles of } SO_2 = \frac{\text{Given volume of } SO_2 \text{ at S.T.P}}{\text{Molar volume } SO_2 \text{ at S.T.P}}$$

$$\text{Number of moles of } SO_2 = \frac{2.24}{22.4}$$

$= 0.1$ mole

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molecular mass}}$$

Mass = No. of moles × molecular mass

Mass = 0.1×64

Mass of SO_2 = 6.4 g

3) 1.51×10^{23} molecules of water

Molecular mass of H_2O = 18

$$\text{Number of moles} = \frac{\text{Number of molecules of water}}{\text{Avogadro's number}}$$

$= 1.51 \times 10^{23} / 6.023 \times 10^{23}$

$= 1 / 4$

$= 0.25$ mole

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molecular mass}}$$

0.25 = mass / 18

Mass = 0.25×18

Mass = 4.5 g

4) 5×10^{23} molecules of glucose ?

Molecular mass of glucose = 180

$$\text{Mass of glucose} = \frac{\text{Molecular mass} \times \text{number of particles}}{\text{Avogadro's number}}$$

$= (180 \times 5 \times 10^{23}) / 6.023 \times 10^{23}$

$= 149.43$ g

IV. Calculation based on number of atoms/ molecules.

1) Calculate the number of molecules in 11.2 litre of CO_2 at S.T.P

$$\text{Number of moles of } CO_2 = \frac{\text{Volume at S.T.P}}{\text{Molar volume}}$$

$= 11.2 / 22.4$

$= 0.5$ mole

Number of molecules of CO_2 = number of moles of CO_2 × Avogadro's number

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

$= 3.011 \times 10^{23}$ molecules of CO_2

2) Calculate the number of atoms present in 1 gram of gold (Atomic mass of Au = 198)

$$\text{Number of atoms of Au} = \frac{\text{Mass of Au} \times \text{Avogadro's number}}{\text{Atomic mass of Au}}$$

$$\text{Number of atoms of Au} = \frac{1}{198} \times 6.023 \times 10^{23}$$

Number of atoms of Au $= 3.042 \times 10^{21}$ g

3) Calculate the number of molecules in 54 gm of H_2O?

$$\text{Number of molecules} = \frac{(\text{Avogadro number} \times \text{Given mass})}{\text{Gram molecular mass}}$$

Number of molecules of water $= 6.023 \times 10^{23} \times 54 \ / \ 18$

$= 18.069 \times 10^{23}$ molecules

4) Calculate the number of atoms of oxygen and carbon in 5 moles of CO_2.

- 1 mole of CO_2 contains 2 moles of oxygen
- 5 moles of CO_2 contain 10 moles of oxygen

Number of atoms of oxygen = Number of moles of oxygen × Avogadro's number

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

$= 6.023 \times 10^{24}$ atoms of Oxygen

- 1 mole of CO_2 contains 1 mole of carbon
- 5 moles of CO_2 contains 5 moles of carbon

No. of atoms of carbon = No.of moles of carbon × Avogadro's number

$= 5 \times 6.023 \times 10^{23}$

$= 3.011 \times 10^{24}$ atoms of Carbon

V. Calculation based on molar volume

Calculate the volume occupied by:

1) 2.5 mole of CO_2 at S.T.P

$$\text{Number of moles of } CO_2 = \frac{\text{Given volume at S.T.P}}{\text{Molar volume at S.T.P}}$$

$$2.5 \text{ mole of } CO_2 = \frac{\text{Volume of } CO_2 \text{ at S.T.P}}{22.4}$$

Volume of CO_2 at S.T.P $= 22.4 \times 2.5$

$= 56$ litres.

2) 12.046×10^{23} of ammonia gas molecules

$$\text{Number of moles} = \frac{\text{Number of molecules}}{\text{Avogadro's number}}$$

$= 12.046 \times 10^{23} / 6.023 \times 10^{23}$

$= 2$ moles

Volume occupied by NH_3

= number of moles × molar volume

$= 2 \times 22.4$

$= 44.8$ litres at S.T.P

3) 14 g nitrogen gas

Number of moles $= 14 / 28$

$= 0.5$ mole

Volume occupied by N_2 at S.T.P

= no. of moles × molar volume

$= 0.5 \times 22.4$

$= 11.2$ litres.

VI. Calculation based on % composition

Calculate % of S in H_2SO_4

molecular mass of H_2SO_4

$$= (1 \times 2) + (32 \times 1) + (16 \times 4)$$
$$= 2 + 32 + 64$$
$$= 98 \text{ g}$$

$$\% \text{ of S in } H_2SO_4 = \frac{\text{Mass of sulphur}}{\text{Molecular mass of } H_2SO_4} \times 100$$

$$\% \text{ of S in } H_2SO_4 = \frac{32}{98} \times 100$$
$$= 32.65\ \%$$

Points to Remember

- ❖ Two or more forms of an element having the same atomic number, but different mass number are called Isotopes ($_{17}Cl^{35}$, $_{17}Cl^{37}$).
- ❖ Atoms of different elements having the same mass number, but different atomic numbers are called Isobars ($_{18}Ar^{40}$, $_{20}Ca^{40}$).
- ❖ Atoms of different elements having the same number of neutrons, but different atomic number and different mass number are called Isotones ($_{6}C^{13}$, $_{7}N^{14}$).
- ❖ Relative atomic mass of an element is the ratio between the mass of one atom of the element to 1/12th of the mass of the atom of carbon -12.
- ❖ Average atomic mass of an element is calculated by adding the masses of its isotopes, each multiplied by their natural abundance on the Earth.
- ❖ Relative molecular mass of a molecule is the ratio between the mass of one molecule of the substance to 1/12th of the mass of the atom of carbon – 12.
- ❖ The Avogadro's law states that "equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules".
- ❖ The vapour density is defined as "the ratio between the masses of equal volumes of a gas (or a vapour) and hydrogen under the same condition".
- ❖ Atomicity of a monoatomic element = Molecular mass / Atomic Mass.
- ❖ Molecular mass = 2 × Vapour density.

DLRFYB

TEXTBOOK EVALUATION

I. Choose the best answer.

1. Which of the following has the smallest mass?
   a. 6.023×10^{23} atoms of He
   b. 1 atom of He
   c. 2 g of He
   d. 1 mole atoms of He

2. Which of the following is a triatomic molecule?
   a. Glucose b. Helium
   c. Carbon dioxide d. Hydrogen

3. The volume occupied by 4.4 g of CO_2 at S.T.P
   a. 22.4 litre b. 2.24 litre
   c. 0.24 litre d. 0.1 litre

4. Mass of 1 mole of Nitrogen atom is
   a. 28 amu b. 14 amu
   c. 28 g d. 14 g

5. Which of the following represents 1 amu?
   a. Mass of a C – 12 atom
   b. Mass of a hydrogen atom
   c. 1/12th of the mass of a C – 12 atom
   d. Mass of O – 16 atom

6. Which of the following statement is incorrect?
   a. 12 gram of C – 12 contains Avogadro's number of atoms.
   b. One mole of oxygen gas contains Avogadro's number of molecules.
   c. One mole of hydrogen gas contains Avogadro's number of atoms.
   d. One mole of electrons stands for 6.023×10^{23} electrons.

7. The volume occupied by 1 mole of a diatomic gas at S.T.P is
   a. 11.2 litre   b. 5.6 litre
   c. 22.4 litre   d. 44.8 litre

8. In the nucleus of ${}_{20}Ca^{40}$, there are
   a. 20 protons and 40 neutrons
   b. 20 protons and 20 neutrons
   c. 20 protons and 40 electrons
   d. 40 protons and 20 electrons

9. The gram molecular mass of oxygen molecule is
   a. 16 g   b. 18 g
   c. 32 g   d. 17 g

10. 1 mole of any substance contains ____ molecules.
    a. 6.023×10^{23} b. 6.023×10^{-23}
    c. 3.0115×10^{23} d. 12.046×10^{23}

II. Fill in the blanks

1. Atoms of different elements having ________ mass number, but _________ atomic numbers are called isobars.
2. Atoms of different elements having same number of ____________ are called isotones.
3. Atoms of one element can be transmuted into atoms of other element by _________
4. The sum of the numbers of protons and neutrons of an atom is called its ___________
5. Relative atomic mass is otherwise known as ___________
6. The average atomic mass of hydrogen is ____________ amu.
7. If a molecule is made of similar kind of atoms, then it is called _____________ atomic molecule.
8. The number of atoms present in a molecule is called its _____________
9. One mole of any gas occupies _________ ml at S.T.P
10 Atomicity of phosphorous is ____________

III. Match the following

1.	8 g of O_2	-	4 moles
2.	4 g of H_2	-	0.25 moles
3.	52 g of He	-	2 moles
4.	112 g of N_2	-	0.5 moles
5.	35.5 g of Cl_2	-	13 moles

IV. True or False: (If false give the correct statement)

1. Two elements sometimes can form more than one compound.
2. Noble gases are Diatomic
3. The gram atomic mass of an element has no unit
4. 1 mole of Gold and Silver contain same number of atoms
5. Molar mass of CO_2 is 42g.

V. Assertion and Reason:

Answer the following questions using the data given below:

i) A and R are correct, R explains the A.

ii) A is correct, R is wrong.

iii) A is wrong, R is correct.

iv) A and R are correct, R doesn't explains A.

1. Assertion: The Relative Atomic mass of aluminium is 27

Reason: An atom of aluminium is 27 times heavier than 1/12th of the mass of the C – 12 atom.

2. **Assertion:** The Relative Molecular Mass of Chlorine is 35.5 a.m.u.

   Reason: The natural abundance of Chlorine isotopes are not equal.

VI. Short answer questions

1. Define: Relative atomic mass.
2. Write the different types of isotopes of oxygen and its percentage abundance.
3. Define: Atomicity
4. Give any two examples for heterodiatomic molecules.
5. What is Molar volume of a gas?
6. Find the percentage of nitrogen in ammonia.

VII. Long answer questions

1. Calculate the number of water molecule present in one drop of water which weighs 0.18 g.

2. $N_2 + 3\ H_2 \rightarrow 2\ NH_3$

   (The atomic mass of nitrogen is 14, and that of hydrogen is 1)

   1 mole of nitrogen (________g) +
   3 moles of hydrogen (__________ g) →

   2 moles of ammonia (__________ g)

3. Calculate the number of moles in

   i) 27g of Al ii) 1.51×10^{23} molecules of NH_4Cl

4. Give the salient features of "Modern atomic theory".
5. Derive the relationship between Relative molecular mass and Vapour density.

VIII. HOT question

1. Calculate carbonate is decomposed on heating in the following reaction

$$CaCO_3 \rightarrow CaO + CO_2$$

   i. How many moles of Calcium carbonate are involved in this reaction?

   ii. Calculate the gram molecular mass of calcium carbonate involved in this reaction

   iii. How many moles of CO_2 are there in this equation?

IX. Solve the following problems

1. How many grams are there in the following?

   i. 2 moles of hydrogen molecule, H_2

   ii. 3 moles of chlorine molecule, Cl_2

   iii. 5 moles of sulphur molecule, S_8

   iv. 4 moles of phosphorous molecule, P_4

2. Calculate the % of each element in calcium carbonate. (Atomic mass: C-12, O-16, Ca -40)

3. Calculate the % of oxygen in $Al_2(SO_4)_3$. (Atomic mass: Al-27, O-16, S -32)

4. Calculate the % relative abundance of B-10 and B-11, if its average atomic mass is 10.804 amu.

REFERENCE BOOKS

1) Petrucci, Ralph H et.al. General Chemistry: Principles & Modern Applications (9th Edition). Upper Saddle River, NJ: Pearson Prentice Hall, 2007. Print.

2) Raymond Chang. (2010). Chemistry. New York, NY: The Tata McGraw Hill Companies. Inc.

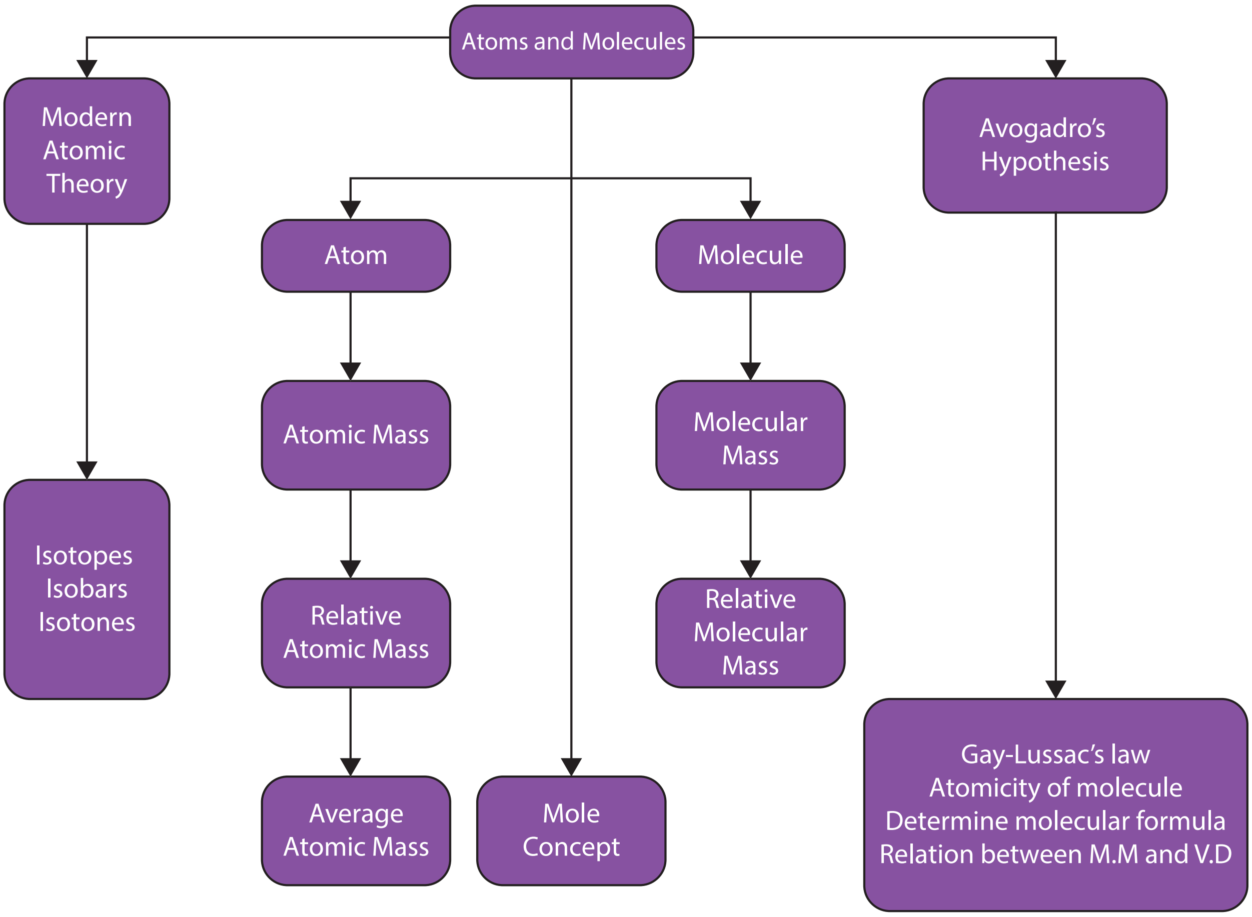
3) Julia Burdge. (2011). Chemistry. New York, NY: The Tata McGraw Hill Companies.Inc.

INTERNET RESOURCES

https://www2.estrellamountain.edu/faculty/farabee/biobk/BioBookCHEM1.html

https://www.toppr.com/guides/chemistry/atoms-and-molecules/

CONCEPT MAP


Atoms and Molecules
Modern Atomic Theory
Avogadro's Hypothesis
Atom
Molecule
Atomic Mass
Molecular Mass
Isotopes Isobars Isotones
Relative Atomic Mass
Relative Molecular Mass
Gay-Lussac's law
Atomicity of molecule
Determine molecular formula
Relation between M.M and V.D
Average Atomic Mass
Mole Concept