

## Chapter 1

Element	Berzelius value	Current value
Hydrogen	1.0000	1.0079
Carbon	12.0000	12.011
Oxygen	16.0000	15.999
Iron	55.845	55.845

## BASIC CONCEPTS

## HISTORICAL BACKGROUND OF ATOM

## Greek philosophers

Greek philosophers thought that matter could be divided into smaller and smaller particles to reach a basic unit which could not be further sub-divided. **Democritus** named these smallest indivisible particles as atoms derived from "atomos" which means indivisible. These ideas of Greek philosophers were not based on experimental evidences.

17<sup>th</sup> Century work

In the late 17<sup>th</sup> century, the quantitative study of the composition of pure substances disclosed that a few elements were the components of many difficult substances. It was also investigated that how elements combined to form compounds and how compounds could be broken down into their constituent elements.

## Dalton's work

In 1808, an English school teacher, John Dalton recognized that law of conservation of mass and law of definite proportions could only be explained by the existence of atoms. He developed a theory about atom called **Dalton's Atomic Theory**. The main postulate of atomic theory is that all matter is composed of atoms of different elements, which differ in their properties.

## Atom

The smallest particle of an element which can take part in a chemical reaction is called **atom**.

or

The smallest particle of an element which may or may not exist independently is called **atom**.

## Examples

Atoms of **He, Ne, Ar, Kr, Xe** and **Rn** can exist independently while atoms of **H, O, N** etc. do not exist independently.

## Sub-atomic particles

According to modern researches, atom is composed of sub-atomic particles like electron, proton, neutron, hypron, boson, neutrino, antineutrino etc. More than 100 such particles are thought to exist in an atom. However, electron, proton and neutron are regarded as fundamental particles of atoms.

## Berzelius's work

Swedish Chemist J. Berzelius (1779 – 1848) has following contribution in chemistry

- He determined the atomic masses of elements. A number of his values are close to the modern values of atomic masses.
- He developed the system of giving element a symbol.

Dalton's atomic theory started chemistry on the road from a branch of philosophy to the science which it is today.

## Element

A substance consisting of atoms which all have the same number of protons i.e. the same atomic number. Elements are chemically the simplest substances and hence cannot be broken down further using chemical methods. Elements can only be changed into other elements using nuclear methods.

J. Berzelius (Best experimental chemist) performed more than 2000 experiments over a 10 years period to determine atomic masses for 50 elements then known.



## Disclaimer

This Blog/Web Site is made available by the lawyer or law firm publisher for educational purpose only as well as to give you general information and a general understanding. We have the Rights to use this document for education purpose. You are not allowed to use this for commercial purpose. It is only for personal use. If you thoughts that this document include something related to you, you can email us at [yAsadBhatti@gmail.com](mailto:yAsadBhatti@gmail.com). We will look up into the matter and if we found anything related to you, we will remove the content from our website.

For Notes, Past Papers, Video Lectures, Education News

Visit Us at:

<https://www.bhattiAcademy.com>

<https://www.youtube.com/bhattiAcademy>

<https://www.facebook.com/bhattiAcademy>

If above links are NOT WORKING contact us at

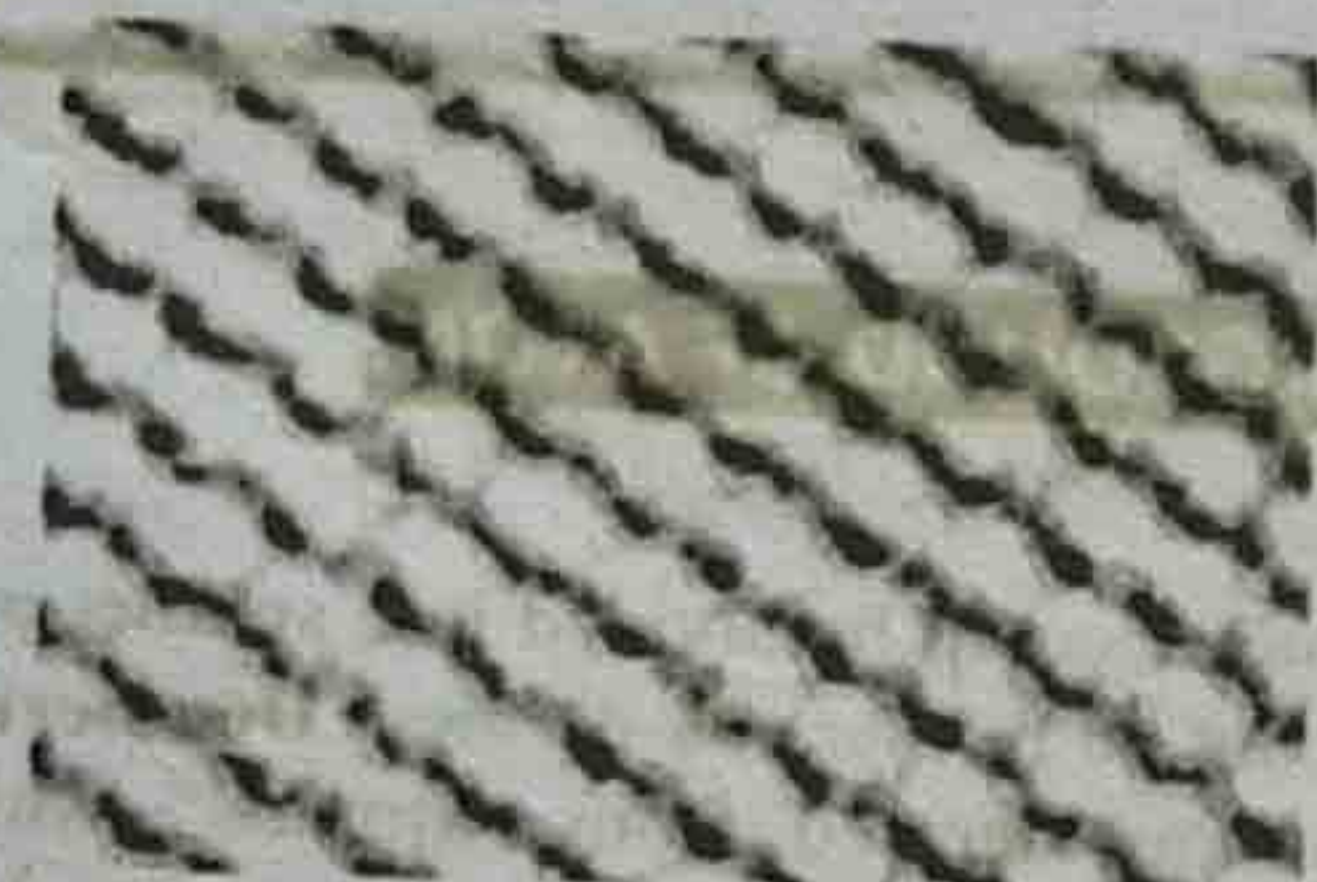
[yAsadBhatti@gmail.com](mailto:yAsadBhatti@gmail.com)



## Comparison of Berzelius's atomic masses with modern values

Element	Berzelius value	Current value
Chlorine	35.41	35.45
Copper	63.00	63.55
Nitrogen	14.05	14.01

Berzelius discovered cerium, thorium, selenium and silicon.



Electron microscopic photograph of graphite

## Evidence of Atoms

It is not possible actually to see the atoms but the nearest possibility to its direct evidence is by using an electron microscope. A clear and accurate image of an object that is smaller than the wavelength of visible light, cannot be obtained.

## Demerit of compound microscope

An ordinary optical microscope can measure the size of an object up to or above 500 nm ( $1 \text{ nm} = 10^{-9} \text{ m}$ ).

## Use of electron microscope

The objects of the size of an atom can be observed in an electron microscope. It uses beams of electrons instead of visible light, because wavelength of electron is much shorter than that of visible light and is most suitable for viewing the atoms.

## Electron microscopic photograph of graphite

The figure shows electron microscopic photograph of a piece of graphite which has been magnified about 15 millions times. The bright band in the figure are layers of carbon atoms.

## Results of X-rays work

In the twentieth century, X-rays work has shown that

- Diameter of atoms are of the order  $2 \times 10^{-10} \text{ m}$  which is 0.2 nm or  $2 \text{ \AA}$ .
- Masses of atoms range from  $10^{-27}$  to  $10^{-25} \text{ kg}$ . They are often expressed in atomic mass units (amu)

$$1 \text{ amu} = 1.661 \times 10^{-27} \text{ kg} (1.661 \times 10^{-24} \text{ g or } 1.661 \times 10^{-21} \text{ mg})$$

## Consequence

This shows that the atoms do exist and they are of an amazingly small size. Even a full stop may have two million atoms present in it.

## Molecule

"The smallest particle of a pure substance (element or compound) which can exist independently is called a molecule."

## Classification of Molecules

On the basis of (i) Nature (ii) Atomicity (iii) Size

## Amazingly Small Atoms:

- If a golf ball is magnified to the size of the earth, then an atom would be the size of marble.
- Take a deep breath; you have just breathed  $10^5$  million atoms.

$$1 \text{ \AA} = 10^{-10} \text{ m}$$

## Scholar's CHEMISTRY - XI (Subjective)

## (i) On the basis of its nature

On the basis of nature, molecules are of two types.

- Homo-atomic molecules
- Hetero-atomic molecules

## (a) Homo-atomic molecule

"A molecule which is composed of same or one kind of atoms is called a homo-atomic molecule or elemental molecule."

Examples:  $\text{O}_2$ ,  $\text{O}_3$ ,  $\text{P}_4$ ,  $\text{S}_8$  etc.

## (b) Hetero-atomic molecule

"A molecule which is composed of different types of atoms is called a hetero-atomic molecule or compound molecule"

Examples:  $\text{CO}$ ,  $\text{SO}_2$ ,  $\text{NH}_3$ ,  $\text{CH}_4$  etc.

## Atomicity

"The number of atoms present in a molecule is called atomicity."

Substance	Molecule	Atomicity
Helium	He	1
Water	$\text{H}_2\text{O}$	3
Glucose	$\text{C}_6\text{H}_{12}\text{O}_6$	24
Sucrose	$\text{C}_{12}\text{H}_{22}\text{O}_{11}$	45

## (ii) On the basis of atomicity

On the basis of atomicity, molecules are of two types.

- Mono-atomic molecule
- Polyatomic molecule

## (a) Mono-atomic molecule

"The molecule which consists of only one atom is called mono-atomic molecule."

Examples Noble gases (He, Ne, Ar, Kr, Xe and Rn) has monoatomic molecules.

## (b) Polyatomic molecule

"A molecule which consists of two or more, same or different kinds of the atoms is called a polyatomic molecule."

Examples  $\text{CO}$ ,  $\text{CO}_2$ ,  $\text{O}_2$ ,  $\text{O}_3$ ,  $\text{C}_6\text{H}_{12}\text{O}_6$  etc.

(It can be diatomic, tri-atomic, tetra-atomic)

## (iii) On the basis of size of molecule

On the basis of size, molecules are of two types.

## (a) Micromolecules

They are small in size. They are simple molecules and exist as monomer.

Examples  $\text{H}_2\text{O}$ ,  $\text{CO}_2$ ,  $\text{C}_6\text{H}_6$  etc.

## (b) Macromolecules

They are large in size having large number of atoms.

Examples Haemoglobin, cellulose, starch, graphite etc.



**Haemoglobin**

- It is a blood protein.
- It transports oxygen from our lungs to all parts of the body.
- Each molecule of haemoglobin is made up of nearly 10,000 atoms.
- It is 68,000 times heavier than a hydrogen atom.
- It contains carbon, hydrogen, nitrogen, oxygen and iron.

**Difference between Atom and Molecule**

ATOM	MOLECULE
1) It is the smallest particle of an element.	1) It is the smallest particle of a pure substance.
2) It is represented by a symbol.	2) It is represented by molecular formula.
3) It shows the properties of element.	3) It shows the properties of the substance.
4) It retains its identity in a chemical reaction.	4) It does not retain its identity in a chemical reaction.
5) It cannot be further sub-divided by ordinary chemical reactions.	5) It can be further sub-divided by ordinary chemical reaction.
6) It may or may not exist in free state.	6) It can exist in free state.

**Ion**

"Those species which carry either positive or negative charge are called ions."

**Types** There are two types of ion

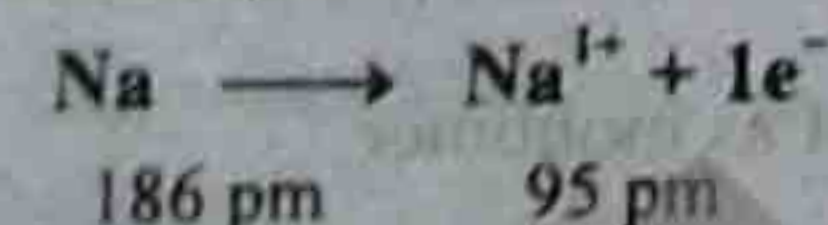
- Positive ion or cation.
- Negative ion or anion.

**(i) Positive Ion**

"It is that ion which is carrying a positive charge. A positive ion is formed when an atom loses one or more electrons."



- A positive ion is also called a cation.
- A positive ion may carry +1, +2 or +3 charge depending upon the number of electrons lost by the atom.
- We have to supply sufficient amount of energy in order to remove an electron from the valence shell of an atom. It means the formation of a positive ion is an endothermic process.
- Formation of cation is an oxidation process.
- Metal atom can easily lose electrons to form positive ions like  $\text{Na}^{+}$ ,  $\text{K}^{+}$ ,  $\text{Ca}^{2+}$  and  $\text{Fe}^{2+}$  ions.
- Size of a cation is smaller than its parent atom.

**(ii) Negative Ion**

"It is that ion which is carrying a negative charge. A negative ion is formed when an atom gains one or more electrons."



- A negative ion is also called an anion.

Q. How many times a He atom is lighter than a haemoglobin molecule?

- 68000
- 34000
- 17000
- 1000

**Cations and Anions:**

During electrolysis, the negative electrode or cathode attracts positive ions, called **cations**. The positive electrode or anode attracts negative ions, called **anions**.

Q. The formation of positive ion is an endothermic process. Justify.

- A negative ion may carry -1, -2 or -3 charge depending upon the number of electrons gained by the atom.
- Energy is released when one electron is added in the valence shell of an isolated neutral atom. It means that the formation of an uni-negative ion is an exothermic process while formation of di-negative ion is an endothermic process.
- Non-metals mostly gain electrons and to form negative ions like  $\text{F}^{-}$ ,  $\text{Cl}^{-}$ ,  $\text{S}^{2-}$  and  $\text{O}^{2-}$  ions etc.
- Formation of anion is a reduction process.
- These cations and anions have entirely different properties from their parent atoms.
- Size of anion is larger than its parent atom.

**Polyatomic Ions**

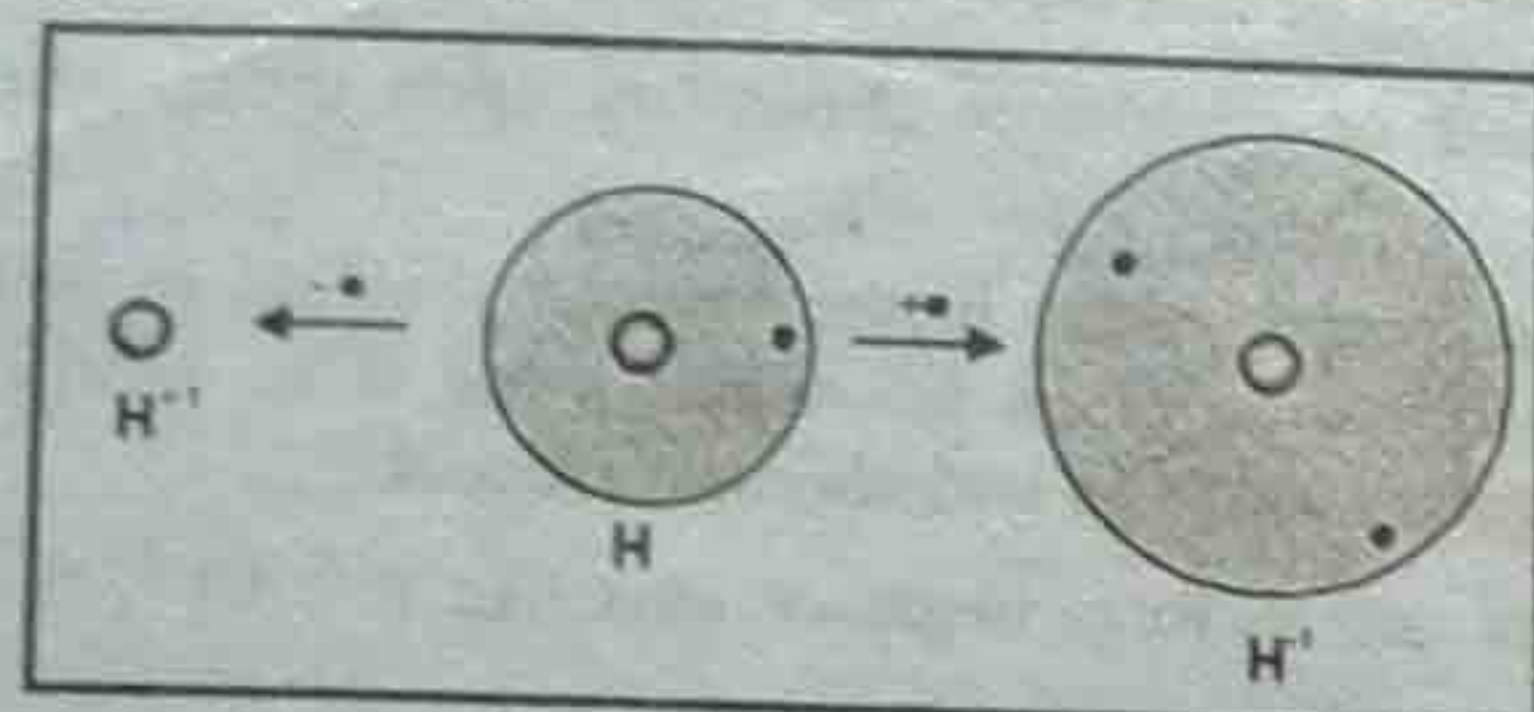
These are the positive or negative ions which consist of group of atoms.

- The positive polyatomic ions are less common. For example,  $\text{NH}_4^{+}$  ions and certain carbocations in organic chemistry.
- There are many examples of negative polyatomic ions like  $\text{OH}^{-}$ ,  $\text{CO}_3^{2-}$  and  $\text{SO}_4^{2-}$  ions  $\text{PO}_4^{3-}$ ,  $\text{MnO}_4^{-}$ ,  $\text{Cr}_2\text{O}_7^{2-}$  etc.

**Difference between Cation and Anion**

Cation (positive ion)	Anion (negative ion)
<ul style="list-style-type: none"> <li>• A cation is formed by the loss of electron or electrons from a neutral atom.</li> </ul> $A \longrightarrow A^{+} + 1e^{-}$	<ul style="list-style-type: none"> <li>• Anion is formed by the gain of electron or electrons by a neutral atom.</li> </ul> $B + 1e^{-} \longrightarrow B^{-}$
<ul style="list-style-type: none"> <li>• Formation of cation is an endothermic process.</li> </ul> $\text{Na}_{(g)} \longrightarrow \text{Na}_{(g)}^{+} + 1e^{-} \quad \Delta H = +496 \text{ kJ mol}^{-1}$	<ul style="list-style-type: none"> <li>• Formation of uninegative ion is an exothermic process whereas the formation of dinegative ion is an endothermic process.</li> </ul> $\text{O}_{(g)} + 1e^{-} \longrightarrow \text{O}_{(g)}^{-} \quad \Delta H = -141 \text{ kJ mol}^{-1}$ $\text{O}_{(g)}^{-} + 1e^{-} \longrightarrow \text{O}_{(g)}^{2-} \quad \Delta H = +780 \text{ kJ mol}^{-1}$
<ul style="list-style-type: none"> <li>• Formation of cation is an oxidation process.</li> </ul>	<ul style="list-style-type: none"> <li>• Formation of anion is a reduction process.</li> </ul>
<ul style="list-style-type: none"> <li>• The size of cation is smaller than its parent atom.</li> </ul> <p>e.g. <math>\text{Na}_{(g)} \longrightarrow \text{Na}_{(g)}^{+} + 1e^{-}</math></p> <p style="text-align: center;">186 pm      95 pm</p>	<ul style="list-style-type: none"> <li>• The size of anion is greater than its parent atom.</li> </ul> <p>e.g. <math>\text{Cl}_{(g)} + 1e^{-} \longrightarrow \text{Cl}_{(g)}^{-}</math></p> <p style="text-align: center;">99 pm      181 pm</p>
<ul style="list-style-type: none"> <li>• The behaviour of neutral atom and cation is different.</li> </ul>	<ul style="list-style-type: none"> <li>• The behaviour of neutral atom and anion is different.</li> </ul>
<ul style="list-style-type: none"> <li>• Usually, electropositive metals lose electrons and form cations.</li> </ul>	<ul style="list-style-type: none"> <li>• Usually, non-metals gain electrons and form anions.</li> </ul>
<ul style="list-style-type: none"> <li>• The charge on cations is equal to number of electrons lost. e.g.,</li> </ul> $\begin{aligned} \text{Na} &\longrightarrow \text{Na}^{+} + 1e^{-} \\ \text{Ca} &\longrightarrow \text{Ca}^{2+} + 2e^{-} \\ \text{Al} &\longrightarrow \text{Al}^{3+} + 3e^{-} \end{aligned}$	<ul style="list-style-type: none"> <li>• The charge on anions is equal to number of electrons gained e.g.,</li> </ul> $\begin{aligned} \text{Cl} + 1e^{-} &\longrightarrow \text{Cl}^{-} \\ \text{O} + 2e^{-} &\longrightarrow \text{O}^{2-} \\ \text{N} + 3e^{-} &\longrightarrow \text{N}^{3-} \end{aligned}$



**Molecular Ion**

"Those ions which are produced by the removal of one or more electron or electrons from the molecule of a substance are called molecular ions."

**Types** There are two types of molecular ion.

**(a) Cationic molecular ions**

They have positive charge. They are more abundant than anionic molecular ions.

**Example**  $\text{N}_2^+$ ,  $\text{CO}^+$ ,  $\text{CH}_4^+$  etc.

**(b) Anionic molecular ions**

They have negative charge. They are less abundant.

**Example**  $\text{O}_2^{2-}$ ,  $\text{N}_3^{2-}$  etc.

**Formation**

These ions can be generated by passing high energy electron beam as  $\alpha$  - particles or X - rays through a gas.

**Application**

The breakdown of molecular ions obtained from the natural products can give important information about their structure.

**Q.4** What are ions? Under what conditions are they produced?

**Ans.** Ions

"Those species which carry either positive or negative charge are called ions."

**Formation of ions**

Ions are formed under following conditions

- By passing different radiations through gaseous mixture.
- By adding the substance into an aqueous solution.
- By heating the substance (in molten state).

**Relative Atomic Mass**

"It is the mass of an atom of an element as compared to the mass of an atom of carbon taken as 12."

- The term atomic mass is preferred over atomic weight because mass is more fundamental unit than weight.
- Atomic mass is a relative term. It tells us how much heavier or lighter an atom of the element is, than an atom of carbon - 12.
- Carbon - 12 is used as a standard because it is stable and exist abundantly.

**Relative Atomic Masses of a Few Elements**

Element	Relative Atomic Mass (amu)	Element	Relative Atomic Mass (amu)
H	1.008	Cl	35.453
O	15.9994	Cu	63.546
Ne	20.1797	U	238.0289

**MCQ's**

Which of the following is not a molecular ion?

- (a)  $\text{CH}_4^+$  (b)  $\text{NH}_4^+$   
(c)  $\text{O}_2^{1-}$  (d)  $\text{NH}_3^+$

**Prove that:**  $1 \text{ amu} = 1.661 \times 10^{-27} \text{ kg}$

**Solution:**

$$6.02 \times 10^{23} \text{ C atom} = 12 \text{ g}$$

$$1 \text{ C atom} = \frac{12}{6.02 \times 10^{23}} \text{ g}$$

$$1 \text{ C atom} = 1.99 \times 10^{-23} \text{ g}$$

$$\frac{1}{12} \text{ of carbon atom} = \frac{1.99 \times 10^{-23}}{12} \text{ g}$$

$$= 1.661 \times 10^{-24} \text{ g}$$

$$= 1.661 \times 10^{-27} \text{ kg}$$

**Atomic Mass Unit**

"1/12th of the mass of one atom of carbon - 12 is called the atomic mass unit. It is abbreviated as amu."

$$1 \text{ amu} = 1.661 \times 10^{-27} \text{ kg} = 1.661 \times 10^{-24} \text{ g} = 1.661 \times 10^{-21} \text{ mg}$$

**Q.** Which one of the following is the relative atomic mass of an element?

- (a)  $\frac{\text{average mass of one atom of the element}}{\text{mass of one atom of } ^1\text{H}}$  (c)  $\frac{\text{average mass of one atom of the element}}{\frac{1}{12} \times \text{average mass of one atom of carbon}}$   
(b)  $\frac{\text{average mass of one atom of the element}}{\text{mass of one atom of } ^{12}\text{C}}$  (d)  $\frac{\text{average of mass of one atom of the element}}{\frac{1}{12} \times \text{mass of one atom of } ^{12}\text{C}}$

**Isotopes**

"Atoms of the same element which have the same atomic number but different mass numbers due to difference in the number of neutrons are called isotopes of that element."

The phenomenon of isotopy was first discovered by Soddy.

**Examples****Isotopes of hydrogen**

Hydrogen consists of three isotopes which are protium ( $^1\text{H}$ ), deuterium ( $^2\text{H}$ ) and tritium ( $^3\text{H}$ ). All these isotopes have the same atomic number i.e. one, but they have different mass numbers 1, 2 and 3 respectively.

**Isotopes of carbon**

Carbon also consists of three isotopes which are C-12 ( $^{12}\text{C}$ ), C-13 ( $^{13}\text{C}$ ) and C-14 ( $^{14}\text{C}$ ). All these isotopes have the same atomic number i.e., six but they have different mass numbers 12, 13 and 14 respectively.

Oxygen has three, nickel has five, calcium has six, palladium has six, cadmium has nine and tin has eleven isotopes. The elements like arsenic, fluorine, iodine and gold etc have only a single isotope. They are called mono - isotopic elements.

**Similarities and Dissimilarities of Isotopes of an Element**

Similarities	Dissimilarities
Isotopes of an element have same	Isotopes of an element have different
• Atomic number	• Atomic mass
• Number of proton	• Number of neutron
• Number of electron	• Radioactive properties due to different composition of nuclei
• Chemical properties due to same electronic configuration	• Physical properties
• Position in modern periodic table	• Half life due to different stabilities

**Relative abundance of isotopes**

"The percentage of each isotope in a mixture of isotopes of an element is called relative abundance."

- Different isotopes have their own natural abundance.
- The relative abundance of isotopes is measured by mass spectrometry.

**Isobars:**

Those atoms which have the same mass numbers but different atomic numbers.  $^{14}\text{C}$ ,  $^{14}\text{N}$ .

**Isoelectronic species:**

Those species i.e. atoms, ions or molecules which have the same number of electrons are called isoelectronic species e.g.  $\text{Ne}$ ,  $\text{Mg}^{2+}$ ,  $\text{O}^{2-}$  are isoelectronic species.

**Isotones:**

Atoms of the different elements having same number of neutrons but different mass numbers are called isotones. e.g.  $^{30}_{14}\text{Si}$ ,  $^{31}_{15}\text{P}$ ,  $^{32}_{16}\text{S}$  are isotones.



- The properties of a particular element, which are mentioned in the literature, mostly correspond to the most abundant isotope of that element.

Natural Abundance of Some Common Isotopes

Element	Isotope	Abundance (%)	Mass (amu)
Hydrogen	$^1\text{H}$ , $^2\text{H}$	99.985, 0.015	1.007825, 2.01410
Carbon	$^{12}\text{C}$ , $^{13}\text{C}$	98.893, 1.107	12.0000, 13.00335
Nitrogen	$^{14}\text{N}$ , $^{15}\text{N}$	99.634, 0.366	14.00307, 15.00011
Oxygen	$^{16}\text{O}$ , $^{17}\text{O}$ , $^{18}\text{O}$	99.759, 0.037, 0.204	15.99491, 16.99914, 17.9916
Sulphur	$^{32}\text{S}$ , $^{33}\text{S}$ , $^{34}\text{S}$ , $^{36}\text{S}$	95.0, 0.76, 4.22, 0.014	31.97207, 32.97146, 33.96786, 35.96709
Chlorine	$^{35}\text{Cl}$ , $^{37}\text{Cl}$	75.53, 24.47	34.96885, 36.96590
Bromine	$^{79}\text{Br}$ , $^{81}\text{Br}$	50.54, 49.49	78.918, 80.916

The distribution of isotopes among the elements varied and complex as it is evident from above table.

### Occurrence of isotopes

At present more than 280 different isotopes occur in nature.

- 40 radioactive isotopes are also included in this number (280).
- Almost 300 unstable radioactive isotopes of different elements have been produced by the artificial disintegration.
- In general elements with odd atomic number almost never possess more than two stable isotopes.
- The elements of even atomic number usually have larger number of isotopes.
- The isotopes whose mass numbers are multiple of four are particularly abundant. For example  $^{16}\text{O}$ ,  $^{24}\text{Mg}$ ,  $^{28}\text{Si}$ ,  $^{40}\text{Ca}$  and  $^{56}\text{Fe}$ . These isotopes exist abundantly and form about 50% of the earth crust.
- Out of 280 isotopes that occur in nature, 154 isotopes have even atomic number and even mass number.

### Monoisotopic Elements

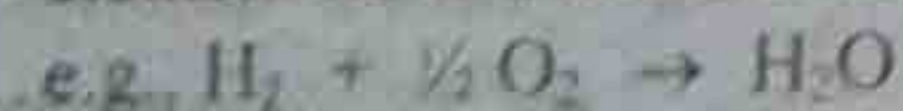
Elements having only one isotope are called mono-isotopic elements.

### Examples:

Arsenic, Fluorine, Iodine, Gold etc have only single isotope.

Q. Why isotopes have identical chemical properties but different physical properties?

Ans. The chemical properties of a substance depends upon its outer shell electronic configuration. Since isotopes of an element have same electronic configuration, so they show similar chemical properties.



Moreover, due to different nuclear composition physical properties of isotopes are different.

## DETERMINATION OF RELATIVE ATOMIC MASS OF ISOTOPES BY MASS SPECTROMETRY

### Mass Spectrometer

"It is an instrument which is used to measure the exact masses of different isotopes of an element along with their relative abundance."

### Principle of mass spectrometry

In this technique, a substance is first volatilized and then ionized with the help of high energy beam of electrons. The gaseous positive ions, thus formed, are separated on the basis of their mass to charge ratio ( $m/e$ ) and then recorded in the form of peaks.

### Aston's mass spectrograph

First mass spectrograph was designed by Aston in 1919. It was designed for the identification of isotopes of an element on the basis of their atomic masses.

### Dempster's mass spectrometer

It was designed for identification of elements which were available in solid state.

Q. Write down principle of mass spectrometry.

## Determination of Relative Atomic Masses of Isotopes by

### Dempster's Mass Spectrometer

Different steps involved in the determination of exact atomic masses and the relative abundances of different isotopes of an element are given below

#### (i) Vapourization

The substance whose analysis for the separation of isotopes is required, is converted into the vapour state. The pressure of these vapours is kept very low, that is,  $10^{-6}$  to  $10^{-7}$  torr.

#### (ii) Ionization

These vapours are then allowed to enter the ionization chamber where fast moving electrons are thrown upon them. The atoms of isotopic element present in the form of vapours, are ionized. These positively charged ions of isotopes of an element have different masses depending upon the nature of the isotopes present in them.

#### (iii) Acceleration

The positive ions of each isotope has its own ( $m/e$ ) value. When a potential difference ( $E$ ) of 500-2000 volts is applied between perforated accelerating plates, then these positive ions are strongly attracted towards the negative plate. In this way, the ions are accelerated.

#### (iv) Deflection

The beam of accelerated positive ions is then allowed to pass through a strong magnetic field of the strength  $H$ . This magnetic field is applied in a direction which is perpendicular to the path of the positive ions. The applied magnetic field will help us in the separation of positive ions on the basis of their  $m/e$  values. The magnetic field makes the ions to move in a circular path. The ions of definite  $m/e$  value will move in the form of groups one after the other and fall on the electrometer.

#### (v) Mathematical explanation

The mathematical relationship between  $m/e$  values and deflection in the circular path is

$$m/e = H^2 r / E$$

Where,  $H$  = strength of magnetic field

$E$  = Strength of electric field

$r$  = Radius of circular path

If  $E$  is increased, by keeping  $H$  constant then radius will increase and positive ion of a particular  $m/e$  will fall at a different place as compared to the first place. This can also be done by changing the magnetic field. Smaller the ( $m/e$ ) of an isotope, smaller the radius of curvature produced by the magnetic field according to above equation. Each ion sets up a minute electrical current.

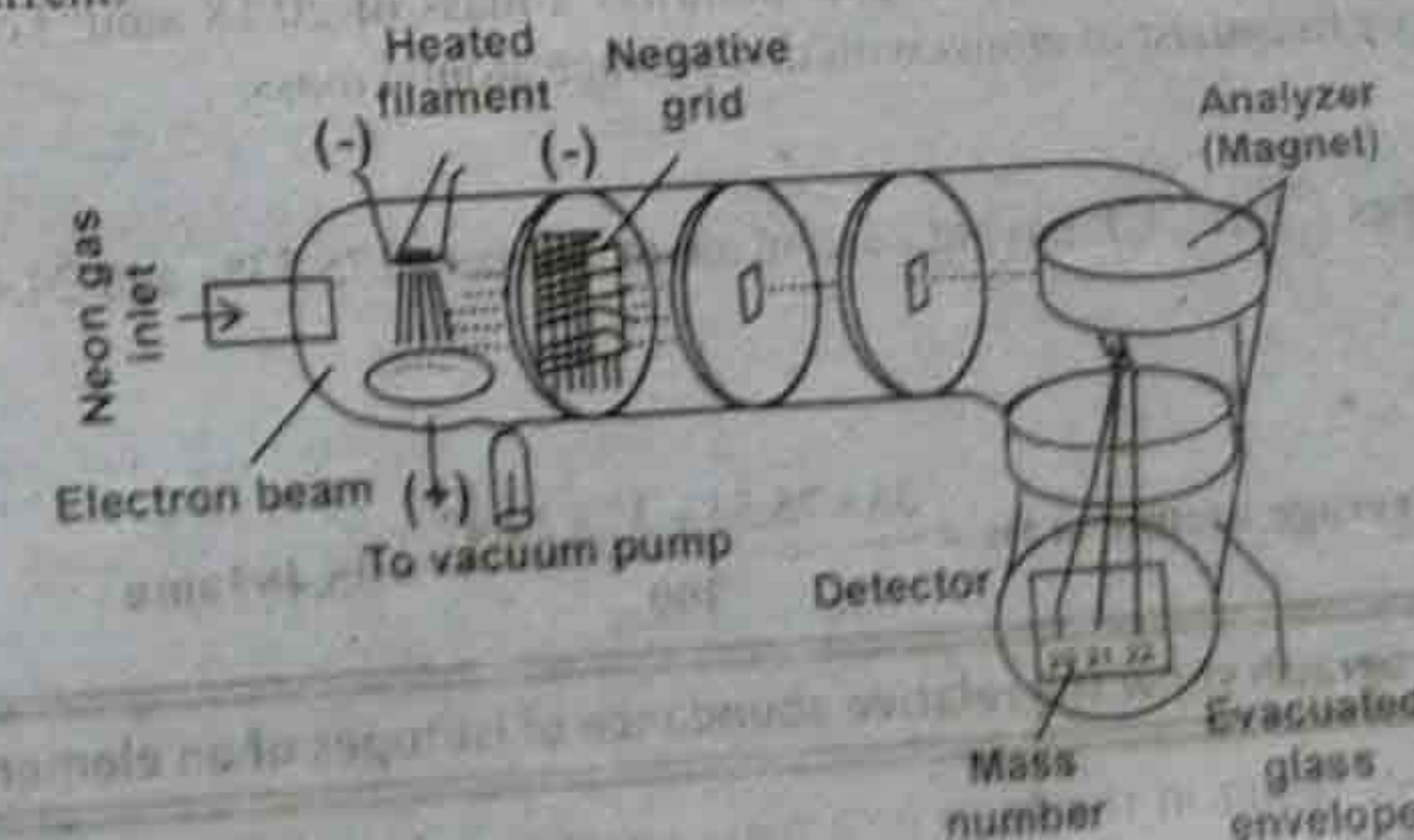


Diagram of a simple mass spectrometer

### The need for a vacuum

The space inside a mass spectrometer is connected to a vacuum pump. The ions under study must be able to move freely. The mass spectrometer would not work properly if the ions collided with the molecules ( $\text{O}_2$  &  $\text{N}_2$ ) present in the atmosphere.

Q. How the ions in the mass spectrometer are accelerated and deflected?

If an element exists as diatomic molecules i.e.  $\text{Cl}_2$ ,  $\text{O}_2$ ,  $\text{H}_2$  etc. then the spectrum will contain peaks both for the separate atoms and for the molecules. e.g. the mass spectrum of chlorine will have peaks for  $\text{Cl}^+$ , with  $\frac{m}{e}$  values of 35 and 37 and peaks for  $\text{Cl}_2^+$  with  $\frac{m}{e}$  values of 70, 72 and 74.



**(vi) Electrometer (ion collector)**

Electrometer develops the electrical current. The strength of the current thus measured gives the relative abundance of ions of a definite  $m/e$  value.

**(vii) Comparison with Carbon-12**

Similarly, the ions of other isotopes having different masses are made to fall on the collector and the current strength is measured. The current strength in each case gives the relative abundance of each of the isotopes. The same experiment is performed with  $C-12$  isotope and the current strength is compared. This comparison allows us to measure the exact mass number of the isotope.

**Modern Spectrograph**

In modern spectrographs, each ion strikes a detector, the ionic current is amplified and is fed to the recorder. The recorder makes a graph showing the relative abundance of isotopes plotted against the mass number.

**Separation of isotopes**

Since isotopes of an element have same chemical properties, so they cannot be separated by chemical methods. Following physical methods are used for their separation

1. Gaseous diffusion
2. Thermal diffusion
3. Distillation
4. Ultracentrifuge
5. Electromagnetic separation
6. Laser separation

**Q.5 (a)** How do you deduce the fractional atomic masses of elements from the relative isotopic abundance? Give two examples in support of your answer.

**Average Atomic Masses**

"The mass of an element which is obtained from isotopic mass and relative abundance of its isotopes is called average atomic mass".

**Example # 1**

A sample of neon is found to consist of  $^{20}_{10}\text{Ne}$ ,  $^{21}_{10}\text{Ne}$ ,  $^{22}_{10}\text{Ne}$  in the percentages of 90.92%, 0.26% and 8.82% respectively. Calculate the fractional atomic mass of neon.

**Solution**

The overall atomic mass of neon is the average of the determined atomic masses of individual isotopes. Hence,

$$\text{Average atomic mass} = \frac{20 \times 90.92 + 21 \times 0.26 + 22 \times 8.82}{100} = 20.18 \text{ amu}$$

Hence, the average atomic mass of neon is 20.18 amu

It is important to realize that no individual neon atom has a mass of 20.18 amu. For most laboratory purposes, however, we consider the sample to consist of atoms with this average atomic mass.

**Example # 2**

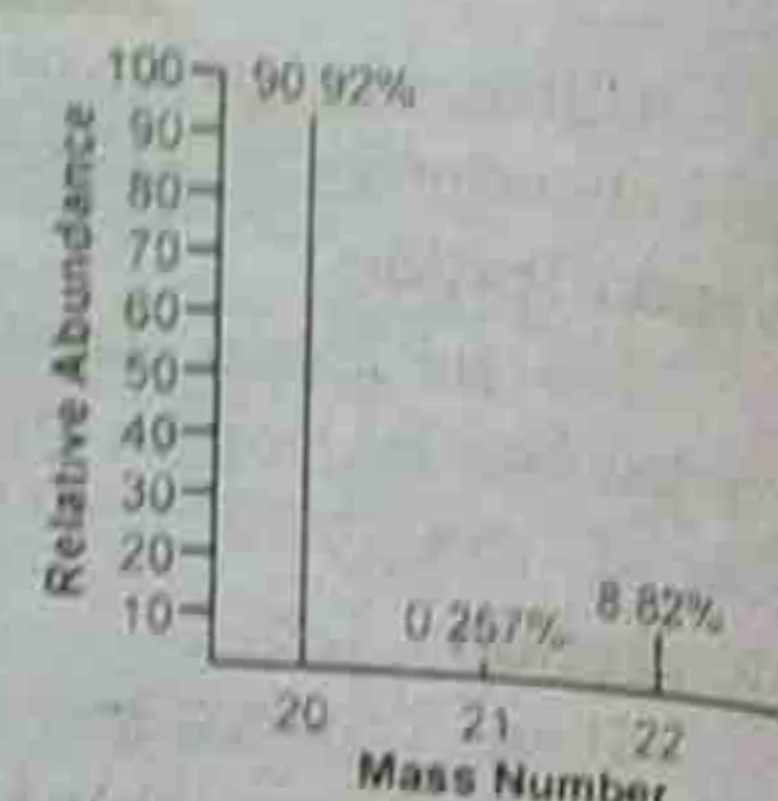
Chlorine has two isotopes  $^{35}_{17}\text{Cl}$ ,  $^{37}_{17}\text{Cl}$  having percentage existence as 75.53% and 24.47% respectively. Calculate average atomic mass of Cl.

**Solution**

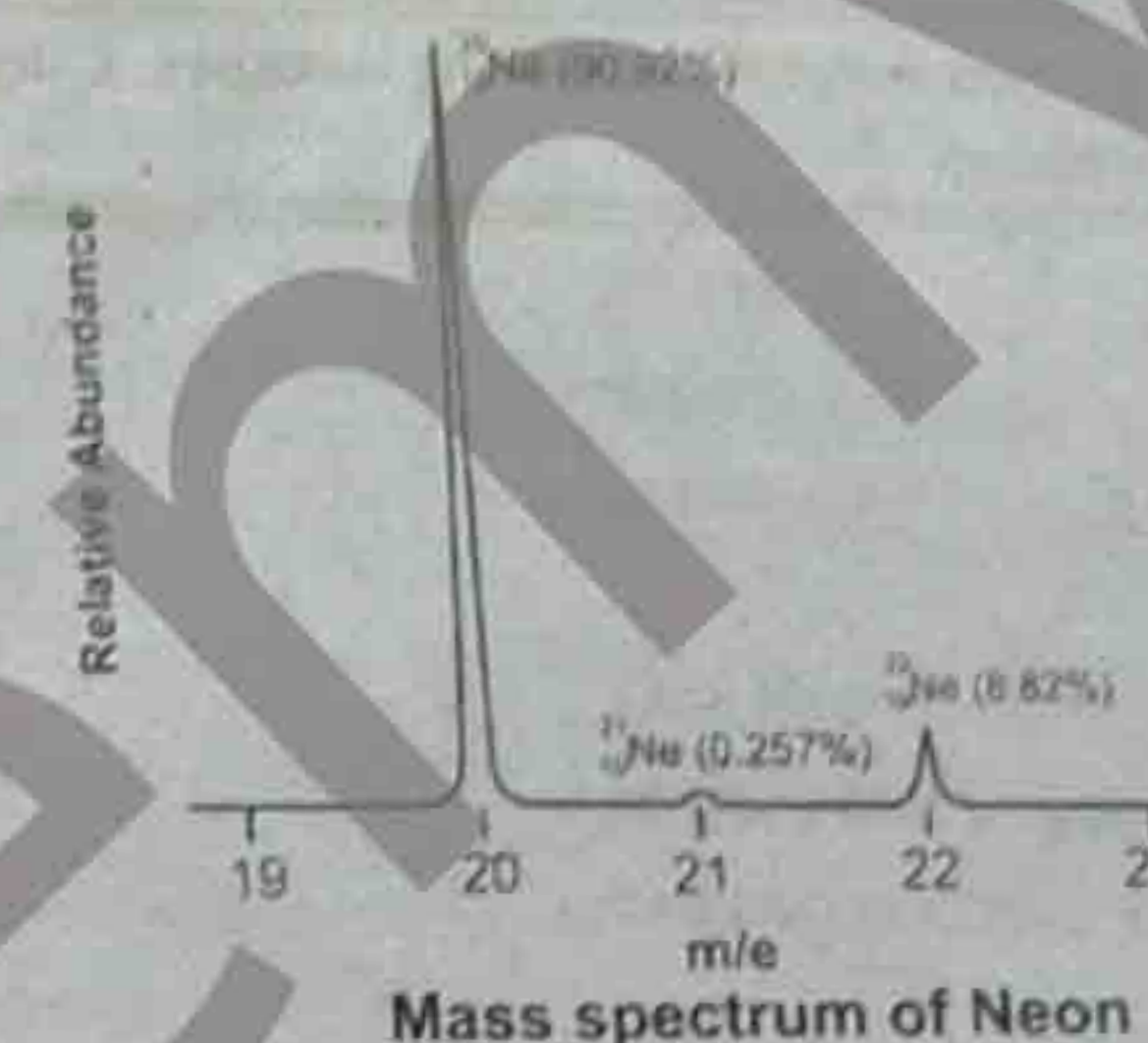
$$\text{Average atomic mass} = \frac{35 \times 75.53 + 37 \times 24.47}{100} = 35.489 \text{ amu}$$

**Q5b.** How does a mass spectrograph show the relative abundance of isotopes of an element?

**Ans.** A mass spectrograph gives result in the form of a mass spectrum. Actually mass spectrum is a plot of data in such a way that  $(m/e)$  of an isotope is plotted as abscissa (X-axis) and relative number of ions as ordinate (Y-axis).



Computer plotted graph for the isotopes of neon.



Mass spectrum of Neon

- Height of peaks in the graph is proportional to relative natural abundance of isotope of an element.
- Number of Peaks gives number of possible isotopes.

**Q5c.** What is the justification of two strong peaks in the mass spectrum for bromine; while for iodine only one peak at 127 amu is indicated?

**Ans.** In mass spectrum height of peaks indicates relative abundance of isotope of an element where as number of peaks gives us number of possible isotopes of that element.

In case of bromine, we observe two strong peaks of almost equal heights. It shows that Br has two naturally occurring isotopes i.e.,  $^{79}\text{Br}$  and  $^{81}\text{Br}$  with a relative abundance of 50.54% and 49.46% respectively.

On the other hand iodine being mono isotopic  $^{127}\text{I}$  gives only one peak at 127 amu (atomic mass of iodine).

**Q6.** Silver has atomic number 47 and has 16 known isotopes but two occur naturally i.e. Ag-107 and Ag-109. Given the following mass spectrometric data, calculate the average atomic mass of silver.

Isotopes	mass (amu)	Percentage abundance
$^{107}\text{Ag}$	106.90509	51.84
$^{109}\text{Ag}$	108.90476	48.16

**Ans. Given data**

Mass of $^{107}\text{Ag}$	= 106.90509 amu
Mass of $^{109}\text{Ag}$	= 108.90476 amu
Percentage of $^{107}\text{Ag}$	= 51.84%
Percentage of $^{109}\text{Ag}$	= 48.16%

**Required**

Average atomic mass of Ag = ?

**Solution**

$$\begin{aligned} \text{Average atomic mass of Ag} &= \frac{(\text{Isotopic mass of } ^{107}\text{Ag} \times \% \text{ of } ^{107}\text{Ag}) + (\text{Isotopic mass of } ^{109}\text{Ag} \times \% \text{ of } ^{109}\text{Ag})}{100} \\ &= \frac{(106.90509 \times 51.84) + (108.90476 \times 48.16)}{100} \\ &= \frac{5541.96 + 5244.85}{100} \\ &= \frac{10786.81}{100} \\ &= 107.87 \text{ amu} \end{aligned}$$

Average atomic mass of Ag



Q7. Boron with atomic number 5 has two naturally occurring isotopes. Calculate the percentage abundance of  $^{10}\text{B}$  and  $^{11}\text{B}$  from the following information

Average Atomic mass of Boron	= 10.81 amu
Isotopic mass of $^{10}\text{B}$	= 10.0129 amu
Isotopic mass of $^{11}\text{B}$	= 11.0093 amu

Ans. Given data

Average Atomic mass of Boron	= 10.81 amu
Isotopic mass of $^{10}\text{B}$	= 10.0129 amu
Isotopic mass of $^{11}\text{B}$	= 11.0093 amu

Required

Percentage abundance of $^{10}\text{B}$	= ?
Percentage abundance of $^{11}\text{B}$	= ?

Solution

Suppose

$$\text{Percentage abundance of } ^{10}\text{B} = x$$

$$\text{Percentage abundance of } ^{11}\text{B} = (100 - x)$$

$$\text{Average Atomic mass of boron} = \frac{(\text{Isotopic mass of } ^{10}\text{B} \times \% \text{ of } ^{10}\text{B}) + (\text{Isotopic mass of } ^{11}\text{B} \times \% \text{ of } ^{11}\text{B})}{100}$$

Putting the values

$$10.81 = \frac{(10.0129 \times x) + [11.0093 \times (100 - x)]}{100}$$

$$10.81 \times 100 = (10.0129x) + (1100.93 - 11.0093x)$$

$$1081 - 1100.93 = 10.0129x - 11.0093x$$

$$-19.93 = -0.9964x$$

$$\frac{-19.93}{-0.9964} = x$$

$$x = 20.002\%$$

$$x = 20.002\%$$

$$\text{Percentage abundance of } ^{10}\text{B} = 20.002\%$$

$$\text{Percentage abundance of } ^{11}\text{B} = 100 - 20.002 = 79.998\%$$

### DETERMINATION OF PERCENTAGE OF ELEMENTS IN A COMPOUND

The percentage of an element in a compound is the number of grams of that element present in 100 grams of the compound, and is calculated as

$$\text{Percentage of an element} = \frac{\text{Mass of element in the compound}}{\text{Formula mass of the compound}} \times 100$$

### Empirical Formula

"A chemical formula which shows the simplest whole number ratio between the atoms of different elements present in a compound is called empirical formula."

Examples

Substance	Empirical formula
Glucose	$\text{CH}_2\text{O}$
Benzene	$\text{CH}$
Water	$\text{H}_2\text{O}$
Sodium Chloride	$\text{NaCl}$

### Steps involved in determination of empirical formula

The following steps are involved in the determination of empirical formula.

1) Determination of the percentage composition

$$\% \text{ of an element} = \frac{\text{mass of the element}}{\text{molar mass of compound}} \times 100$$

2) Finding the number of gram atoms of each element. For this purpose divide the percentage of an element by its atomic mass.

$$\text{Number of gram atoms (moles)} = \frac{\% \text{ of an element}}{\text{atomic mass}}$$

3) Determination of the atomic ratio of each element. To get this, divide the number of moles of each element (gram atoms) by the smallest number of moles.

$$\text{Atomic ratio} = \frac{\text{number of moles}}{\text{smallest number of moles}}$$

4) If the atomic ratio is simple whole number, it gives the empirical formula, otherwise multiply with a suitable digit to get the whole number atomic ratio.

### EMPIRICAL FORMULA FROM COMBUSTION ANALYSIS

Q. Combustion analysis is used to determine?

Combustion:

The process in which an organic compound is burnt in excess of  $\text{O}_2$ .  $\text{CO}_2$  and  $\text{H}_2\text{O}$  are formed with evolution of energy.



$$(\Delta H = -891 \text{ kJ/mol})$$

### Combustion Analysis:

It is an experimental procedure by which amounts of various elements present in the given amount of a compound are determined by burning.

Organic compounds containing only C, H and O are analyzed by combustion analysis.

Sole Product

$\text{CO}_{2(g)}$  and  $\text{H}_2\text{O}_{(g)}$  are sole (main) products during combustion of an organic compound.

Procedure

A weighed sample of the organic compound is placed in the combustion tube. This combustion tube is fitted in a furnace. Oxygen is supplied to burn the organic compound. Hydrogen is converted into  $\text{H}_2\text{O}$  and carbon is converted into  $\text{CO}_2$ .

$\text{H}_2\text{O}$  absorber

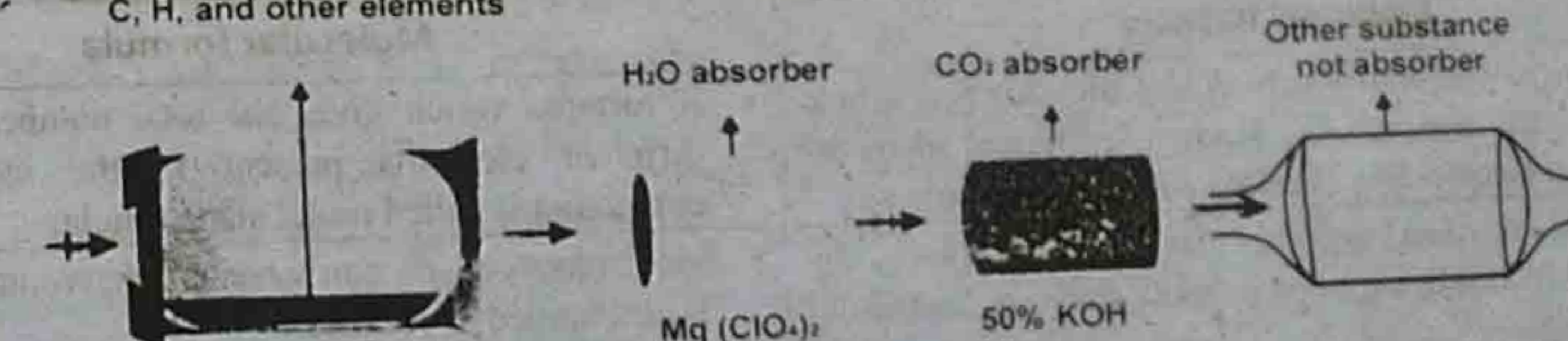
Next to combustion tube there is  $\text{H}_2\text{O}$  absorber chamber. It contains  $\text{Mg}(\text{ClO}_4)_2$  which absorbs  $\text{H}_2\text{O}$  produced during combustion.

$\text{CO}_2$  absorber

After  $\text{H}_2\text{O}$  absorber chamber, next is the  $\text{CO}_2$  absorber chamber. It contains 50%  $\text{KOH}$  which absorbs  $\text{CO}_2$  produced during combustion.

The difference in the amounts or masses of these absorbers gives us the amounts of  $\text{H}_2\text{O}$  and  $\text{CO}_2$  produced.

Sample of compound containing C, H, and other elements



### Combustion analysis

Calculations

$$\% \text{ of Carbon} = \frac{\text{Mass of } \text{CO}_2}{\text{Mass of organic compound}} \times \frac{12.00}{44.00} \times 100$$



$$\% \text{ of Hydrogen} = \frac{\text{Mass of H}_2\text{O}}{\text{Mass of organic compound}} \times \frac{2.016}{18} \times 100$$

The percentage of oxygen is obtained by the method of difference.

$$\% \text{ of Oxygen} = 100 - (\% \text{ of carbon} + \% \text{ of hydrogen})$$

### Molecular Formula

The formula, which shows the exact number of atoms of each element present in one molecule of a compound, is called molecular formula.

#### Examples

Substance	Molecular Formula
Water	H <sub>2</sub> O
Benzene	C <sub>6</sub> H <sub>6</sub>
Glucose	C <sub>6</sub> H <sub>12</sub> O <sub>6</sub>
Sulphuric acid	H <sub>2</sub> SO <sub>4</sub>

Q. Why some compounds have same empirical and molecular formula?

#### Relationship between empirical and molecular formula

- 1) A compound may have same empirical and molecular formula. For example, CH<sub>4</sub>, H<sub>2</sub>O, CO<sub>2</sub>, NH<sub>3</sub>, C<sub>12</sub>H<sub>22</sub>O<sub>11</sub> are the empirical and molecular formulas of the respected compounds. For such compounds the value of "n" is unity (1).
- 2) The molecular formula may be integral multiple of empirical formula. For example, Molecular formula of glucose is C<sub>6</sub>H<sub>12</sub>O<sub>6</sub> which is six times the empirical formula of glucose (CH<sub>2</sub>O). The molecular formula of benzene is C<sub>6</sub>H<sub>6</sub> which is six times the empirical formula of benzene (CH). So molecular formula is related to empirical formula as

$$\text{Molecular formula} = n \times \text{Empirical formula}$$

- 3) A compound may have empirical formula but no molecular formula. For example, sodium chloride has the empirical formula NaCl but no molecular formula. All the ionic compounds have empirical formula but no molecular formula.
- 4) Molecular compounds having same molecular formula are called isomers e.g. ethyl alcohol or dimethyl ether have same molecular formula C<sub>2</sub>H<sub>6</sub>O.

### DIFFERENCE BETWEEN EMPIRICAL AND MOLECULAR FORMULA

Empirical formula	Molecular formula
<ul style="list-style-type: none"> <li>A chemical formula which shows the simplest whole number ratio between the atoms of different elements present in a compound is called empirical formula.</li> <li>Empirical formula can be determined directly by different methods e.g., combustion analysis, elemental analysis etc.</li> <li>Empirical formula is used for all compounds.</li> </ul>	<ul style="list-style-type: none"> <li>A formula which gives the total number of atoms of different elements present in the molecule of a compound is called molecular formula.</li> <li>Molecular formula can never be determined directly. It is determined as follows: Molecular formula = n × (Empirical formula)</li> <li>Molecular formula is used for only molecular compounds.</li> </ul>
<b>Examples</b> Empirical formula of glucose = CH <sub>2</sub> O Empirical formula of water = H <sub>2</sub> O	<b>Examples</b> Molecular formula of glucose = C <sub>6</sub> H <sub>12</sub> O <sub>6</sub> Molecular formula of water = H <sub>2</sub> O

Q8. Define the following terms and give three examples of each

- |                        |                    |                          |
|------------------------|--------------------|--------------------------|
| (i) Gram atom          | (v) Molar volume   | (ii) Gram molecular mass |
| (vi) Avogadro's number | (iii) Gram formula | (vii) Stoichiometry      |
| (iv) Gram ion          |                    |                          |

### CONCEPT OF MOLE

#### Gram Atom

"Atomic mass of an element expressed in grams is called a gram atom. It is also called one gram mole or simply a mole of that element."

#### Formula

$$\text{Number of gram atoms or mole of an element} = \frac{\text{Mass of element in grams}}{\text{Atomic mass of an element}}$$

#### Examples

- |                              |   |          |
|------------------------------|---|----------|
| (i) 1 gram atom of hydrogen  | = | 1.008 g  |
| (ii) 1 gram atom of carbon   | = | 12.000 g |
| (iii) 1 gram atom of uranium | = | 238.0 g  |

#### Gram Molecule

"Molecular mass of a compound expressed in grams is called gram molecule or gram mole or simply the mole of a substance."

#### Formula

$$\text{Number of gram molecules or moles of a molecular substance} = \frac{\text{Mass of the molecular substance in grams}}{\text{Molecular mass of the substance}}$$

#### Examples

- |  |   |         |
|--|---|---------|
| (i) 1 gram molecule of water                           | = | 18.0 g  |
| (ii) 1 gram molecule of H <sub>2</sub> SO <sub>4</sub> | = | 98.0 g  |
| (iii) 1 gram molecule of sucrose                       | = | 342.0 g |

#### Gram Formula

"The formula mass of an ionic compound expressed in grams is called gram formula of the substance. It is also called gram mole or simply mole."

#### Formula

$$\text{Number of gram formulas or moles of ionic substance} = \frac{\text{Mass of the ionic substance in grams}}{\text{Formula mass of the ionic substance}}$$

#### Examples

- |  |   |         |
|--|---|---------|
| (i) 1 gram formula of NaCl                             | = | 58.50 g |
| (ii) 1 gram formula of Na <sub>2</sub> CO <sub>3</sub> | = | 106.0 g |
| (iii) 1 gram formula of AgNO <sub>3</sub>              | = | 170.0 g |

#### Gram Ion

"Ionic mass of an ionic specie expressed in grams is called one gram ion or one mole of ions."

#### Formula

$$\text{Number of gram ions or moles of ionic specie} = \frac{\text{Mass of ionic specie in grams}}{\text{Formula mass of the ionic specie}}$$

#### Particle

A single particle of a substance may refer to an atom, a molecule, an ion, an electron or to any identifiable particle. Chemists refer to a collection of particles as chemical species.



## Examples

- (i) 1 gram formula of  $\text{OH}^-$  = 17 g  
 (ii) 1 gram ion of  $\text{SO}_4^{2-}$  = 96 g  
 (iii) 1 gram ion of  $\text{CO}_3^{2-}$  = 60 g

## Mole

"When the atomic mass of an element, molecular mass of a molecular substance, formula mass of an ionic compound or ionic mass of an ionic species is expressed in grams then it is called **mole**."

or

"The amount of a substance which contains Avogadro's number of particles (atoms, molecules, formula units, or ions) is called **mole**."

- It is denoted by "n".
- It is abbreviated as "mol".
- It is the SI unit of measuring the quantity of substance.

## Examples

Element	1 mole of Na	= 23 g
Molecular compound	1 mole of $\text{H}_2\text{O}$	= 18 g
Ionic compound	1 mole of NaCl	= 58.5 g
Ionic specie	1 mole of $\text{HCO}_3^-$	= 61 g

## Formula

$$\text{Mole} = \frac{\text{Given mass of the substance}}{\text{Atomic mass / Molecular mass / Formula mass}}$$

$$\text{Mole} = \frac{\text{Given mass of the substance}}{\text{Molar mass}}$$

## Avogadro's Number

"The number of particles (atoms, molecules, formula units or ions) which are present in one mole of a substance is called Avogadro's number."

or

"The number of atoms, molecules or ions in one gram atom of an element, one gram molecule of a compound and one gram ion of a substance, respectively is called Avogadro's number."

Symbol	It is represented by " $N_A$ ".
Value	It's value is $6.022 \times 10^{23}$ .
Unit	It's unit is $\text{mol}^{-1}$ .

## Examples

1 mole of Na	= 23 g = $6.022 \times 10^{23}$ atoms
1 mole of $\text{H}_2\text{SO}_4$	= 98 g = $6.022 \times 10^{23}$ molecules
1 mole of $\text{CaCl}_2$	= 111 g = $6.022 \times 10^{23}$ formula units
1 mole of $\text{OH}^-$	= 17 g = $6.022 \times 10^{23}$ ions

## Relationships

The relationships between amounts of substances in term of their masses and number of particles present in them are

$$1) \text{ Number of atoms of an element} = \frac{\text{Mass of the element}}{\text{Atomic mass}} \times N_A$$

The word 'mole' was introduced around 1896 by Wilhelm Ostwald who derived the term from Latin word moles meaning a 'heap' or 'pile'.

$$n = \frac{m}{M}$$

## A mole of substance represents:

- $6.023 \times 10^{23}$  particles
- 22.4 L of gas at STP
- 1 gram atom of an element
- 1 gram molecular mass of a substance
- 1 gram formula mass of an ionic substance

Q. What is the mass of one molecule of water?

- (a)  $\frac{6.0 \times 10^{23}}{18}$  g  
 (b)  $\frac{6.0}{18} \times 10^{-23}$  g  
 (c)  $\frac{18}{6.0 \times 10^{23}}$  g  
 (d)  $\frac{1}{6.0 \times 18 \times 10^{23}}$  g

Q. Which of the following contains the largest number of atoms?

- (a) 8g of methane  
 (b) 15g of hydrogen fluoride  
 (c) 15.6g of benzene  
 (d) 21.8g of bromoethane

- 2) Number of molecules of a compound =  $\frac{\text{Mass of compound}}{\text{Molecular mass}} \times N_A$   
 3) Number of ions of ionic specie =  $\frac{\text{Mass of the ion}}{\text{Ionic mass}} \times N_A$

Q9. (a) 23 g of sodium and 238 g of uranium have equal number of atoms in them.

Ans. Sodium and uranium both are elements. 23g & 238g are the molar masses of sodium & uranium respectively.

$$23 \text{ g of Na} = 1 \text{ mol}$$

$$238 \text{ g of uranium} = 1 \text{ mol}$$

According to definition of Avogadro's number, 1 mol of all the elements have same number of atoms in them i.e.,  $6.02 \times 10^{23}$ . Therefore

$$23 \text{ g of sodium} = 1 \text{ mol} = 6.02 \times 10^{23} \text{ sodium atoms}$$

$$238 \text{ g of uranium} = 1 \text{ mol} = 6.02 \times 10^{23} \text{ uranium atoms}$$

Q9 (b) Mg atom is twice heavier than that of carbon atom.

Ans. We know the molar masses of each element i.e.

$$1 \text{ mol of magnesium} = 24 \text{ g} = 6.02 \times 10^{23} \text{ magnesium atoms}$$

$$1 \text{ mol of carbon} = 12 \text{ g} = 6.02 \times 10^{23} \text{ carbon atoms}$$

By this information, we can calculate the mass of each magnesium and carbon atom by dividing molar masses with Avogadro's number as follows

$$6.02 \times 10^{23} \text{ magnesium atoms has mass} = 24 \text{ g}$$

$$1 \text{ magnesium atom has mass} = \frac{24 \text{ g}}{6.02 \times 10^{23}} = 3.9867 \times 10^{-23} \text{ g}$$

Similarly,

$$6.02 \times 10^{23} \text{ carbon atoms has mass} = 12 \text{ g}$$

$$1 \text{ carbon atom has a mass} = \frac{12}{6.02 \times 10^{23}} = 1.9933 \times 10^{-23} \text{ g}$$

By comparing both masses

Mg	C
$3.9867 \times 10^{-23} \text{ g}$	$1.9933 \times 10^{-23} \text{ g}$
Simplify them	
$\frac{3.9867 \times 10^{-23}}{1.9933 \times 10^{-23}}$	$\frac{1.9933 \times 10^{-23}}{1.9933 \times 10^{-23}}$
2	1

We can easily conclude that magnesium atom is twice heavier in mass than that of a carbon atom.

Q9. (c) 180 g of glucose and 342 g of sucrose have same number of molecules but different number of atoms present in them.

Ans. Molar mass of glucose ( $\text{C}_6\text{H}_{12}\text{O}_6$ ) = 180 g  $\text{mol}^{-1}$   
 Molar mass of sucrose ( $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ ) = 342 g  $\text{mol}^{-1}$

Handwritten notes on the right page:  
 Mg 24g, C 12g  
 1 mol Mg = 24g, 1 mol C = 12g  
 1 mol contains  $6.02 \times 10^{23}$  atoms  
 Therefore, 1 atom of Mg =  $\frac{24}{6.02 \times 10^{23}}$  g  
 1 atom of C =  $\frac{12}{6.02 \times 10^{23}}$  g  
 Hence, Mg atom is twice heavier than C atom.



Both glucose and sucrose are molecular species. Therefore one mole contains same number of molecules i.e., Avogadro's number ( $6.02 \times 10^{23}$ ).

180 g of glucose = 1 mol =  $6.02 \times 10^{23}$  molecules of glucose

342 g of sucrose = 1 mol =  $6.02 \times 10^{23}$  molecules of sucrose

One molecule of glucose has different number of atoms than one molecule of sucrose so one mole of each molecule contains different number of atoms in them.

As 1 molecule of glucose contains = 24 atoms

So 1 mole glucose molecules contain =  $24 \times 6.02 \times 10^{23}$  atoms

Similarly

As 1 molecule of sucrose contains = 45 atoms

So 1 mole sucrose molecules contain =  $45 \times 6.02 \times 10^{23}$  atoms

**Q9. (d)** 4.9 g of  $\text{H}_2\text{SO}_4$  when completely ionized in water, have equal number of positive and negative charges but the number of positively charged ions are twice the number of negatively charged ions.

**Ans. Given data**

Mass of  $\text{H}_2\text{SO}_4$  = 4.9 g

**Required**

Number of  $\text{H}^+$  = ?

Number of  $\text{SO}_4^{2-}$  ions = ?

Total positive charge = ?

Total negative charge = ?

**Solution**

(i) Firstly we calculate number of moles

$$\text{Number of moles of } \text{H}_2\text{SO}_4 = \frac{\text{Mass in grams}}{\text{Molar mass}}$$

$$\text{Molar mass of } \text{H}_2\text{SO}_4 = 98 \text{ g mol}^{-1}$$

$$\text{So number of moles} = \frac{4.9 \text{ g}}{98 \text{ g mol}^{-1}} = 0.05 \text{ moles}$$

(ii) Now calculate molecules of  $\text{H}_2\text{SO}_4$

$$\begin{aligned} \text{Number of molecules of } \text{H}_2\text{SO}_4 &= \text{Number of moles of } \text{H}_2\text{SO}_4 \times N_A \\ &= 0.05 \times 6.02 \times 10^{23} \\ &= 3.01 \times 10^{22} \text{ molecules} \end{aligned}$$

(iii) Ionization of  $\text{H}_2\text{SO}_4$  in water



Now we can calculate

**Total positive ions**

$\text{H}_2\text{SO}_4$	:	$\text{H}^+$
1	:	2
$3.01 \times 10^{22}$	:	$2 \times 3.01 \times 10^{22}$
	:	$6.02 \times 10^{22}$ ions

**Total negative ions**

$\text{H}_2\text{SO}_4$	:	$\text{SO}_4^{2-}$
1	:	1
$3.01 \times 10^{22}$	:	$3.01 \times 10^{22}$ ions

**Comparison of  $\text{H}^+$  and  $\text{SO}_4^{2-}$**

$\text{H}^+$	$\text{SO}_4^{2-}$
$6.02 \times 10^{22}$	$3.01 \times 10^{22}$
$6.02 \times 10^{22}$	$3.01 \times 10^{22}$
$3.01 \times 10^{22}$	$3.01 \times 10^{22}$
2	1

♦ So number of positive ions are twice than that of negative ions

**Total positive charge**

Number of positive ions  $\times$  charge on one positive ion

$$= 6.02 \times 10^{22} \times (+1)$$

$$= +6.02 \times 10^{22}$$

**Total negative charge**

Number of negative ions  $\times$  charge on one negative ion

$$= 3.01 \times 10^{22} \times (-2)$$

$$= -6.02 \times 10^{22}$$

♦ So charges are same.

(e) One mg of  $\text{K}_2\text{CrO}_4$  has thrice the number of ions than the number of formula units when ionized in water.

**Ans. Given data**

Mass of  $\text{K}_2\text{CrO}_4$  = 1 mg =  $10^{-3}$  g

**Required**

Total ions of  $\text{K}_2\text{CrO}_4$  = ?

**Solution**

$$\text{Molar mass of } \text{K}_2\text{CrO}_4 = 194 \text{ g mol}^{-1}$$

$$\text{Number of moles of } \text{K}_2\text{CrO}_4 = \frac{\text{Mass in grams}}{\text{Molar mass}}$$

$$= \frac{10^{-3} \text{ g}}{194 \text{ g mol}^{-1}} = 5.15 \times 10^{-6} \text{ moles}$$

$$\text{Total formula units of } \text{K}_2\text{CrO}_4 = \text{Number of moles of } \text{K}_2\text{CrO}_4 \times N_A$$

$$= 5.15 \times 10^{-6} \times 6.02 \times 10^{23}$$

$$= 3.10 \times 10^{18} \text{ formula units}$$

In water,  $\text{K}_2\text{CrO}_4$  ionizes as



1 formula unit of  $\text{K}_2\text{CrO}_4$  produces ions = 3

$$\begin{aligned} 3.10 \times 10^{18} \text{ formula units of } \text{K}_2\text{CrO}_4 \text{ produce ions} &= 3 \times 3.10 \times 10^{18} \\ &= 9.30 \times 10^{18} \text{ ions} \end{aligned}$$

So one mg of  $\text{K}_2\text{CrO}_4$  produces thrice the number of ions than the number of formula units.

**Molar Mass**

"The mass of one mole of a substance is called molar mass."

**Unit** Its unit is g/mole.



## Examples

Molar mass of Na	= 23 g mol <sup>-1</sup>
Molar mass of (C <sub>12</sub> H <sub>22</sub> O <sub>11</sub> )	= 342 g mol <sup>-1</sup>
Molar mass of CaCl <sub>2</sub>	= 111 g mol <sup>-1</sup>
Molar mass of OH <sup>-</sup>	= 17 g mol <sup>-1</sup>

## Molar Volume

"The volume occupied by one mole of an ideal gas at standard temperature and pressure (STP) i.e. 0°C and 1 atm is called molar volume."

**Symbol** Its symbol is V<sub>m</sub>

**Value** Its value is 22.414 dm<sup>3</sup> (0.022414 m<sup>3</sup> or 22414 cm<sup>3</sup>)

## Examples

- (i) 2g of H<sub>2</sub> = 1 mole of H<sub>2</sub> = 22.414 dm<sup>3</sup>  
 (ii) 16g of CH<sub>4</sub> = 1 mole of CH<sub>4</sub> = 22.414 dm<sup>3</sup>

From above examples it is clear that 1 mole of different gases have same number of molecules and same volume at STP but different masses.

Q9. (f) 2 g of H<sub>2</sub>, 16 g of CH<sub>4</sub> and 44 g of CO<sub>2</sub> occupy separately the volumes of 22.414 dm<sup>3</sup>, although the sizes and masses of three gases are very different from each other.

Ans. 2g of H<sub>2</sub> = 1 mol H<sub>2</sub> = 6.02 × 10<sup>23</sup> molecule = 22.414 dm<sup>3</sup>  
 16g of CH<sub>4</sub> = 1 mol CH<sub>4</sub> = 6.02 × 10<sup>23</sup> molecules = 22.414 dm<sup>3</sup>  
 44g of CO<sub>2</sub> = 1 mol CO<sub>2</sub> = 6.02 × 10<sup>23</sup> molecules = 22.414 dm<sup>3</sup>

According to Avogadro's law, equal moles of gases at same temperature and pressure (STP) occupy same volume. One mole of any gas can occupy 22.414 dm<sup>3</sup> at STP. So 2g of H<sub>2</sub>, 16g of CH<sub>4</sub> and 44g of CO<sub>2</sub> occupy 22.414 dm<sup>3</sup>.

Volume occupied by a gas does not depend on size and mass of gas molecule. It only depends on number of molecules. Reason is that at STP, distance between gas molecules is 300 times greater than their own diameters.

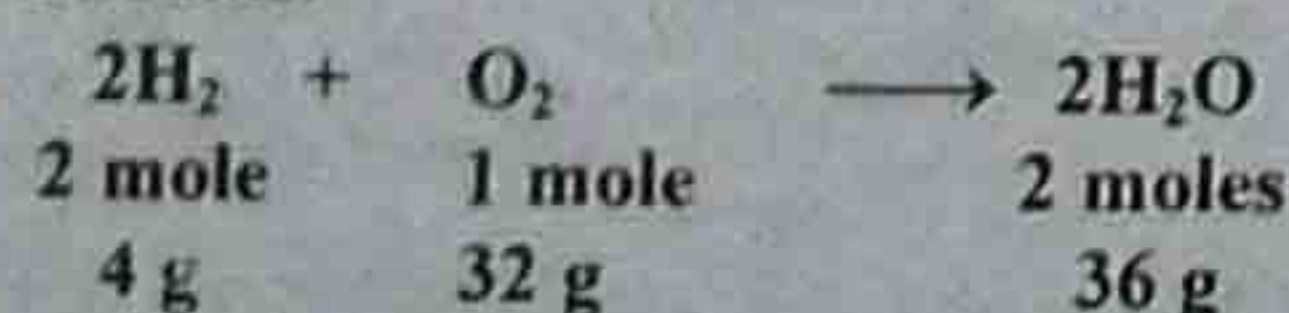
Q23. (a) What is stoichiometry? Give its assumptions? Mention two important laws, which help to perform the stoichiometric calculations?

## Stoichiometry

"The branch of chemistry which deals with the study of quantitative relationship between reactants and products in a balanced chemical equation is called stoichiometry."

## Stoichiometric Amounts

"The amounts of the reactants or the products as given by the balanced chemical equation are called stoichiometric amounts."



## Assumptions

When stoichiometric calculations are performed, we have to assume the following conditions

- All the reactants are completely converted into products.
- No side reaction occurs.

During stoichiometric calculations, law of conservation of mass and law of definite proportions are obeyed.

## Relationships

## (i) Mass-mass relationship

If we are given mass of one substance, we can calculate the mass of the other substance involved in the chemical reaction.

## (ii) Mass-mole relationship

If we are given mass of one substance, we can calculate the moles of the other substance and vice versa.

## (iii) Mass-volume relationship

If we are given mass of one substance, we can calculate the volume of the other substance and vice versa. Similarly, mole-mole calculations can also be performed.

## Law of Conservation of mass:

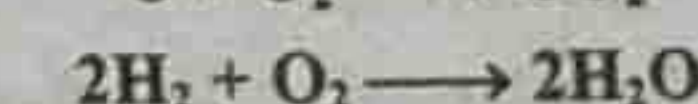
- In all physical and chemical changes, the total mass of the reactants is equal to that of the products.
- It was studied by Antoine Lavoisier (1774).

## The Law of Definite proportions or Constant composition:

- The law was given by J.L. Proust and deals with the composition of elements present in a compound.
- The law states that: The same compound always contains the same elements combined together in the same fixed proportion by weight.
- The composition of the compound is always the same irrespective of the method by which it is produced.

## Chemical Equation

"A statement that describes a chemical reaction in terms of symbols and chemical formulas is called a chemical equation."



## Balancing a chemical equation

There are three methods to balance a chemical equation

- Hit and trial method
- Redox method
- Ion electron method

## Limitations of a chemical equation

The demerits of a chemical equation are as follows

- A chemical equation does not tell the rate of the reaction.
- It does not tell the conditions, necessary for the reaction.
- It does not give colour, odour or state of the reactants and the products.
- We can even write the chemical equation of the reaction that does not occur.

## Limiting Reactant

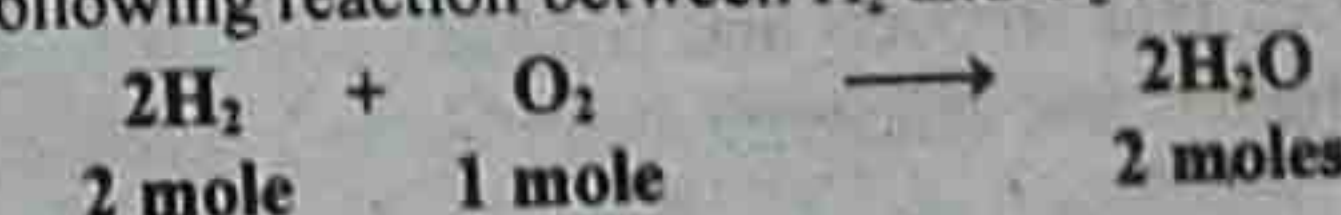
"A reactant which is used or consumed earlier due to its lesser amount and controls the amount of product formed in a chemical reaction is called limiting reactant."

A limiting reactant is that which

- controls the amount of product formed
- is taken in lesser amount
- is consumed earlier
- produces least amount of product

## Examples

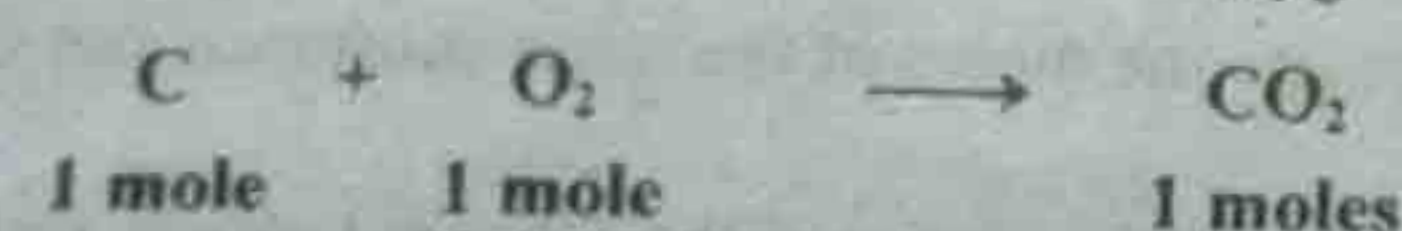
(1) Consider the following reaction between H<sub>2</sub> and O<sub>2</sub> to form water



Suppose, we allow 2 moles of hydrogen (4g) to react with 2 moles of oxygen (64g). According to the above equation 2 moles of hydrogen (4g) will react with only one mole of oxygen (32g) to produce two moles of water (36g). Hence, one mole of oxygen (32g) will be left un-reacted because the whole of the given amount of H<sub>2</sub> has been consumed. Therefore, no more reaction will take place. In this case, H<sub>2</sub> is the limiting reactant because it is consumed first during the chemical reaction and controls the chemical reaction between hydrogen and oxygen.

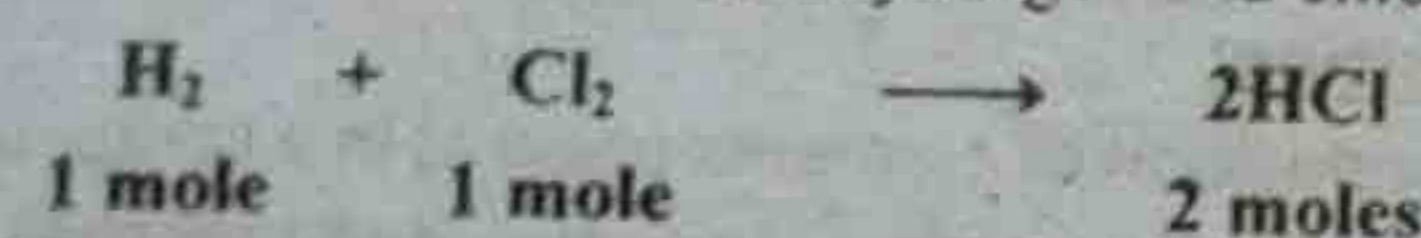


(2) Consider the following reaction between carbon and oxygen to produce carbon dioxide.



Suppose, we allow 2 moles of carbon to react with one mole of oxygen. According to the above equation one mole of oxygen will react with only one mole of carbon to form one mole of carbon dioxide. Hence, one mole of carbon will be left un-reacted because the whole of the amount of oxygen has been consumed. Therefore, no more reaction will take place. In this case oxygen will be the limiting reactant because it is consumed first during the chemical reaction and controls the chemical reaction between carbon and oxygen.

(3) Consider the following reaction between hydrogen and chlorine to form hydrochloric acid.



Suppose, we allow one mole of hydrogen to react with 2 moles of chlorine. According to the above equation one mole of hydrogen will react with only one mole of chlorine to form two moles of hydrochloric acid. Hence, one mole of chlorine will be left un-reacted because the whole of the given amount of hydrogen has been consumed. Therefore, no more reaction will take place. In this case hydrogen will be the limiting reactant because it is consumed first during the chemical reaction and it is controlling the chemical reaction.

#### Identification of Limiting Reactant

The following three steps should be followed to find out the limiting reactant

- Calculate the number of moles from the given amount of reactant.
- Find out the number of moles of product with the help of a balanced chemical equation.
- Identify the reactant which produces the least amount of product as limiting reactant.

Q. Write down steps involve in the determination of limiting reactant.

Q23b. What is a limiting reactant? How does it control the quantity of the product formed? Explain with three examples.

Ans. Limiting reactant

"A limiting reactant is that reactant which controls the amount of product as it consumes earlier due to its smaller amount."

#### Examples

- If 2 moles  $\text{H}_2$  and 2 moles  $\text{O}_2$  are allowed to react then only 2 moles  $\text{H}_2\text{O}$  is produced.  
 $2\text{H}_2 + \text{O}_2 \longrightarrow 2\text{H}_2\text{O}$   
 Here  $\text{H}_2$  consumes completely and one mole  $\text{O}_2$  left behind unreacted so  $\text{H}_2$  is limiting reactant and limits the amount of  $\text{H}_2\text{O}$  upto 2 moles.
- During burning  $\text{O}_2$  is in excess and a combustible material (coal, paper, candle etc.) is in smaller quantity. Combustible material is fully consumed and controls the amount of product so coal, candle or paper is limiting reactant.
- If we have 20 slices and 9 kababs, we can only make 9 sandwiches as kababs are limiting reactant.

Q. Concept of limiting reactant is not applicable on reversible reactions. Justify

Ans: "A reactant which consumes earlier due to its smaller amount and produces least amount of product is called limiting reactant."

During a reversible reaction, reactants are converted into products and products convert back into reactants. So reactants are not completely consumed. As a result a limiting reactant cannot be identified in a reversible reaction.

#### Yield

"The amount of products obtained in a chemical reaction is called yield."

#### Types

Yield is of three types

#### (i) Theoretical yield

"The amount of the products calculated from a balanced chemical equation is called theoretical yield."

- It is also known as calculated yield or expected yield.
- It is the maximum amount of the product which can be produced by a given amount of the reactant according to a balanced chemical equation.
- Theoretical yield of a reaction is always greater than the actual yield of the same reaction.

#### (ii) Actual yield

"The amount of the products obtained with a given amount of the reactant in an actual experiment is called actual yield."

- It is also known as experimental yield.
- The actual yield of a chemical reaction is always lesser than the theoretical yield.

#### (iii) Percentage yield

"It is equal to the ratio of the actual yield to the theoretical yield multiplied by 100."

Formula  $\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$

The efficiency of a chemical reaction is determined with the help of percentage yield.

#### Differentiate between Theoretical and Actual Yield

Theoretical Yield	Actual Yield
1) The amount of product calculated from a balanced chemical equation is called <b>theoretical yield</b> .	1) The amount of product obtained while performing a chemical reaction is called <b>actual yield</b> .
2) It is also called expected yield or calculated yield.	2) It is also called experimental yield.
3) Theoretical yield is always greater than actual yield.	3) Actual yield is always lesser than theoretical yield.
4) No need to perform experiment. Just to calculate from balanced chemical equation.	4) In order to get actual yield experiment has to be performed.

Q. Actual yield is always lesser than theoretical yield. Why?

Ans.

It is due to :

- In experience workers.
- Side reaction may take place.
- Reaction may be reversible.
- Mechanical loss i.e. loss of product during filtration, crystallisation or distillation etc.

Q 24. (a) How do we calculate the percentage yield of a chemical reaction.

Ans. We can calculate the percentage yield of a chemical reaction with the help of actual yield and theoretical yield of the reaction. The efficiency of a reaction is also expressed in the form of percentage yield.

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

Greater the percentage yield, more amount of reactants will convert into product and high will be the efficiency of reaction.

#### KEY POINTS

- Atoms are the building blocks of matter. Atoms can combine to form molecules. Covalent compounds mostly exist in the form of molecules. Atoms and molecules can either gain or lose electrons, forming charged particles called ions. Metals tend to lose electrons, becoming positively charged ions. Non-metals tend to gain electrons forming negatively charged ions. When X-rays or  $\alpha$ -particles are passed through molecules in gaseous state, they are converted into molecular ions.
- The atomic mass of an element is determined with reference to the mass of carbon as a standard element and is expressed in amu. The fractional atomic masses can be calculated from the relative abundance of isotopes. The separation and identification of isotopes can be carried out by mass spectrograph.
- The composition of a substance is given by its chemical formula. A molecular substance can be represented by empirical or molecular formulas. The empirical and molecular formulas are related through a simple integer.
- Combustion analysis is one of the techniques to determine the empirical formula and then the molecular formula of a substance by knowing its molar mass.



- 5) A mole of any substance is the Avogadro's number of atoms or molecules or formula units of that substance.
- 6) The study of quantitative relationship between the reactants and the products in a balanced equation is known as stoichiometry. The mole concept can be used to calculate the relative quantities of reactants and products in a balanced chemical equation.
- 7) The concept of molar volume of gases helps to relate solids and liquids with gases in a quantitative manner.
- 8) A limiting reactant is completely consumed in a reaction and controls the quantity of products formed.
- 9) The theoretical yield of a reaction is the quantity of the products calculated with the help of a balanced chemical equation. The actual yield of a reaction is always less than the theoretical yield. The efficiency of a chemical reaction can be checked by calculating its percentage yield.

### SOLVED OBJECTIVE EXERCISE

Q1. Select the most suitable answer from the given ones in each questions

- (i) Isotopes differ in
  - (a) properties depend upon mass
  - (b) arrangement of electron in orbitals
  - (c) chemical properties
  - (d) the extent to which they may be affected in electromagnetic field
- (ii) Which of the following statement is not true
  - (a) isotopes with even atomic masses are comparatively abundant
  - (b) isotopes with odd atomic masses are comparatively less abundant
  - (c) isotopes with even atomic masses and even atomic numbers are comparatively abundant
  - (d) isotopes with even atomic masses and odd atomic numbers are comparatively abundant
- (iii) Many elements have fractional atomic masses. This is because
  - (a) the mass of atom is itself fractional
  - (b) atomic masses are average masses of isobars
  - (c) atomic masses are average masses of isotopes
  - (d) atomic masses are average masses of isotopes proportional to their relative abundance
- (iv) The mass of one mole of electron is
  - (a) 1.008 mg
  - (b) 0.55 mg
  - (c) 0.184 mg
  - (d) 1.637 mg
- (v) 27 g of Al will react completely with how much mass of  $O_2$  to produce  $Al_2O_3$ 
  - (a) 8 g of oxygen
  - (b) 16 g of oxygen
  - (c) 32 g of oxygen
  - (d) 24 g of oxygen
- (vi) The number of moles of  $CO_2$  which contain 8.0 g of oxygen
  - (a) 0.25
  - (b) 0.50
  - (c) 1.0
  - (d) 1.5
- (vii) The largest number of molecules are present in
  - (a) 3.6 g of  $H_2O$
  - (b) 4.8 g of  $C_2H_5OH$
  - (c) 2.8 g of  $CO$
  - (d) 5.4 g of  $N_2O_5$
- (viii) One mole of  $SO_2$  contains
  - (a)  $6.02 \times 10^{23}$  of oxygen atoms
  - (b)  $18.1 \times 10^{23}$  molecules of  $SO_2$
  - (c)  $6.02 \times 10^{23}$  of sulphur atoms
  - (d) 4 gram atoms of  $SO_2$
- (ix) The volume occupied by 1.4 g of  $N_2$  at STP is
  - (a) 2.24  $dm^3$
  - (b) 22.4  $dm^3$

- (c) 1.12  $dm^3$
- (d) 112  $cm^3$
- (x) A limiting reactant is the one which
  - (a) is taken in lesser quantity in grams as compared to the other reactants
  - (b) is taken in lesser quantity in volume as compared to the other reactants
  - (c) gives the maximum amount of the product which is required
  - (d) gives the minimum amount of product under consideration

### Solved Exercise MCQ's

Q. No	Answer	Reason
(i)	(a) properties depend upon mass	Isotopes are sister atoms of same element which differ by their atomic masses due to different number of neutrons. They show different deflection in magnetic field due to their different m/e values.
(ii)	(c) isotopes with even atomic masses and even atomic numbers are comparatively abundant	The elements of even atomic number usually have larger number of isotopes. The isotopes whose mass numbers are multiple of four are particularly abundant. For example $^{16}O$ , $^{24}Mg$ , $^{28}Si$ , $^{40}Ca$ and $^{56}Fe$ . These isotopes exist abundantly and form about 50% of the earth crust. Out of 280 isotopes that occur in nature, 154 isotopes have even atomic number and even mass number.
(iii)	(d) atomic masses are average masses of isotopes proportional to their relative abundance	The atomic mass of an element is calculated from i) Isotopic masses ii) Relative abundance
(iv)	(b) 0.55 mg	1 mole of electron = $6.022 \times 10^{23}$ Mass of one electron = $9.1095 \times 10^{-31} kg$ Mass of 1 mole of electron = $6.022 \times 10^{23} \times 9.1095 \times 10^{-31}$ $= 54.85 \times 10^{-8} kg$ $= 54.85 \times 10^{-8} \times 10^6 kg = 54.85 \times 10^{-2} mg$ $= 0.55 mg$
(v)	(d) 24 g of oxygen	$4Al + 3O_2 \rightarrow 2Al_2O_3$ Mol of Al = $\frac{27}{27} = 1 mol$ Al : $O_2$ 4 ml : 3 mol 1 mol : $\frac{3}{4} = 0.75 mol$ Mass of $O_2 = 0.75 \times 32 = 24g$ of oxygen
(vi)	(a) 0.25	C : $O_2$ $\downarrow \quad \downarrow$ 12g : 32g = 44g = 1mol $\downarrow$ 16g = 22g = 0.50mol $\downarrow$ 8g = 11g = 0.25mol
(vii)	(a) 3.6 g of $H_2O$	Greater the number of moles of a substance greater will be its number of



		molecules. (a) $\frac{3.6}{18} = 0.2 \text{ mol}$ (b) $\frac{4.8}{48} = 0.1 \text{ mol}$ (c) $\frac{2.8}{28} = 0.1 \text{ mol}$ (d) $\frac{5.4}{108} = 0.05 \text{ mol}$
(viii)	(c) $6.02 \times 10^{23}$ of sulphur atoms	1 mole of $\text{SO}_3 = 6.022 \times 10^{23}$ $\text{SO}_3$ molecules 1 molecule of $\text{SO}_3$ contains one sulphur atom. So $6.022 \times 10^{23}$ $\text{SO}_3$ molecules contain $6.022 \times 10^{23}$ S-atoms
(ix)	(c) $1.12 \text{ dm}^3$	Molecular mass of $\text{N}_2 = 28 \text{ g mol}^{-1}$ 28g of $\text{N}_2$ at STP occupies volume = $22.414 \text{ dm}^3$ 14g of $\text{N}_2$ at STP occupies volume = $11.207 \text{ dm}^3$ 1.4g of $\text{N}_2$ at STP occupies volume = $1.12 \text{ dm}^3$
(x)	(d) gives the minimum amount of product under consideration	A reactant which is used or consumed earlier due to its lesser amount and controls the amount of product formed in a chemical reaction is called limiting reactant.

## Q2. Fill in the blanks

- The unit of relative atomic mass is \_\_\_\_\_.
- The exact masses of isotopes can be determined by \_\_\_\_\_ spectrograph.
- The phenomenon of isotopy was first discovered by \_\_\_\_\_.
- The empirical formula can be determined by combustion analysis for those compounds which have \_\_\_\_\_ and \_\_\_\_\_ in them.
- A limiting reactant is that which controls the quantities of \_\_\_\_\_.
- 1 mole of glucose has \_\_\_\_\_ atoms of carbon, \_\_\_\_\_ atoms of oxygen and \_\_\_\_\_ atoms of hydrogen.
- 4 g of  $\text{CH}_4$  at  $0^\circ\text{C}$  and 1 atm pressure has \_\_\_\_\_ molecules of  $\text{CH}_4$ .
- Stoichiometric calculations can be performed only when \_\_\_\_\_ is obeyed.

## ANSWERS

(i) atomic mass unit (a.m.u)	(ii) Aston's mass
(iii) Soddy	(iv) Carbon, hydrogen
(v) product	(vi) $3.612 \times 10^{24}$ , $3.612 \times 10^{24}$ , $7.224 \times 10^{24}$
(vii) $1.505 \times 10^{23}$	(viii) law of conservation of mass and law of definite proportion

## Q3. Indicate 'True' or 'False'

- Ne has three isotopes and fourth one with atomic mass of 20.18 a.m.u.
- Empirical formula gives us the information about the total number of atoms present in the molecule.
- During combustion analysis  $\text{Mg}(\text{ClO}_4)_2$  is employed to absorb water vapours.
- Molecular formula is the integral multiple of empirical formula and the integral multiple can never be unity.
- The number of atoms in 1.79 g of gold and 0.023 g of sodium are equal.
- The number of electrons in the molecules of CO and  $\text{N}_2$  are 14 each, so 1 mg of each gas will have same number of electrons.
- Avogadro's hypothesis is applicable to all types of gases i.e., ideal and non-ideal.
- Actual yield of a chemical reaction may be greater than the theoretical yield.

## ANSWERS

(i) False	(ii) False	(iii) True	(iv) False
(v) False	(vi) True	(vii) False	(viii) False

## SHORT ANSWERS TO EXERCISE

Q24. (b) What are the factors which are mostly responsible for the low yield of the products in chemical reactions? Or Why actual yield is always less than theoretical yield.

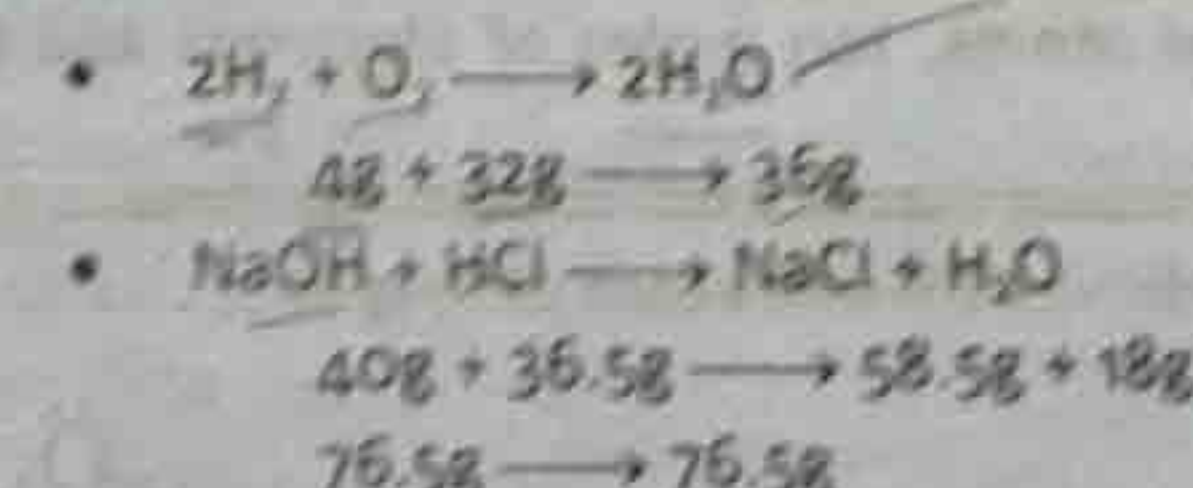
- Ans. Actual yield of a chemical reaction is usually lesser than theoretical yield due to many reasons. They are:
- Practically inexperienced worker has many shortcomings and cannot get expected yield.
  - Mechanical loss during experimentation e.g., filtration, separation by distillation or by a separating funnel, washing, drying and crystallization if not properly carried out, decrease the yield.
  - Some of the reactants might take part in a competing side reaction.
  - The reaction might be reversible.
  - The reactants might be impure.
  - Sometimes the reaction conditions are not suitably maintained.

Q25. Explain the following with reasons.

- (i) Law of conservation of mass has to be obeyed during stoichiometric calculations.

Ans. According to law of conservation of mass, "Mass can neither be created nor destroyed during a chemical reaction but it changes its form from one form to another."  
In stoichiometric calculations, we use balanced chemical equation where total mass of reactants is equal to the total mass of products.

For example



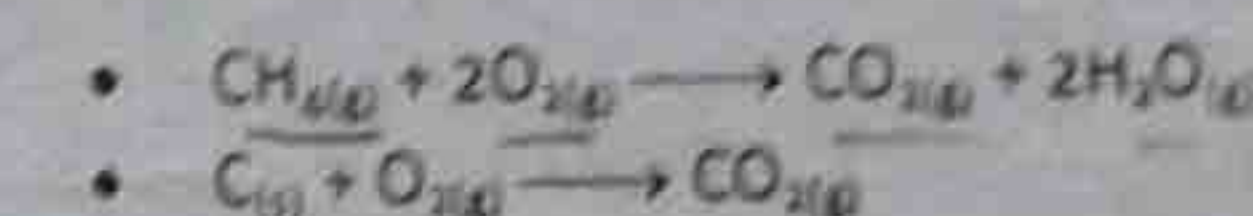
So while doing stoichiometric calculations law of conservation of mass has to be obeyed.

- (ii) Many chemical reactions taking place in our surrounding involve the limiting reactants.

Ans. "The reactant which is consumed earlier and controls the amount of product in a chemical reaction is called a limiting reactant."

Mostly combustion reactions taking place in our surroundings involve the limiting reactant.

For example



In above examples  $\text{O}_2$  gas is present in excess while  $\text{CH}_4$  and C are the reactants which consume earlier so they are limiting reactants.

- (iii) No individual neon atom in the sample of the element has a mass of 20.18 amu.

Ans. 20.18 amu is the average atomic mass of neon which is obtained by using different isotopic masses and relative abundances of Neon as follows.  
Neon has three isotopes  $^{20}\text{Ne}$ ,  $^{21}\text{Ne}$  and  $^{22}\text{Ne}$  with a relative abundance of 90.92%, 0.26% and 8.82% respectively.

$$\text{Average atomic mass of neon} = \frac{\text{Sum of product of isotopic masses and relative abundances}}{100}$$



$$= \frac{(90.92 \times 20) + (0.26 \times 21) + (8.82 \times 22)}{100}$$

$$= 20.18 \text{ amu}$$

So no individual Ne atom has a mass of 20.18 amu.

- (iv) One mole of  $\text{H}_2\text{SO}_4$  should completely react with two moles of  $\text{NaOH}$ . How does Avogadro's number help to explain it?

Ans. "A process in which one mole of water is produced by its ionic components i.e.,  $\text{H}^+$  and  $\text{OH}^-$  coming from acid and base respectively is called neutralization."



$\text{H}_2\text{SO}_4$  being a dibasic acid can produce 2 mole  $\text{H}^+$  ions on ionization in water.



On the other hand  $\text{NaOH}$  being monoacid produces only one mole  $\text{OH}^-$  in aqueous solution.



To neutralize 2 moles  $\text{H}^+$  ions of  $\text{H}_2\text{SO}_4$ , we need 2 moles of  $\text{OH}^-$  so



$$2 \text{ moles of } \text{H}^+ = 2 \times 6.02 \times 10^{23} \text{ ions} = 12.04 \times 10^{23} \text{ ions}$$

$$2 \text{ moles of } \text{OH}^- = 2 \times 6.02 \times 10^{23} \text{ ions} = 12.04 \times 10^{23} \text{ ions}$$

so

$12.04 \times 10^{23} \text{ H}^+ \text{ ions} + 12.04 \times 10^{23} \text{ OH}^- \text{ ions}$  are used to produce  $12.04 \times 10^{23} \text{ H}_2\text{O}$  molecules.

Hence one mole of  $\text{H}_2\text{SO}_4$  is neutralized by 2 moles of  $\text{NaOH}$ .



- (v) One mole of  $\text{H}_2\text{O}$  has two moles of bonds, three moles of atoms, ten moles of electrons and twenty eight moles of total fundamental particles in it.

Ans. Water molecule is formed when two atoms of hydrogen combine with one atom of oxygen.

$$^8_8\text{O} = 8e + 8p + 8n$$

$$^1_1\text{H} = 1e + 1p + 0n$$

$$^1_1\text{H} = 1e + 1p + 0n$$

Total sub atomic particles of  $\text{H}_2\text{O} = 10e + 10p + 8n = 28$  particles

As shown by the formula of one molecule of water

(i) 1 molecule of  $\text{H}_2\text{O}$  contains bonds = 2

1 mol  $\text{H}_2\text{O}$  molecules contains bonds = 2 moles

(ii) 1 molecule of  $\text{H}_2\text{O}$  contains atoms = 3

1 mole  $\text{H}_2\text{O}$  molecules contain atoms = 3 moles

(iii) 1  $\text{H}_2\text{O}$  molecule contains electrons =  $8 + 1 + 1 = 10$

1 mol  $\text{H}_2\text{O}$  molecule contain electrons = 10 moles

(iv) 1  $\text{H}_2\text{O}$  molecule has subatomic particles = 28

1 mol  $\text{H}_2\text{O}$  molecule have subatomic particles = 28 moles

- (vi)  $\text{N}_2$  and  $\text{CO}$  have the same number of electrons, protons and neutrons.

Ans.	$\text{N}_2$	$\text{CO}$
	$^{14}_7\text{N} = 7e + 7p + 7n$	$^{12}_6\text{C} = 6e + 6p + 6n$
		$^{16}_8\text{O} = 8e + 8p + 8n$
	One molecule of nitrogen contains two nitrogen atoms	One molecule of carbon monoxide contains one carbon and one oxygen atom

so

$$\text{N} = 7e + 7p + 7n$$

$$+ \text{N} = 7e + 7p + 7n$$

$$\text{Total } \text{N}_2 = 14e + 14p + 14n$$

so

$$\text{C} = 6e + 6p + 6n$$

$$+ \text{O} = 8e + 8p + 8n$$

$$\text{Total } \text{CO} = 14e + 14p + 14n$$

So  $\text{N}_2$  and  $\text{CO}$  have same number of electrons, protons and neutrons.

### NUMERICALS OF EXERCISE

Q10. Calculate each of the following quantities.

- (a) Mass in grams of 2.74 moles of  $\text{KMnO}_4$ .

Ans. Given data

$$\text{Number of moles of } \text{KMnO}_4 = 2.74 \text{ moles}$$

Required

$$\text{Mass in grams of } \text{KMnO}_4 = ?$$

Solution

$$\text{Number of moles of } \text{KMnO}_4 = \frac{\text{Mass of } \text{KMnO}_4}{\text{Formula mass of } \text{KMnO}_4}$$

$$\text{Formula mass of } \text{KMnO}_4 = 39 + 55 + 64 = 158 \text{ g mol}^{-1}$$

$$2.74 = \frac{\text{Mass of } \text{KMnO}_4}{158}$$

$$158 \times 2.74 = \text{Mass of } \text{KMnO}_4$$

$$432.92 \text{ g} = \text{Mass of } \text{KMnO}_4$$

- (b) Moles of O atoms in 9.00 g of  $\text{Mg}(\text{NO}_3)_2$ .

Ans. Given data

$$\text{Mass of } \text{Mg}(\text{NO}_3)_2 = 9.00 \text{ g}$$

Required

$$\text{Number of moles of O atoms} = ?$$

Solution

$$\text{Number of moles of } \text{Mg}(\text{NO}_3)_2 = \frac{\text{Mass of } \text{Mg}(\text{NO}_3)_2}{\text{Molar mass of } \text{Mg}(\text{NO}_3)_2}$$

$$\text{Molar Mass of } \text{Mg}(\text{NO}_3)_2 = 24 + 2(14 + 3 \times 16) = 148 \text{ g mol}^{-1}$$

$$\text{Number of moles of } \text{Mg}(\text{NO}_3)_2 = \frac{9}{148} = 0.06 \text{ moles}$$

$$1 \text{ mole of } \text{Mg}(\text{NO}_3)_2 \text{ contains O atoms} = 6 \text{ moles}$$

$$0.06 \text{ moles of } \text{Mg}(\text{NO}_3)_2 \text{ contain moles of O atoms} = 0.06 \times 6 = 0.36 \text{ atoms}$$

- (c) Number of O atoms in 10.037 g of  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

Ans. Given data

$$\text{Mass of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O} = 10.037 \text{ g}$$

Required

$$\text{Number of O atoms} = ?$$

Solution

$$\text{Number of moles of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O} = \frac{\text{Mass of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O}}{\text{Molar mass of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O}}$$



$$\text{Molar Mass of CuSO}_4 \cdot 5\text{H}_2\text{O} = 63.5 + 32 + (4 \times 16) + (5 \times 18) = 249.5 \text{ g mol}^{-1}$$

$$\text{Number of moles of CuSO}_4 \cdot 5\text{H}_2\text{O} = \frac{10.037}{249.5}$$

$$= 0.04 \text{ moles}$$

$$1 \text{ moles of CuSO}_4 \cdot 5\text{H}_2\text{O} \text{ have moles of O} = 9 \text{ moles}$$

$$0.04 \text{ moles of CuSO}_4 \cdot 5\text{H}_2\text{O} \text{ have moles of O} = 9 \times 0.04$$

$$= 0.36 \text{ moles}$$

$$\text{Number of Oxygen atoms} = \text{Number of moles of oxygen atoms} \times N_A$$

$$= 0.36 \times 6.02 \times 10^{23}$$

$$\text{Number of Oxygen atoms} = 2.167 \times 10^{23} \text{ atoms}$$

(d) Mass in kilograms of  $2.6 \times 10^{20}$  molecules of  $\text{SO}_2$ .

Ans. Given data

$$\text{Number of SO}_2 \text{ molecules} = 2.6 \times 10^{20} \text{ molecules}$$

Required

$$\text{Mass in kg of SO}_2 \text{ molecules} = ?$$

Solution

$$\text{Number of SO}_2 \text{ molecules} = \frac{\text{Mass of SO}_2 \text{ molecules}}{\text{Molar mass of SO}_2} \times N_A$$

$$\text{Molar Mass of SO}_2 = 32 + (16 \times 2) = 64 \text{ g mol}^{-1}$$

$$2.6 \times 10^{20} = \frac{\text{Mass of SO}_2 \text{ molecules}}{64} \times 6.02 \times 10^{23}$$

$$\frac{2.6 \times 10^{20} \times 64}{6.02 \times 10^{23}} = \text{Mass of SO}_2 \text{ molecules}$$

$$27.641 \times 10^{-3} \text{ g} =$$

$$\frac{27.641 \times 10^{-3}}{1000} =$$

$$2.764 \times 10^{-5} \text{ kg} = \text{Mass of SO}_2 \text{ molecules}$$

(e) Moles of Cl atoms in 0.822 g  $\text{C}_2\text{H}_4\text{Cl}_2$ .

Ans. Given data

$$\text{Mass of C}_2\text{H}_4\text{Cl}_2 = 0.822 \text{ g}$$

Required

$$\text{Number of moles of Cl atoms} = ?$$

Solution

$$\text{Number of moles of C}_2\text{H}_4\text{Cl}_2 = \frac{\text{Mass of C}_2\text{H}_4\text{Cl}_2}{\text{Molar mass of C}_2\text{H}_4\text{Cl}_2}$$

$$\text{Molar Mass of C}_2\text{H}_4\text{Cl}_2 = 24 + 4 + 71 = 99 \text{ g mol}^{-1}$$

$$\text{Number of moles of C}_2\text{H}_4\text{Cl}_2 = \frac{0.822}{99}$$

$$= 0.0083 \text{ moles}$$

$$1 \text{ mole of C}_2\text{H}_4\text{Cl}_2 \text{ contains moles of Cl atoms} = 2 \text{ moles}$$

$$0.0083 \text{ moles of C}_2\text{H}_4\text{Cl}_2 \text{ contain moles of Cl atoms} = 0.0083 \times 2 = 0.0166 \text{ moles}$$

# Scholar's CHEMISTRY – XI (Subjective)

31

(f) Mass in grams of 5.136 moles of  $\text{Ag}_2\text{CO}_3$ .

Ans. Given data

$$\text{Moles of Ag}_2\text{CO}_3 = 5.136 \text{ moles}$$

Required

$$\text{Mass of Ag}_2\text{CO}_3 = ?$$

Solution

$$\text{Number of moles of Ag}_2\text{CO}_3 = \frac{\text{Mass of Ag}_2\text{CO}_3}{\text{Formula Mass of Ag}_2\text{CO}_3}$$

$$\text{Formula mass of Ag}_2\text{CO}_3 = (108 \times 2) + 12 + (16 \times 3) = 276 \text{ g mol}^{-1}$$

$$5.136 = \frac{\text{Mass of Ag}_2\text{CO}_3}{276}$$

$$5.136 \times 276 = \text{Mass of Ag}_2\text{CO}_3$$

$$1417.54 \text{ g} = \text{Mass of Ag}_2\text{CO}_3$$

(g) Mass in grams of  $2.78 \times 10^{21}$  molecules of  $\text{CrO}_2\text{Cl}_2$ .

Ans. Given data

$$\text{Number of molecules of CrO}_2\text{Cl}_2 = 2.78 \times 10^{21} \text{ molecules}$$

Required

$$\text{Mass of molecules of CrO}_2\text{Cl}_2 \text{ in gram} = ?$$

Solution

$$\text{Number of molecules of CrO}_2\text{Cl}_2 = \frac{\text{Mass of CrO}_2\text{Cl}_2}{\text{Molar mass of CrO}_2\text{Cl}_2} \times N_A$$

$$\text{Molar Mass of CrO}_2\text{Cl}_2 = 52 + 32 + 71 = 155 \text{ g mol}^{-1}$$

$$2.78 \times 10^{21} = \frac{\text{Mass of CrO}_2\text{Cl}_2}{155} \times 6.02 \times 10^{23}$$

$$\frac{2.78 \times 10^{21} \times 155}{6.02 \times 10^{23}} = \text{Mass of CrO}_2\text{Cl}_2$$

$$0.7158 \text{ g} = \text{Mass of CrO}_2\text{Cl}_2$$

(h) Number of moles and formula units in 100 g of  $\text{KClO}_3$ .

Ans. Given data

$$\text{Mass of KClO}_3 = 100 \text{ g}$$

Required

$$\text{Number of moles of KClO}_3 = ?$$

$$\text{Number of formula units of KClO}_3 = ?$$

Solution

$$\text{Number of moles of KClO}_3 = \frac{\text{Mass of KClO}_3}{\text{Molar mass of KClO}_3}$$

$$\text{Molar Mass of KClO}_3 = 39 + 35.5 + 48 = 122.5 \text{ g mol}^{-1}$$

$$\text{Number of moles of KClO}_3 = \frac{100}{122.5} = 0.816 \text{ moles}$$

$$1 \text{ mole of KClO}_3 \text{ contains formula units} = 6.02 \times 10^{23}$$

$$0.816 \text{ moles of KClO}_3 \text{ contain formula units} = 0.816 \times 6.02 \times 10^{23}$$

$$= 4.91 \times 10^{23} \text{ formula units}$$



(i) Number of  $K^+$  ions,  $ClO_3^-$  ions, Cl atoms, and O atoms in 100g of  $KClO_3$  (h).

Ans. 1 mole of  $KClO_3$  contains  $K^+$  ions  $= 6.02 \times 10^{23}$   
 0.816 moles of  $KClO_3$  contain  $K^+$  ions  $= 0.816 \times 6.02 \times 10^{23}$   
 $= 4.91 \times 10^{23}$  ions  
 1 mole of  $KClO_3$  contains  $ClO_3^-$  ions  $= 6.02 \times 10^{23}$   
 0.816 moles of  $KClO_3$  contain  $ClO_3^-$  ions  $= 0.816 \times 6.02 \times 10^{23} = 4.91 \times 10^{23}$  ions

Similarly,

1 mole of  $KClO_3$  contains Cl atoms  $= 6.02 \times 10^{23}$   
 0.816 mol of  $KClO_3$  contain Cl atoms  $= 4.91 \times 10^{23}$  atoms  
 1 mole of  $KClO_3$  contains O atoms  $= 3 \times 6.02 \times 10^{23}$  atoms  
 0.816 moles of  $KClO_3$  contain O atoms  $= 0.816 \times 3 \times 6.02 \times 10^{23}$   
 $= 1.474 \times 10^{24}$  atoms

Q11. Aspartame, the artificial sweetener, has a molecular formula of  $C_{14}H_{18}N_2O_5$ .

(a) What is the mass of one mole of aspartame?

Ans. 1 mole of  $C_{14}H_{18}N_2O_5$   $= (12 \times 14) + (1 \times 18) + (14 \times 2) + (16 \times 5)$   
 $= 168 + 18 + 28 + 80$   
 $= 294 \text{ g mol}^{-1}$

(b) How many moles are present in 52 g of aspartame?

Ans. Given data

Mass of aspartame  $= 52 \text{ g}$

Required

Number of moles of aspartame  $= ?$

Solution

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

Molar mass of  $C_{14}H_{18}N_2O_5$   $= (12 \times 14) + 18 + 28 + (16 \times 5) = 294 \text{ g mol}^{-1}$

Number of moles of aspartame  $= \frac{52}{294} = 0.177 \text{ moles}$

(c) What is the mass in grams of 10.122 moles of aspartame?

Ans. Given data

moles of Aspartame  $= 10.122 \text{ moles}$

Required

Mass of aspartame  $= ?$

Solution

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

Molar mass of  $C_{14}H_{18}N_2O_5$   $= (12 \times 14) + 18 + 28 + (16 \times 5) = 294 \text{ g mol}^{-1}$

10.122  $= \frac{\text{Mass of aspartame}}{294}$

10.122  $\times$  294  $=$  Mass of aspartame

2975.87 g  $=$  Mass of aspartame

(d) How many hydrogen atoms are present in 2.43 g of aspartame?

Ans. Given data

Molecular formula of Aspartame  $= C_{14}H_{18}N_2O_5$

Mass of Aspartame  $= 2.43 \text{ g}$

Required

Number of hydrogen atoms  $= ?$

Solution

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

Molar mass of  $C_{14}H_{18}N_2O_5$   $= (12 \times 14) + (1 \times 18) + (2 \times 14) + (16 \times 5)$

$= 168 + 18 + 28 + 80$

$= 294 \text{ g mol}^{-1}$

$= \frac{2.43}{294} = 0.00826 \text{ moles}$

Moles of hydrogen in one mole of aspartame  $= 18 \text{ moles}$

0.00826 moles of aspartame have moles of hydrogen  $= 0.00826 \times 18 = 0.1487$

Number of atoms of hydrogen  $= \text{Number of moles of aspartame} \times N_A$

$= 0.1487 \times 6.02 \times 10^{23}$

$= 0.8951 \times 10^{23}$

Number of atoms of hydrogen  $= 8.96 \times 10^{22} \text{ atoms}$

Q12. A sample of 0.600 moles of a metal M reacts completely with excess of fluorine to form 46.8 g of  $MF_2$ .

(a) How many moles of F are present in the sample of  $MF_2$  that forms?

(b) Which element is represented by the symbol M?

Ans. Given data

Number of moles of M  $= 0.600 \text{ moles}$

Mass of  $MF_2$   $= 46.8 \text{ g}$

Required

Number of moles of F  $= ?$

Actual name of Metal M  $= ?$

Solution

$M + F_2 \longrightarrow MF_2$

Comparison between number of moles of M and  $MF_2$

M  $MF_2$

1  $1$

0.6  $0.6$

Number of moles of  $MF_2$   $= 0.6 \text{ moles}$

1 mole of  $MF_2$  contain moles of F atoms  $= 2$

0.6 moles of  $MF_2$  contain moles of F atoms  $= 0.6 \times 2 = 1.2 \text{ moles}$



$$\text{Number of moles of MF}_2 = \frac{\text{Mass of MF}_2}{\text{Molar mass of MF}_2}$$

$$0.6 \text{ mol} = \frac{46.8 \text{ g}}{\text{Molar mass of MF}_2}$$

$$\text{Molar mass of MF}_2 = \frac{46.8 \text{ g}}{1.2 \text{ mol}} = 78 \text{ g mol}^{-1}$$

$$\text{Atomic mass of M} + (19 \times 2) = 78 \text{ g mol}^{-1}$$

$$M + 38 = 78$$

$$M = 78 - 38$$

$$\text{Atomic mass of M} = 40 \text{ g mol}^{-1}$$

$$\text{So, Actual name of Metal M} = \text{Calcium (Ca)}$$

(Atomic Mass of F = 19)

Q13. In each pair, choose the larger of the indicated quantity, or state if the samples are equal.

(a) Individual particles 0.4 moles of oxygen molecules or 0.4 moles of oxygen atoms.

Ans. Both are equal.

Reason

0.4 moles of  $\text{O}_2$  molecules and 0.4 moles of O atoms, both are equimolar quantities so, they have equal number of particles. i.e.

$$0.4 \times 6.02 \times 10^{23} = 2.408 \times 10^{23} \text{ moles}$$

(d) Individual particles 4.0 g of  $\text{N}_2\text{O}_4$  or 3.3 g of  $\text{SO}_2$ .

Given data

$$\text{Mass of N}_2\text{O}_4 = 4 \text{ g}$$

$$\text{Mass of SO}_2 = 3.3 \text{ g}$$

Required

$$\text{Individual particles N}_2\text{O}_4 = ?$$

$$\text{Individual particles SO}_2 = ?$$

Solution

$$(i) \quad \text{Number of molecules of N}_2\text{O}_4 = \frac{\text{Mass of N}_2\text{O}_4}{\text{Molar mass of N}_2\text{O}_4} \times N_A$$

$$\text{Molar Mass of N}_2\text{O}_4 = (14 \times 2) + (16 \times 4) = 92 \text{ g mol}^{-1}$$

$$= \frac{4}{92} \times 6.02 \times 10^{23}$$

$$= 0.258 \times 10^{23}$$

$$\text{Number of molecules of N}_2\text{O}_4 = 2.58 \times 10^{22} \text{ molecules}$$

$$(ii) \quad \text{Number of molecules of SO}_2 = \frac{\text{Mass of SO}_2}{\text{Molar mass of SO}_2} \times N_A$$

$$\text{Molar Mass of SO}_2 = 32 + 32 = 64 \text{ g mol}^{-1}$$

$$= \frac{3.3}{64} \times 6.02 \times 10^{23}$$

$$= 0.3104 \times 10^{23}$$

$$\text{Number of molecules of SO}_2 = 3.1 \times 10^{22} \text{ molecules}$$

3.3 g of  $\text{SO}_2$  has larger number of individual particles.

(e) Total ions 2.3 moles of  $\text{NaClO}_3$  or 2.0 moles of  $\text{MgCl}_2$ ?

Ans. Given data

$$\text{Number of moles of NaClO}_3 = 2.3 \text{ moles}$$

$$\text{Number of moles of MgCl}_2 = 2 \text{ moles}$$

Required

$$\text{Number of ions in 2.3 moles of NaClO}_3 = ?$$

$$\text{Number of ions in 2 moles of MgCl}_2 = ?$$

Solution

$$(i) \quad 1 \text{ mole of NaClO}_3 \text{ contains Na}^+ \text{ ions} = 6.02 \times 10^{23}$$

$$2.3 \text{ moles of NaClO}_3 \text{ contain Na}^+ \text{ ions} = 2.3 \times 6.02 \times 10^{23}$$

$$= 13.846 \times 10^{23} \text{ ions}$$

$$1 \text{ mole of NaClO}_3 \text{ contains ClO}_3^- \text{ ions} = 6.02 \times 10^{23}$$

$$2.3 \text{ moles of NaClO}_3 \text{ contain ClO}_3^- \text{ ions} = 2.3 \times 6.02 \times 10^{23}$$

$$= 13.846 \times 10^{23} \text{ ions}$$

$$\text{Total ions of NaClO}_3 = \text{Na}^+ \text{ ions} + \text{ClO}_3^- \text{ ions}$$

$$= 13.846 \times 10^{23} + 13.846 \times 10^{23}$$

$$= 2.7692 \times 10^{24} \text{ ions}$$

$$(ii) \quad 1 \text{ mole of MgCl}_2 \text{ contain Mg}^{2+} \text{ ions} = 6.02 \times 10^{23}$$

$$2 \text{ mole of MgCl}_2 \text{ contain Mg}^{2+} \text{ ions} = 2 \times 6.02 \times 10^{23}$$

$$= 12.04 \times 10^{23} \text{ ions}$$

$$1 \text{ mole of MgCl}_2 \text{ contain Cl}^- \text{ ions} = 2 \times 6.02 \times 10^{23}$$

$$2 \text{ moles of MgCl}_2 \text{ contain Cl}^- \text{ ions} = 2 \times 2 \times 6.02 \times 10^{23}$$

$$= 24.04 \times 10^{23}$$

$$\text{Total ions of MgCl}_2 = \text{Mg}^{2+} \text{ ions} + \text{Cl}^- \text{ ions}$$

$$= 12.04 \times 10^{23} + 24.04 \times 10^{23}$$

$$= 3.612 \times 10^{24} \text{ ions}$$

2 moles of  $\text{MgCl}_2$  contain larger number of ions.

(f) Molecules 11.0 g  $\text{H}_2\text{O}$  or 11.0 g  $\text{H}_2\text{O}_2$ .

Ans. Given data

$$\text{Mass of H}_2\text{O} = 11.0 \text{ g}$$

$$\text{Mass of H}_2\text{O}_2 = 11.0 \text{ g}$$

Required

$$\text{Number of molecules in 11 g of H}_2\text{O} = ?$$

$$\text{Number of molecules in 11g of H}_2\text{O}_2 = ?$$

Solution

$$(i) \quad \text{Number of molecules of H}_2\text{O} = \frac{\text{Mass of H}_2\text{O}}{\text{Molar mass of H}_2\text{O}} \times N_A$$

$$\text{Molar Mass of H}_2\text{O} = 2 + 16 = 18 \text{ g mol}^{-1}$$

$$\text{Number of molecules of H}_2\text{O} = \frac{11}{18} \times 6.02 \times 10^{23}$$

$$= 3.67 \times 10^{23} \text{ molecules}$$

$$(ii) \quad \text{Number of molecules of H}_2\text{O}_2 = \frac{\text{Mass of H}_2\text{O}_2}{\text{Molar mass of H}_2\text{O}_2} \times N_A$$

$$\text{Molar Mass of H}_2\text{O}_2 = 2 + 32 = 34 \text{ g mol}^{-1}$$

$$\text{Number of moles of H}_2\text{O}_2 = \frac{11}{34} \times 6.02 \times 10^{23}$$

$$= 1.926 \times 10^{23} \text{ molecules}$$

11g of  $\text{H}_2\text{O}$  contain larger number of molecules.



(g)  $\text{Na}^+$  ion 0.500 moles of NaBr or 0.0145 kg of NaCl.

Ans. Given data

Moles of NaBr = 0.500 moles

Mass of NaCl = 0.0145 kg = 14.5 g

Required

(i)  $\text{Na}^+$  ions in 0.500 moles of NaBr = ?

(ii)  $\text{Na}^+$  ions in 0.0145 kg of NaCl = ?

Solution

(i) Number of formula units of NaBr = Number of moles of NaBr  $\times N_A$

$$= 0.5 \times 6.02 \times 10^{23}$$

Number of formula units of NaBr =  $3.01 \times 10^{23}$  formula units

Since one NaBr contains one  $\text{Na}^+$  and one  $\text{Br}^-$ , So

$$\text{Number of } \text{Na}^+ \text{ ions} = 1 \times 3.01 \times 10^{23} \text{ ions}$$

$$\text{Number of } \text{Na}^+ \text{ ions} = 3.01 \times 10^{23} \text{ ions}$$

(ii) Number of formula units of NaCl =  $\frac{\text{Mass of NaCl}}{\text{Formula Mass of NaCl}} \times N_A$

$$\text{Molar Mass of NaCl} = 23 + 35.5 = 58.5 \text{ g mol}^{-1}$$

$$\text{Number of formula units of NaCl} = \frac{14.5}{58.5} \times 6.02 \times 10^{23}$$

$$\text{Number of formula units of NaCl} = 1.49 \times 10^{23} \text{ formula units}$$

Since one NaCl contains one  $\text{Na}^+$  and one  $\text{Cl}^-$ , So

$$\text{Number of } \text{Na}^+ \text{ ions} = 1 \times 1.49 \times 10^{23} \text{ ions}$$

$$\text{Number of } \text{Na}^+ \text{ ions} = 1.49 \times 10^{23} \text{ ions}$$

Number of  $\text{Na}^+$  ions in 0.5 moles of NaBr is larger.

(h) Mass  $6.02 \times 10^{23}$  atoms of  $^{235}\text{U}$  or  $6.02 \times 10^{23}$  atoms of  $^{238}\text{U}$ .

Ans. Given data

Number of atoms =  $6.02 \times 10^{23}$  atoms

Required

Mass of atoms = ?

Solution

Isotopic mass of  $^{235}\text{U}$  = 235 (1 mole)

Therefore,

$$6.02 \times 10^{23} \text{ atoms of } ^{235}\text{U} \text{ have mass} = 235 \text{ g}$$

Similarly,

$$6.02 \times 10^{23} \text{ atoms of } ^{238}\text{U} \text{ have mass} = 238 \text{ g}$$

$$6.02 \times 10^{23} \text{ atoms of } ^{238}\text{U} \text{ have larger mass.}$$

Q14. (a) Calculate the percentage of nitrogen in the four important fertilizers i.e.,

(i)  $\text{NH}_3$  (ii)  $\text{NH}_2\text{CONH}_2$  (urea) (iii)  $(\text{NH}_4)_2\text{SO}_4$  (iv)  $\text{NH}_4\text{NO}_3$ .

(i)  $\text{NH}_3$

$$\text{Percentage of nitrogen} = \frac{\text{Mass of nitrogen}}{\text{Formula mass of compound}} \times 100$$

$$\text{Molar mass of } \text{NH}_3 = 14 + (3 \times 1) = 14 + 3 = 17 \text{ g mol}^{-1}$$

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

(ii)  $\text{NH}_2\text{CONH}_2$

$$\text{Molar mass of } \text{NH}_2\text{CONH}_2 = 14 + 2 + 12 + 16 + 14 + 2 = 60 \text{ g mol}^{-1}$$

$$\text{Percentage of nitrogen} = \frac{28}{60} \times 100 = 46.67\%$$

(iii)  $(\text{NH}_4)_2\text{SO}_4$

$$\text{Molar mass of } (\text{NH}_4)_2\text{SO}_4 = (14 \times 2) + (1 \times 8) + 32 + (16 \times 4) = 132 \text{ g mol}^{-1}$$

$$\text{Percentage of nitrogen} = \frac{28}{132} \times 100 = 21.21\%$$

(iv)  $\text{NH}_4\text{NO}_3$

$$\text{Molar mass of } \text{NH}_4\text{NO}_3 = 14 + 4 + 14 + 48 = 80 \text{ g mol}^{-1}$$

$$\text{Percentage of nitrogen} = \frac{28}{80} \times 100 = 35\%$$

(b) Calculate the percentage of Nitrogen and Phosphorus in each of the following

(i)  $\text{NH}_4\text{H}_2\text{PO}_4$

$$\text{Formula mass of } \text{NH}_4\text{H}_2\text{PO}_4 = 14 + 4 + 2 + 31 + 64 = 115 \text{ g mol}^{-1}$$

$$\text{Percentage of nitrogen} = \frac{14}{115} \times 100 = 12.17\%$$

$$\text{Percentage of phosphorus} = \frac{31}{115} \times 100 = 26.96\%$$

(ii)  $(\text{NH}_4)_2\text{HPO}_4$

$$\text{Molar mass of } (\text{NH}_4)_2\text{HPO}_4 = (14 \times 2) + (1 \times 8) + 1 + 31 + 64 = 132 \text{ g mol}^{-1}$$

$$\text{Percentage of nitrogen} = \frac{28}{132} \times 100 = 21.21\%$$

$$\text{Percentage of phosphorus} = \frac{31}{132} \times 100 = 23.48\%$$

(iii)  $(\text{NH}_4)_3\text{PO}_4$

$$\text{Molar mass of } (\text{NH}_4)_3\text{PO}_4 = (14 \times 3) + 12 + 31 + 64 = 149 \text{ g mol}^{-1}$$

$$\text{Percentage of nitrogen} = \frac{42}{149} \times 100 = 28.19\%$$

$$\text{Percentage of phosphorus} = \frac{31}{149} \times 100 = 20.81\%$$

Q15. Glucose  $\text{C}_6\text{H}_{12}\text{O}_6$  is the most important nutrient in the cell for generating chemical potential energy. Calculate the mass % of each element in glucose and determine the number of C, H and O atoms in 10.5 g of the sample.

Ans. Given data

Mass of Glucose = 10.5 g

Required

Percentage of C = ?

Percentage of H = ?

Percentage of O = ?

Number of carbon atoms = ?

Number of hydrogen atoms = ?

Number of oxygen atoms = ?



## Solution

$$\begin{aligned}\text{Molar mass of } C_6H_{12}O_6 &= (6 \times 12) + (12 \times 1) + (6 \times 16) \\ &= 72 + 12 + 96 = 180 \text{ g mol}^{-1}\end{aligned}$$

$$\text{Percentage of an element} = \frac{\text{Mass of element}}{\text{Molar mass}} \times 100$$

$$\text{Percentage of carbon} = \frac{72}{180} \times 100 = 40\%$$

$$\text{Percentage of hydrogen} = \frac{12}{180} \times 100 = 6.66\%$$

$$\text{Percentage of oxygen} = \frac{96}{180} \times 100 = 53.33\%$$

$$\begin{aligned}\text{Number of molecules of Glucose} &= \frac{\text{Mass}}{\text{Molar mass}} \times N_A \\ &= \frac{10.5}{180} \times 6.02 \times 10^{23}\end{aligned}$$

$$\text{Number of molecules of } C_6H_{12}O_6 = 0.351 \times 10^{23} \text{ molecules}$$

$$\begin{aligned}1 \text{ molecule of glucose contain carbon atoms} &= 6 \times 0.351 \times 10^{23} \\ &= 2.107 \times 10^{23} \text{ atoms}\end{aligned}$$

$$\begin{aligned}1 \text{ molecule of glucose contain hydrogen atoms} &= 12 \times 0.351 \times 10^{23} \\ &= 4.212 \times 10^{23} \text{ atoms}\end{aligned}$$

$$\begin{aligned}1 \text{ molecule of glucose contain oxygen atoms} &= 6 \times 0.351 \times 10^{23} \\ &= 2.107 \times 10^{23} \text{ atoms}\end{aligned}$$

**Q16.** Ethylene glycol is used as automobile antifreeze. It has 38.7% carbon, 9.7% hydrogen and 51.6% oxygen. Its molar mass is 62.1 g mol<sup>-1</sup>. Determine its empirical formula.

## Ans. Given data

$$\text{Percentage of carbon} = 38.7\%$$

$$\text{Percentage of hydrogen} = 9.7\%$$

$$\text{Percentage of oxygen} = 51.6\%$$

## Required

$$\text{Empirical formula} = ?$$

## Solution

## Number of gram atoms

$$\text{Number of gram atoms of element} = \frac{\text{Percentage of element}}{\text{Atomic mass of element}}$$

$$\text{Number of gram atoms of carbon} = \frac{38.7}{12} = 3.225 \text{ mole}$$

$$\text{Number of gram atoms of hydrogen} = \frac{9.7}{1.008} = 9.6 \text{ mole}$$

$$\text{Number of gram atoms of oxygen} = \frac{51.6}{16} = 3.225 \text{ mole}$$

## Atomic ratio

$$\text{Atomic ratio of element} = \frac{\text{No. of gram atoms of element}}{\text{Smallest number}}$$

$$\text{Atomic ratio of carbon} = \frac{3.225}{3.225} = 1$$

$$\text{Atomic ratio of hydrogen} = \frac{9.6}{3.225} = 3$$

## Scholar's CHEMISTRY – XI (Subjective)

$$\text{Atomic ratio of oxygen} = \frac{3.225}{3.225} = 1$$

$$\text{Empirical formula} = CH_3O$$

**Q17.** Serotonin (Molar mass = 176 g mol<sup>-1</sup>) is a compound that conducts nerve impulses in brain and muscles. It contains 68.2% C, 6.86% H, 15.09% N and 9.08% O. What is its molecular formula?

## Ans. Given data

$$\text{Percentage of carbon} = 68.2\%$$

$$\text{Percentage of hydrogen} = 6.86\%$$

$$\text{Percentage of nitrogen} = 15.09\%$$

$$\text{Percentage of oxygen} = 9.08\%$$

$$\text{Molar mass} = 176 \text{ g mol}^{-1}$$

## Required

$$\text{Molecular formula} = ?$$

## Solution

## Number of gram atoms

$$\text{Number of gram atoms of element} = \frac{\text{Percentage of element}}{\text{Atomic mass of element}}$$

$$\text{Number of gram atoms of carbon} = \frac{68.2}{12} = 5.68 \text{ mole}$$

$$\text{Number of gram atoms of hydrogen} = \frac{6.86}{1.008} = 6.80 \text{ mole}$$

$$\text{Number of gram atoms of nitrogen} = \frac{15.09}{14} = 1.08 \text{ mole}$$

$$\text{Number of gram atoms of oxygen} = \frac{9.08}{16} = 0.5675 \text{ mole}$$

## Atomic ratio

$$\text{Atomic ratio of element} = \frac{\text{No. of gram atoms of element}}{\text{Smallest number}}$$

$$\text{Atomic ratio of C} = \frac{5.68}{0.5675} = 10$$

$$\text{Atomic ratio of H} = \frac{6.80}{0.5675} = 12$$

$$\text{Atomic ratio of N} = \frac{1.08}{0.5675} = 2$$

$$\text{Atomic ratio of O} = \frac{0.5675}{0.5675} = 1$$

$$\text{Empirical formula} = C_{10}H_{12}N_2O$$

## Molecular formula

$$\text{Molecular formula} = n \times (\text{Empirical formula}) \dots\dots (i)$$

Also we know that

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

$$\text{Empirical formula mass} = C_{10}H_{12}N_2O$$



$$= (12 \times 10) + (1 \times 12) + (14 \times 2) + 16$$

$$= 176$$

$$n = \frac{176}{176} = 1$$

Putting the value of  $n = 1$  in eq (i)

Molecular formula  $= 1 \times (C_{10}H_{12}N_2O)$

Molecular formula  $= C_{10}H_{12}N_2O$

**Q18.** An unknown metal M reacts with S to form a compound with a formula  $M_2S_3$ . If 3.12 g of M reacts with exactly 2.88 g of sulphur, what are the names of metal M and the compound  $M_2S_3$ ?

**Ans.** Given data

Mass of metal M  $= 3.12$  g

Mass of S  $= 2.88$  g

**Required**

Name of metal M  $= ?$

Name of compound  $M_2S_3$   $= ?$

**Solution**



$$\begin{aligned} \text{Number of moles of sulphur} &= \frac{\text{Mass of sulphur}}{\text{Atomic mass of sulphur}} \\ &= \frac{2.88}{32} \quad (\text{Atomic mass of S} = 32 \text{ g/mol}) \\ &= 0.09 \text{ moles} \end{aligned}$$

Number of moles of S  $= 0.09$  moles

Comparison between moles of S and moles of M

S	M
3	2
1	2/3
0.09	$2/3 \times 0.09$
0.09	0.06

Number of moles of M  $= 0.06$  moles

$$\begin{aligned} \text{Number of moles of M} &= \frac{\text{Mass of M}}{\text{Atomic mass of M}} \\ 0.06 \text{ moles} &= \frac{3.12}{\text{Atomic mass of M}} \end{aligned}$$

$$\text{Atomic mass of M} = \frac{3.12}{0.06}$$

$$\text{Atomic mass of M} = 52 \text{ g mol}^{-1}$$

From the Atomic mass of M, it is clear that

Name of M  $= \text{Chromium (Cr)}$

Name of  $M_2S_3$   $= \text{Cr}_2S_3$  (Chromium sulphide)

**Q19.** The octane present in gasoline burns according to the following equation



**(a)** How many moles of  $O_2$  are needed to react fully with 4 moles of octane?

**Ans.** Given data

Number of moles of octane  $= 4$  moles

**Required**

Number of moles of  $O_2 = ?$

**Solution**



Comparison between moles of  $C_8H_{18}$  and  $O_2$

$C_8H_{18}$	$O_2$
2	25
1	25/2
4	$25/2 \times 4$
4	50

Number of moles of  $O_2 = 50$  moles

**(b)** How many moles of  $CO_2$  can be produced from one mole of octane?

**Ans.** Given data

Number of moles of octane  $= 1$  mole

**Required**

Number of moles of  $CO_2 = ?$

**Solution**



Comparison between moles of  $C_8H_{18}$  and moles of  $CO_2$

$C_8H_{18}$	$CO_2$
2	16
1	16/2
1	8

Number of moles of  $CO_2 = 8$  moles

**(c)** How many moles of water are produced by the combustion of 6 moles of octane?

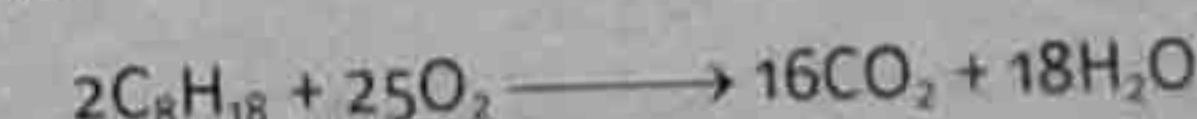
**Ans.** Given data

Number of moles of octane  $= 6$  moles

**Required**

Number of moles of  $H_2O = ?$

**Solution**



Comparison between moles of  $C_8H_{18}$  and moles of  $H_2O$

$C_8H_{18}$	$H_2O$
2	18
1	18/2
6	$18/2 \times 6$
6	54

Number of moles of  $H_2O = 54$  moles

**(d)** If this reaction is to be used to synthesize 8 moles of  $CO_2$ , how many grams of oxygen are needed? How many grams of octane will be used?

**Ans.** Given data

Number of moles of  $CO_2 = 8$  moles

**Required**

Mass of  $O_2 = ?$



Mass of  $C_8H_{18}$  = ?

Solution

Comparison between moles of  $CO_2$  and moles of  $O_2$ 

$CO_2$	:	$O_2$
16	:	25
1	:	$25/16$
8	:	$25/16 \times 8$
8	:	12.5

Number of moles of  $O_2$  : 12.5 molesMolar Mass of  $O_2$  =  $(16 \times 2) = 32 \text{ g mol}^{-1}$ Number of moles =  $\frac{\text{Mass of oxygen}}{\text{Molar mass of oxygen}}$ 12.5 moles =  $\frac{\text{Mass of oxygen}}{32 \text{ g mol}^{-1}}$  $12.5 \times 32$  = Mass of oxygen

400 g = Mass of oxygen

Comparison between moles of  $CO_2$  and moles of  $C_8H_{18}$ 

$CO_2$	:	$C_8H_{18}$
16	:	2
1	:	$2/16$
8	:	$\frac{2}{16} \times 8$
8	:	1

Number of moles of  $C_8H_{18}$  : 1 moleMolar Mass of  $C_8H_{18}$  =  $(12 \times 8) + (1 \times 18) = 114 \text{ g mol}^{-1}$ Number of moles of  $C_8H_{18}$  =  $\frac{\text{Mass of } C_8H_{18}}{\text{Molar mass of } C_8H_{18}}$ 1 mol =  $\frac{\text{Mass of } C_8H_{18}}{114 \text{ g mol}^{-1}}$  $1 \times 114$  = Mass of  $C_8H_{18}$ 114 g = Mass of  $C_8H_{18}$ 

**Q20.** Calculate the number of grams of  $Al_2S_3$  which can be prepared by the reaction of 20g of Al and 30g of sulphur. How much the non-limiting reactant is in excess?

Ans. Given data

Mass of Aluminium = 20 g

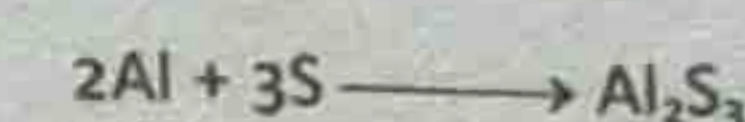
Mass of Sulphur = 30 g

Required

Mass of  $Al_2S_3$  = ?

Non-limiting reactant in excess = ?

Solution



Number of moles of reactant

Number of moles of Al =  $\frac{\text{Mass of Al}}{\text{Atomic mass of Al}}$ Atomic mass of Al =  $27 \text{ g mol}^{-1}$ Number of moles of Al =  $\frac{20}{27} = 0.740 \text{ moles}$ Number of moles of S =  $\frac{\text{Mass of S}}{\text{Atomic mass of S}}$ Molar mass of S =  $32 \text{ g/mol}$ Number of moles of S =  $\frac{30}{32} = 0.9375 \text{ moles}$ 

Number of moles of product

Comparing number of moles of Al and  $Al_2S_3$ 

Al	:	$Al_2S_3$
2	:	1
1	:	$1/2$
0.740	:	$1/2 \times 0.740$
0.740	:	0.37

Number of moles of  $Al_2S_3$  : 0.37Comparing number of moles of S and  $Al_2S_3$ 

S	:	$Al_2S_3$
3	:	1
1	:	$1/3$
0.9375	:	$1/3 \times 0.9375$
0.9375	:	0.3125

Number of moles of  $Al_2S_3$  : 0.3125S is a limiting reactant because it has given less amount of  $Al_2S_3$ .Mass of  $Al_2S_3$ Number of moles of  $Al_2S_3$  =  $\frac{\text{Mass of } Al_2S_3}{\text{Molar mass of } Al_2S_3}$ Molar mass of  $Al_2S_3$  =  $(27 \times 2) + (32 \times 3) = 150 \text{ g mol}^{-1}$ 0.3125 mol =  $\frac{\text{Mass of } Al_2S_3}{150 \text{ g mol}^{-1}}$  $0.3125 \times 150$  = Mass of  $Al_2S_3$ 46.87 g = Mass of  $Al_2S_3$ (ii)  $2Al + 3S \longrightarrow Al_2S_3$ 

Comparison between sulphur and Aluminium

S	:	Al
3	:	2
1	:	$2/3$
0.937	:	$2/3 \times 0.937$
0.937	:	0.624

Number of moles of Al = 0.624 moles

Excess moles of Al =  $0.74 - 0.624$ 

= 0.116 moles



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

$$\text{Atomic mass of Al} = 27 \text{ g mol}^{-1}$$

$$0.116 \text{ mol} = \frac{\text{Mass of Al}}{27 \text{ g mol}^{-1}}$$

$$0.116 \times 27 = \text{Mass of Al}$$

$$3.132 \text{ g} = \text{Mass of Al (Non-limiting reactant in excess)}$$

Q21. A mixture of two liquids, hydrazine  $\text{N}_2\text{H}_4$  and  $\text{N}_2\text{O}_4$  are used as a fuel in rockets. They produce  $\text{N}_2$  and water vapours. How many grams of  $\text{N}_2$  gas will be formed by reacting 100 g of  $\text{N}_2\text{H}_4$  and 200 g of  $\text{N}_2\text{O}_4$ .



Ans. Given data

$$\text{Mass of } \text{N}_2\text{H}_4 = 100 \text{ g}$$

$$\text{Mass of } \text{N}_2\text{O}_4 = 200 \text{ g}$$

Required

$$\text{Mass of } \text{N}_2 = ?$$

Solution



Number of moles of reactant

$$\text{Number of moles of } \text{N}_2\text{H}_4 = \frac{\text{Mass of } \text{N}_2\text{H}_4}{\text{Molar mass of } \text{N}_2\text{H}_4}$$

$$\text{Molar mass of } \text{N}_2\text{H}_4 = (14 \times 2) + (1 \times 4) = 32 \text{ g mol}^{-1}$$

$$\text{Number of moles of } \text{N}_2\text{H}_4 = \frac{100}{32}$$

$$= 3.125 \text{ moles}$$

$$\text{Number of moles of } \text{N}_2\text{O}_4 = \frac{\text{Mass of } \text{N}_2\text{O}_4}{\text{Molar mass of } \text{N}_2\text{O}_4}$$

$$\text{Molar mass of } \text{N}_2\text{O}_4 = (14 \times 2) + (16 \times 4) = 92 \text{ g mol}^{-1}$$

$$\text{Number of moles of } \text{N}_2\text{O}_4 = \frac{200}{92}$$

$$= 2.17 \text{ moles}$$

Number of moles of product

Comparing number of moles of  $\text{N}_2\text{H}_4$  and  $\text{N}_2$

$$\text{N}_2\text{H}_4 : \text{N}_2$$

$$2 : 3$$

$$1 : 3/2$$

$$3.125 : 3/2 \times 3.125$$

$$3.125 : 4.68$$

$$\text{Number of moles of } \text{N}_2 = 4.68 \text{ moles}$$

Comparing number of moles of  $\text{N}_2\text{O}_4$  and  $\text{N}_2$

$$\text{N}_2\text{O}_4 : \text{N}_2$$

$$1 : 3$$

$$2.17 : 3 \times 2.17$$

$$2.17 : 6.51$$

$$\text{Number of moles of } \text{N}_2 = 6.51 \text{ moles}$$

Hydrazine is a limiting reactant because, it has given less amount of  $\text{N}_2$ .

Mass of  $\text{N}_2$

$$\text{Number of moles of } \text{N}_2 = \frac{\text{Mass of } \text{N}_2}{\text{Molar mass of } \text{N}_2}$$

$$\text{Molar mass of } \text{N}_2 = (14 \times 2) = 28 \text{ g mol}^{-1}$$

$$4.68 \text{ mol} = \frac{\text{Mass of } \text{N}_2}{28 \text{ g mol}^{-1}}$$

$$4.68 \times 28 = \text{Mass of } \text{N}_2$$

$$131.04 \text{ g} = \text{Mass of } \text{N}_2$$

Q22. Silicon Carbide (SiC) is an important ceramic material. It is produced by allowing sand ( $\text{SiO}_2$ ) to react with carbon at high temperature.



When 100 kg sand is reacted with excess of carbon, 51.4 kg of SiC is produced. What is the percentage yield of SiC?

Ans. Given data

$$\text{Mass of } \text{SiO}_2 = 100 \text{ kg}$$

$$= 100 \times 1000 = 100,000 \text{ g}$$

$$\text{Mass of SiC} = 51.4 \text{ kg}$$

$$= 51.4 \times 1000 = 51,400 \text{ g}$$

Required

$$\text{Percentage yield of SiC} = ?$$

Solution



Number of moles of reactant

$$\text{Number of moles of } \text{SiO}_2 = \frac{\text{Mass of } \text{SiO}_2}{\text{Molar mass of } \text{SiO}_2}$$

$$\text{Molar mass of } \text{SiO}_2 = 28 + (16 \times 2) = 60 \text{ g mol}^{-1}$$

$$\text{Number of moles of } \text{SiO}_2 = \frac{100,000 \text{ g}}{60 \text{ g mol}^{-1}}$$

$$\text{Number of moles of } \text{SiO}_2 = 1666.66 \text{ moles}$$

Number of moles of product

Comparing number of moles of  $\text{SiO}_2$  and SiC

$$\text{SiO}_2 : \text{SiC}$$

$$1 : 1$$

$$1666.6 : 1666.6$$

$$\text{Number of moles of SiC} = 1666.6 \text{ moles}$$

Mass of SiC

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

$$\text{Molar mass of SiC} = 28 + 12 = 40 \text{ g mol}^{-1}$$



$$\begin{aligned}
 1666.6 \text{ mol} &= \frac{\text{Mass of SiC}}{40 \text{ g mol}^{-1}} \\
 1666.6 \times 40 &= \text{Mass of SiC} \\
 66,666.4 \text{ g} &= \text{Mass of SiC} \\
 \text{Percentage yield} &= \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 \\
 &= \frac{51400}{66,666.4} \times 100 \\
 \text{Percentage yield of SiC} &= 77\%
 \end{aligned}$$

### SOLVED EXAMPLES

#### Example (1)

A sample of neon is found to consist of  $^{20}\text{Ne}$ ,  $^{21}\text{Ne}$ ,  $^{22}\text{Ne}$  in the percentages of 90.92%, 0.26% and 8.82% respectively. Calculate the fractional atomic mass of neon.

Ans. Given data

Mass of  $^{20}\text{Ne}$  = 20amu

Mass of  $^{21}\text{Ne}$  = 21amu

Mass of  $^{22}\text{Ne}$  = 22amu

Percentage of  $^{20}\text{Ne}$  = 90.92%

Percentage of  $^{21}\text{Ne}$  = 0.26%

Percentage of  $^{22}\text{Ne}$  = 8.82%

Required

Fractional atomic mass of Ne = ?

Solution

Average Atomic mass of Ne =

$$\frac{(\text{Mass of } ^{20}\text{Ne} \times \% \text{ of } ^{20}\text{Ne}) + (\text{Mass of } ^{21}\text{Ne} \times \% \text{ of } ^{21}\text{Ne}) + (\text{Mass of } ^{22}\text{Ne} \times \% \text{ of } ^{22}\text{Ne})}{100}$$

$$= \frac{(20 \times 90.92) + (21 \times 0.26) + (22 \times 8.82)}{100}$$

$$= \frac{1818.4 + 5.46 + 194.04}{100} = \frac{2017.9}{100}$$

$$= 20.179 \text{ amu}$$

Fractional Atomic Mass of Ne = 20.179 amu

#### Example (2)

8.657g of a compound were decomposed into its elements and gave 5.217g of carbon, 0.962g of hydrogen, 2.478g of oxygen. Calculate the percentage composition of the compound under study.

Ans. Given data

Mass of compound = 8.657g

Mass of carbon = 5.217g

Mass of hydrogen = 0.962g

Mass of oxygen = 2.478g

Required

Percentage of carbon = ?

Percentage of hydrogen = ?

Percentage of oxygen = ?

Solution

Formula

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

$$\text{Percentage of carbon} = \frac{5.217}{8.657} \times 100 = 60.26\%$$

$$\text{Percentage of hydrogen} = \frac{0.962}{8.657} \times 100 = 11.11\%$$

$$\text{Percentage of oxygen} = \frac{2.478}{8.657} \times 100 = 28.62\%$$

The above results tell us that in one hundred gram of the given compound, there are 60.62g of C, 11.11g of H and 28.62g of O.

#### Example (3)

Ascorbic acid (vitamin C) contains 40.92% carbon, 4.58% hydrogen and 54.5% of oxygen by mass. What is the empirical formula of the ascorbic acid?

Ans. Given data

Percentage of carbon = 40.92%

Percentage of hydrogen = 4.58%

Percentage of oxygen = 54.5%

Required

Empirical formula of Ascorbic acid = ?

Solution:

Number of gram atoms

$$\text{Number of gram atoms of element} = \frac{\text{Percentage of element}}{\text{Atomic mass of element}}$$

$$\text{Number of gram atoms of C} = \frac{40.92}{12.0} = 3.41 \text{ gram atoms}$$

$$\text{Number of gram atoms of H} = \frac{4.58}{1.008} = 4.54 \text{ gram atoms}$$

$$\text{Number of gram atoms of O} = \frac{54.5}{16} = 3.406 \text{ gram atoms}$$

Atomic Ratio

$$\text{Atomic ratio of element} = \frac{\text{Number of gram atoms of element}}{\text{Smallest number}}$$

$$\text{Atomic ratio of C} = \frac{3.41}{3.406} = 1$$

$$\text{Atomic ratio of H} = \frac{4.54}{3.406} = 1.33$$

$$\text{Atomic ratio of O} = \frac{3.406}{3.406} = 1$$

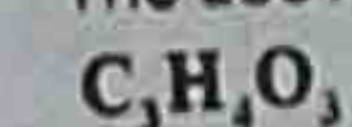
Conversion of atomic ratio into whole numbers

To convert the atomic ratio into whole number, multiply with three

$$\text{C : H : O} = 3(1 : 1.33 : 1) = 3 : 4 : 3$$

Empirical Formula

The above whole number ratio gives us the subscript for empirical formula of the Ascorbic acid





**Example (4)**

A sample of liquid consisting of carbon, hydrogen and oxygen was subjected to combustion analysis. 0.5439g of the compound gave 1.039g of  $\text{CO}_2$ , 0.6369g of  $\text{H}_2\text{O}$ . Determine the empirical formula of the compound.

**Ans. Given data**

Mass of organic compound = 0.5439g

Mass of carbon dioxide = 1.039g

Mass of water = 0.6369g

**Required**

Empirical formula of the compound = ?

**Solution**

**Percentage composition of the sample**

$$\begin{aligned}\text{Percentage of carbon} &= \frac{\text{Mass of } \text{CO}_2}{\text{Mass of organic compound}} \times \frac{12}{44} \times 100 \\ &= \frac{1.039\text{g}}{0.5439\text{g}} \times \frac{12.00}{44.00} \times 100 = 52.08\%\end{aligned}$$

$$\begin{aligned}\text{Percentage of hydrogen} &= \frac{\text{Mass of } \text{H}_2\text{O}}{\text{Mass of organic compound}} \times \frac{2.016}{18} \times 100 \\ &= \frac{0.6369\text{g}}{0.5439\text{g}} \times \frac{2.016}{18} \times 100 = 13.11\%\end{aligned}$$

$$\begin{aligned}\text{Percentage of oxygen} &= 100 - (\% \text{ of C} + \% \text{ of H}) \\ &= 100 - (52.08 + 13.11) = 34.77\%\end{aligned}$$

**Number of gram atoms**

$$\text{Number of gram atoms of element} = \frac{\text{Percentage of element}}{\text{Atomic mass of element}}$$

$$\text{Number of gram atoms of C} = \frac{52.08}{12} = 4.34 \text{ gram atoms}$$

$$\text{Number of gram atoms of H} = \frac{13.11}{1.008} = 13.00 \text{ gram atoms}$$

$$\text{Number of gram atoms of O} = \frac{34.77}{16.00} = 2.17 \text{ gram atoms}$$

**Atomic ratio**

$$\text{Atomic ratio of element} = \frac{\text{number of gram atoms of element}}{\text{Smallest number}}$$

$$\text{Atomic ratio of C} = \frac{4.34}{2.17} = 2$$

$$\text{Atomic ratio of H} = \frac{12.13}{2.17} = 6$$

$$\text{Atomic ratio of O} = \frac{2.17}{2.17} = 1$$

$$\text{Empirical formula} = \text{C}_2\text{H}_6\text{O}$$

**Example (5)**

The combustion analysis of an organic compound shows it to contain 65.44% carbon, 5.50% hydrogen and 29.06% oxygen. What is the Empirical formula of the compound? If the molecular mass of this compound is  $110.15 \text{ g mol}^{-1}$ . Calculate the molecular formula of the compound.

**Ans. Given data**

Percentage of carbon = 65.44%

Percentage of hydrogen = 5.50%

Percentage of oxygen = 29.06%

Molecular Mass =  $110.15 \text{ g/mol}$

**Required**

Empirical formula = ?

Molecular formula = ?

**Solution**

**Number of Gram atoms**

$$\text{Number of gram atoms of element} = \frac{\text{Percentage of element}}{\text{Atomic mass of element}}$$

$$\text{Number of gram atoms of C} = \frac{65.44}{12} = 5.45 \text{ gram atoms}$$

$$\text{Number of gram atoms of H} = \frac{5.50}{1.008} = 5.45 \text{ gram atoms}$$

$$\text{Number of gram atoms of O} = \frac{29.06}{16.00} = 1.82 \text{ gram atoms}$$

**Atomic ratio**

$$\text{Atomic ratio of element} = \frac{\text{number of gram atoms of element}}{\text{Smallest number}}$$

$$\text{Atomic ratio of C} = \frac{5.45}{1.82} = 3$$

$$\text{Atomic ratio of H} = \frac{5.45}{1.82} = 3$$

$$\text{Atomic ratio of O} = \frac{1.82}{1.82} = 1$$

**Empirical formula of the compound**

C, H and O are present in the compound in the ratio of 3: 3: 1 so the empirical formula  $\text{C}_3\text{H}_3\text{O}$ .

**Molecular formula**

In order to determine the molecular formula, first calculate the empirical formula mass.

Empirical formula Mass of  $\text{C}_3\text{H}_3\text{O} = 12 \times 3 + 1.008 \times 3 + 16 \times 1 = 55.05 \text{ g mol}^{-1}$

Molecular Mass =  $110.15 \text{ g mol}^{-1}$

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

$$= \frac{110.15}{55.05} = 2$$

Molecular formula =  $n \times (\text{Empirical formula})$

$$= 2 \times \text{C}_3\text{H}_3\text{O}$$

$$= \text{C}_6\text{H}_6\text{O}_2$$

**Example (6)**

Calculate the gram atoms (moles) in

(a) 0.1g of Sodium

(b) 0.1kg of Silicon

**Ans. Given data**

Mass of Sodium = 0.1g

Mass of Silicon =  $0.1 \text{ kg} = 0.1 \times 1000 = 100 \text{ g}$



Required

- (a) Number of gram atoms (moles) of sodium = ?  
 (b) Number of gram atoms (moles) of silicon = ?

Solution

(a) Formula

$$\text{Number of gram atoms} = \frac{\text{Mass of element in grams}}{\text{Atomic mass of element}}$$

$$\text{Atomic mass of Na} = 23 \text{ g mol}^{-1}$$

$$\text{Number of gram atoms of Na} = \frac{0.1 \text{ g}}{23 \text{ g mol}^{-1}} = 0.0043 \text{ moles}$$

$$(b) \text{ Atomic mass of Si} = 28.086 \text{ g mol}^{-1}$$

$$\text{Number of gram atoms of Si} = \frac{100 \text{ g}}{28.086 \text{ g mol}^{-1}} = 3.56 \text{ moles}$$

Example (7)

Calculate the mass of  $10^{-3}$  moles of  $\text{MgSO}_4$ .

Ans. Given data

$$\text{Moles of } \text{MgSO}_4 = 10^{-3}$$

Required

$$\text{Mass of } \text{MgSO}_4 = ?$$

Solution

$\text{MgSO}_4$  is an ionic compound. We will consider its formula mass in place of molecular mass.

Number of gram formula or

$$\text{Mole of the substance} = \frac{\text{Mass of ionic substance}}{\text{Formula mass of ionic substance}}$$

$$\text{Formula mass of } \text{MgSO}_4 = 24 + 32 + 16 \times 4$$

$$= 24 + 32 + 64 = 120 \text{ g mol}^{-1}$$

Applying the formula

$$10^{-3} \text{ mol} = \frac{\text{Mass of } \text{MgSO}_4}{120 \text{ g mol}^{-1}}$$

$$12 \times 10^{-3} = \text{Mass of } \text{MgSO}_4$$

$$0.12 \text{ g} = \text{Mass of } \text{MgSO}_4$$

Example (8)

How many molecules of water are there in 10.0g of ice? Also calculate the number of atoms of hydrogen and oxygen separately, the total number of atoms and the covalent bonds present in the sample.

Ans. Given data

$$\text{Mass of ice} = 10.0 \text{ g}$$

Required

$$\text{Number of molecules of water} = ?$$

$$\text{Number of atoms of hydrogen} = ?$$

$$\text{Number of atoms of oxygen} = ?$$

$$\text{Total number of atoms} = ?$$

$$\text{Total number of covalent bonds} = ?$$

Solution

Number of molecules of water

Formula

$$\text{Number of molecules of } \text{H}_2\text{O} = \frac{\text{Mass of } \text{H}_2\text{O}}{\text{Molar mass of } \text{H}_2\text{O}} \times N_A$$

## Scholar's CHEMISTRY - XI (Subjective)

$$\text{Molar Mass of } \text{H}_2\text{O} = 2 + 16 = 18 \text{ g mol}^{-1}$$

$$= \frac{10.0}{18} \times 6.02 \times 10^{23}$$

$$= 0.55 \times 6.02 \times 10^{23}$$

$$= 3.31 \times 10^{23} \text{ molecules}$$

$$\text{Number of molecules of water} = 3.31 \times 10^{23} \text{ molecules}$$

Number hydrogen atoms

$$\text{One molecule of water contains hydrogen atoms} = 2$$

$$3.31 \times 10^{23} \text{ molecules of water contain hydrogen atoms} = 2 \times 3.31 \times 10^{23}$$

$$= 6.62 \times 10^{23}$$

Number of oxygen atoms

$$\text{One molecule of water contains oxygen atom} = 1$$

$$3.31 \times 10^{23} \text{ molecules of water contain oxygen atom} = 1 \times 3.31 \times 10^{23}$$

$$= 3.31 \times 10^{23}$$

Total number of atoms of hydrogen and oxygen

$$= 6.62 \times 10^{23} + 3.31 \times 10^{23}$$

$$= 9.93 \times 10^{23}$$

Total Number of Covalent Bonds

$$\text{One molecule of water contains covalent bonds} = 2$$

$$3.31 \times 10^{23} \text{ molecules of water contain covalent bonds} = 2 \times 3.31 \times 10^{23}$$

$$= 6.62 \times 10^{23}$$

Example (9)

10.0g of  $\text{H}_3\text{PO}_4$  has been dissolved in excess of water to dissociate into ions.

Calculate,

$$(a) \text{ Number of molecules in 10.0g of } \text{H}_3\text{PO}_4.$$

$$(b) \text{ Number of positive and negative ions in case of complete dissociation in water.}$$

$$(c) \text{ Masses of individual ions.}$$

$$(d) \text{ Number of positive and negative charges dispersed in the solution.}$$

Ans. Given data

$$\text{Mass of } \text{H}_3\text{PO}_4 = 10.0 \text{ g}$$

$$\text{Molar Mass of } \text{H}_3\text{PO}_4 = 3 + 31 + 64 = 98 \text{ g mol}^{-1}$$

Required

$$(a) \text{ Number of molecules in 10.0g of } \text{H}_3\text{PO}_4 = ?$$

$$(b) \text{ Number of positive and negative ions in case of complete dissociation in water} = ?$$

$$(c) \text{ Masses of individual ions} =$$

$$(d) \text{ Number of positive and negative charges dispersed in the solution} = ?$$

Solution

(a) Number of molecules

Formula

$$\text{Number of Molecule} = \frac{\text{Mass of } \text{H}_3\text{PO}_4}{\text{Molar mass of } \text{H}_3\text{PO}_4} \times N_A$$

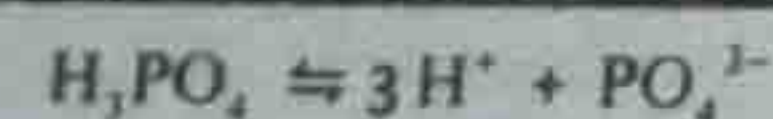


$$= \frac{10.0\text{g}}{98.0\text{g mol}^{-1}} \times 6.02 \times 10^{23}$$

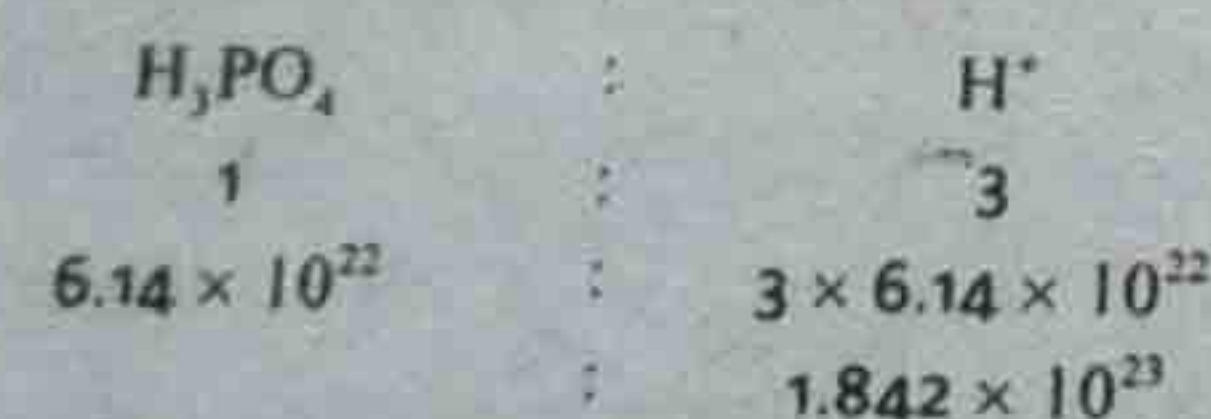
$$= 0.102 \times 6.02 \times 10^{23}$$

$$= 0.614 \times 10^{23} \text{ molecules}$$

(b)  $\text{H}_3\text{PO}_4$  dissolves in water and ionised as follows.



According to the balance chemical equation



Hence, the number of  $\text{H}^+$  is  $1.842 \times 10^{23}$



Hence, the number of  $\text{PO}_4^{3-}$  ions is  $6.14 \times 10^{22}$

(c) Masses of individual ions

Mass of positive ions

$$\text{Number of } \text{H}^+ = \frac{\text{Mass of } \text{H}^+}{\text{ionic mass of } \text{H}^+} \times N_A$$

$$1.842 \times 10^{23} = \frac{\text{Mass of } \text{H}^+}{1.008} \times 6.02 \times 10^{23}$$

$$\text{Mass of } \text{H}^+ = \frac{1.842 \times 10^{23} \times 1.008}{6.02 \times 10^{23}} = 0.308\text{g}$$

$$\text{Mass of } \text{H}^+ = 0.308\text{g}$$

$$\text{Mass of } \text{PO}_4^{3-} \text{ ions}$$

$$\text{Number of } \text{PO}_4^{3-} = \frac{\text{Mass of } \text{PO}_4^{3-}}{\text{ionic mass of } \text{PO}_4^{3-}} \times N_A$$

$$\text{Ionic Mass of } \text{PO}_4^{3-} = 31 + 64 = 95 \text{ g mol}^{-1}$$

$$6.14 \times 10^{22} = \frac{\text{Mass of } \text{PO}_4^{3-}}{95} \times 6.02 \times 10^{23}$$

$$\text{Mass of } \text{PO}_4^{3-} = 9.689\text{g}$$

(d) Number of positive and negative charges

One molecules of  $\text{H}_3\text{PO}_4$  positive charges = 3

$$6.14 \times 10^{22} \text{ molecules of } \text{H}_3\text{PO}_4 \text{ gives positive charges} = 3 \times 6.14 \times 10^{22}$$

$$= 1.842 \times 10^{23}$$

$$= 1.842 \times 10^{23} \text{ positive charges}$$

Number of positive and negative charges are always equal.

So number of negative charges =  $1.842 \times 10^{23}$  negative charges

Example (10)

A well known ideal gas is enclosed in a container having volume  $500 \text{ cm}^3$  at S.T.P. Its mass comes out to be  $0.72\text{g}$ . What is the molar mass of this gas?

Ans. Given data

Volume of the ideal gas =  $500 \text{ cm}^3$

Mass of the ideal gas =  $0.72\text{g}$

Required

Molar mass of the gas = ?

Solution

We can calculate the number of moles of the ideal gas at S.T.P from the given volume.

$22.414 \text{ dm}^3$  or  $22414 \text{ cm}^3$  of the ideal gas at S.T.P = 1mole

$$1 \text{ cm}^3 \text{ of the ideal gas at S.T.P} = \frac{1}{22414}$$

$$500 \text{ cm}^3 \text{ of the ideal gas at S.T.P} = \frac{1}{22414} \times 500$$

$$= 0.0223 \text{ moles}$$

We know that

$$\text{Number of moles of gas} = \frac{\text{Mass of the gas}}{\text{Molar mass of the gas}}$$

$$0.0223 \text{ mole} = \frac{0.72\text{g}}{\text{Molar mass of the gas}}$$

$$\text{Molar mass of the gas} = \frac{0.72\text{g}}{0.0223 \text{ mole}}$$

$$\text{Molar mass of the gas} = 32 \text{ g mol}^{-1}$$

Example (11)

Calculate the number of grams of  $\text{K}_2\text{SO}_4$  and water produced when  $14\text{g}$  of  $\text{KOH}$  are reacted with excess of  $\text{H}_2\text{SO}_4$ . Also calculate the number of molecules of water produced.

Ans. Given data

Mass of  $\text{KOH}$  =  $14\text{g}$

Molar mass of  $\text{KOH}$  =  $39 + 16 + 1 = 56 \text{ g mol}^{-1}$

Required

Number of grams (Mass) of  $\text{K}_2\text{SO}_4$  = ?

Mass of  $\text{H}_2\text{O}$  = ?

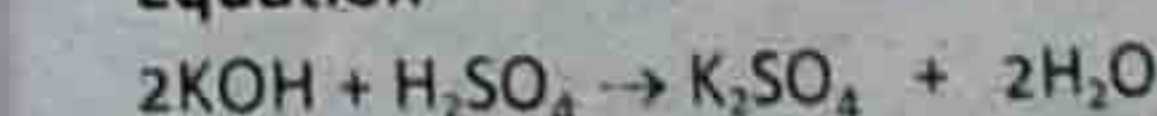
Number of molecules of  $\text{H}_2\text{O}$  = ?

Solution

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

$$= \frac{14.0\text{g}}{56\text{g/mole}} = 0.25 \text{ moles}$$

Equation



Comparison between Number of moles of  $\text{KOH}$  and  $\text{K}_2\text{SO}_4$

$\text{KOH}$	$\text{K}_2\text{SO}_4$
2	1
1	$\frac{1}{2}$



$$\frac{0.25}{0.25} = \frac{\frac{1}{2} \times 0.25}{0.125}$$

Number of moles of  $K_2SO_4 = 0.125 \text{ mole}$

$$\text{Number of moles of } K_2SO_4 = \frac{\text{Mass of } K_2SO_4}{\text{Molar mass of } K_2SO_4}$$

$$\begin{aligned} \text{Molar Mass of } K_2SO_4 &= 39 \times 2 + 32 + 16 \times 4 \\ &= 78 + 32 + 64 = 174 \text{ g mol}^{-1} \end{aligned}$$

$$0.125 \text{ moles} = \frac{\text{Mass of } K_2SO_4}{174 \text{ g mol}^{-1}}$$

$$0.125 \times 174 = \text{Mass of } K_2SO_4$$

$$21.75 \text{ g} = \text{Mass of } K_2SO_4$$

Comparison between number of moles of KOH and  $H_2O$

$$\begin{array}{ccc} \text{KOH} & : & H_2O \\ 2 & : & 2 \\ 1 & : & 1 \\ 0.25 & : & 0.25 \end{array}$$

Number of moles of  $H_2O = 0.25 \text{ moles}$

$$\text{Number of moles of } H_2O = \frac{\text{Mass of } H_2O}{\text{Molar mass of } H_2O}$$

$$\text{Molar mass of } H_2O = 2 + 16 = 18 \text{ g mol}^{-1}$$

$$0.25 \text{ moles} = \frac{\text{Mass of } H_2O}{18 \text{ g mol}^{-1}}$$

$$0.25 \times 18 = \text{Mass of } H_2O$$

$$4.5 \text{ g} = \text{Mass of } H_2O$$

Number of molecules of  $H_2O$

$$\text{Number of molecules of } H_2O = \text{Number of moles of } H_2O \times N_A$$

$$= 0.25 \times 6.02 \times 10^{23}$$

$$= 1.505 \times 10^{23} \text{ molecules}$$

$$\text{Number of molecules of } H_2O = 1.55 \times 10^{23} \text{ molecules}$$

#### Example (12)

Mg metal reacts with HCl to give  $H_2$  gas. What is the minimum volume of HCl solution (27% by Mass) required to produce 12.1g of  $H_2$ ? The density of HCl solution is  $1.14 \text{ g cm}^{-3}$ .



Ans.

Given data

Mass of  $H_2$  produced = 12.1g

Density of HCl solution =  $1.14 \text{ g cm}^{-3}$

Percentage of HCl solution = 27%

Required

Volume of HCl solution = ?

Solution

First of all convert the Mass of  $H_2$  into moles, then compare the moles of  $H_2$  and moles of HCl according to balance chemical equation.

Formula

$$\text{Number of moles of } H_2 = \frac{\text{Mass of } H_2}{\text{Molar mass of } H_2}$$

$$\text{Molar Mass of } H_2 = 2.016 \text{ g mol}^{-1}$$

$$\text{Number of moles of } H_2 = \frac{12.1 \text{ g}}{2.016 \text{ g mol}^{-1}} = 6 \text{ mole}$$

Equation



Comparison between Number of Moles of  $H_2$  and HCl

$$\begin{array}{ccc} H_2 & : & HCl \\ 1 & : & 2 \\ 6 & : & 2 \times 6 \\ 6 & : & 12 \end{array}$$

Moles of HCl = 12 moles

$$\text{Molar Mass of HCl} = 1 + 35.5 = 36.5 \text{ g mol}^{-1}$$

$$\text{Mass of HCl} = \text{Number of Moles of HCl} \times \text{Molar Mass of HCl}$$

$$= 12 \text{ moles} \times 36.5 \text{ g mol}^{-1} = 438 \text{ g}$$

Mass of HCl = 438 grams

We know that the percentage of HCl = 27% by weight, it means that

27g of HCl are present in HCl solution = 100g

$$1 \text{ g of HCl is present in HCl solution} = \frac{100}{27} \text{ g}$$

$$438 \text{ g of HCl are present in HCl solution} = \frac{100}{27} \times 438$$

$$= 1622.2 \text{ g}$$

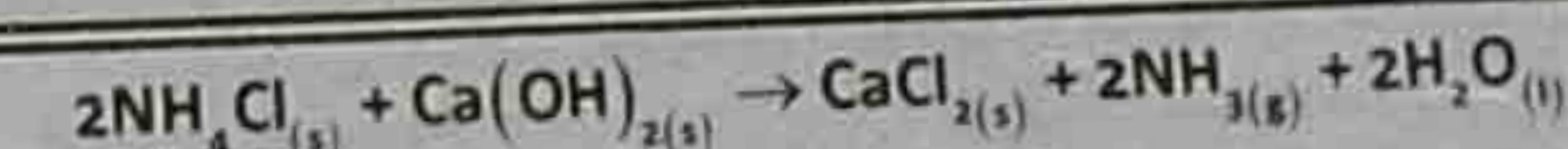
$$\text{Volume of HCl} = \frac{\text{Mass of HCl solution}}{\text{Density of HCl}}$$

$$\begin{aligned} \text{Volume of HCl} &= \frac{1622.2 \text{ g}}{1.14 \text{ g cm}^{-3}} \\ &= 1422.98 \text{ cm}^3 \end{aligned}$$

#### Example (13)

$NH_3$  gas can be prepared by heating together two solids,  $NH_4Cl$  and  $Ca(OH)_2$ . If a mixture containing 100g of each solid is heated then

(a) Calculate the number of grams of  $NH_3$  produced (b) Calculate the excess amount of reagent left unreacted



Ans. Given data

Mass of  $NH_4Cl$  = 100g

Mass of  $Ca(OH)_2$  = 100g

Required

(a) Number of grams (mass) of  $NH_3$  = ?

(b) Excess amount of reagent left unreacted = ?

Solution:

Number of moles of reactants

Formula

$$\text{Number of moles} = \frac{\text{Mass in grams}}{\text{Molar mass}}$$



Molar mass of  $\text{NH}_4\text{Cl} = 14 + 4 + 35.5 = 53.5 \text{ g mol}^{-1}$

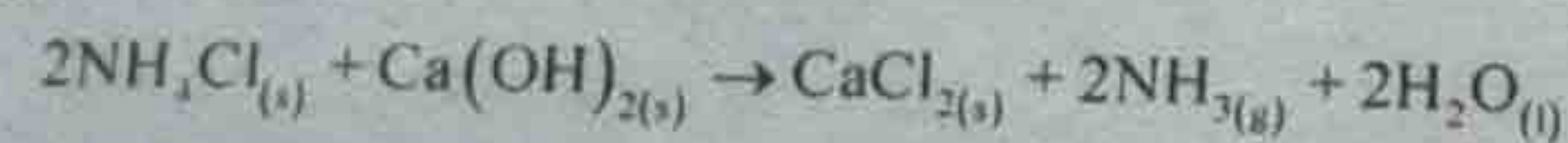
Number of moles of  $\text{NH}_4\text{Cl} = \frac{100\text{g}}{53.5 \text{ g mol}^{-1}} = 1.87 \text{ mole}$

Molar mass of  $\text{Ca}(\text{OH})_2 = 40 + 16 \times 2 + 2$   
 $= 40 + 32 + 2 = 74 \text{ g mol}^{-1}$

Number of moles of  $\text{Ca}(\text{OH})_2 = \frac{100\text{g}}{74 \text{ g mol}^{-1}} = 1.35 \text{ mole}$

Number of moles of product ( $\text{NH}_3$ )

Equation



Comparison between number of moles of  $\text{NH}_4\text{Cl}$  and  $\text{NH}_3$

$\text{NH}_4\text{Cl}$	:	$\text{NH}_3$
2	:	2
1	:	1
1.87	:	1.87

Number of moles of  $\text{NH}_3 = 1.87 \text{ moles}$

Comparison between number of moles of  $\text{Ca}(\text{OH})_2$  and  $\text{NH}_3$

$\text{Ca}(\text{OH})_2$	:	$\text{NH}_3$
1	:	2
1.35	:	$2 \times 1.35$
1.35	:	2.70

Number of moles of  $\text{NH}_3 = 2.70 \text{ moles}$

Since the number of moles of  $\text{NH}_3$  produced by 100g of  $\text{NH}_4\text{Cl}$  are less, so  $\text{NH}_4\text{Cl}$  is the limiting reactant. The other reactant,  $\text{Ca}(\text{OH})_2$  is present in excess.

Hence

$$\begin{aligned} \text{Mass of } \text{NH}_3 \text{ produced} &= 1.87 \text{ moles} \times 17 \text{ g mol}^{-1} \\ &= 31.79 \text{ g} \end{aligned}$$

(b) Amount of the reagent present in excess

To calculate the amount of the reagent present in excess compare the moles of  $\text{NH}_4\text{Cl}$  and  $\text{Ca}(\text{OH})_2$

$\text{NH}_4\text{Cl}$	:	$\text{Ca}(\text{OH})_2$
2	:	1
1	:	$\frac{1}{2}$
1.87	:	$\frac{1}{2} \times 1.87$
1.87	:	0.935

Hence the number of moles of  $\text{Ca}(\text{OH})_2$  which completely react with 1.87 moles of  $\text{NH}_4\text{Cl}$  is 0.935 moles.

Number of moles of  $\text{Ca}(\text{OH})_2$  taken = 1.35 moles

Number of moles of  $\text{Ca}(\text{OH})_2$  used = 0.935 moles

Number of moles of  $\text{Ca}(\text{OH})_2$  left behind =  $1.35 - 0.935$

$$= 0.415$$

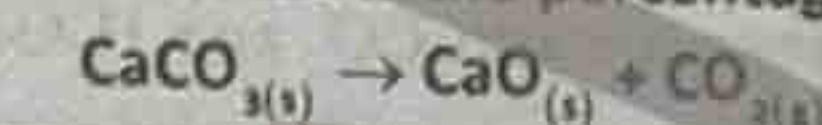
Mass of  $\text{Ca}(\text{OH})_2$  left unreacted (excess) =  $0.415 \times 74$

$$= 30.71 \text{ g}$$

Excess amount of  $\text{Ca}(\text{OH})_2$  left unreacted = 30.71 g

Example (14)

When lime stone  $\text{CaCO}_3$  is roasted, quicklime ( $\text{CaO}$ ) is produced according to following equation. The actual yield of  $\text{CaO}$  is 2.5kg, when 4.5kg of lime stone is roasted. What is the percentage yield of this reaction.



Ans. Given data

Mass of limestone roasted =  $4.5 \text{ kg} = 4.5 \times 1000 = 4500 \text{ g}$

Mass of quick lime (actual yield) =  $2.5 \text{ kg} = 2.5 \times 1000 = 2500 \text{ g}$

Molar Mass of  $\text{CaCO}_3 = 40 + 12 + 16 \times 3$   
 $= 40 + 12 + 48 = 100 \text{ g mol}^{-1}$

Molar mass of  $\text{CaO} = 40 + 16 = 56 \text{ g mol}^{-1}$

Required

Percentage yield of  $\text{CaO} = ?$

Solution

Equation



According to balanced chemical equation

100g of  $\text{CaCO}_3$  give  $\text{CaO} = 56 \text{ g}$

1g of  $\text{CaCO}_3$  gives  $\text{CaO} = \frac{56}{100} \text{ g}$

4500g of  $\text{CaCO}_3$  give  $\text{CaO} = \frac{56}{100} \times 4500 = 2520 \text{ g}$

Theoretical yield of  $\text{CaO} = 2520 \text{ g}$

Actual yield of  $\text{CaO} = 2500 \text{ g}$

Percentage yield =  $\frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$

$$= \frac{2500}{2520} \times 100$$

$$= 99.2\%$$

% yield of  $\text{CaO} = 99.2\%$

### Additional Questions

Q. In industry, costly reactant is always taken as limiting reactant.

Ans. In industry, costly reactant is taken in small amounts and cheaper one in excess. As a result of that costly reactant is completely consumed earlier. Hence, its amount is not wasted. Due to which reaction becomes economical and there is no financial loss of costly reactant.

Q. Differentiate between qualitative and quantitative analysis.

Ans.

Qualitative Analysis	Quantitative Analysis
The analysis which is made to identify different elements present in the compound is called as qualitative analysis.	The analysis which is made to determine the exact amount of each element present in the compound is called as quantitative analysis.
It is the initial analysis to detect the elements present in the compound.	By this analysis we can calculate the %age of element present in the compound.
e.g. Salt analysis, detection of elements and functional groups.	e.g. Volumetric analysis, combustion analysis, gravimetric analysis.



## Disclaimer

This Blog/Web Site is made available by the lawyer or law firm publisher for educational purpose only as well as to give you general information and a general understanding. We have the Rights to use this document for education purpose. You are not allowed to use this for commercial purpose. It is only for personal use. If you thoughts that this document include something related to you, you can email us at [yAsadBhatti@gmail.com](mailto:yAsadBhatti@gmail.com). We will look up into the matter and if we found anything related to you, we will remove the content from our website.

For Notes, Past Papers, Video Lectures, Education News

Visit Us at:

<https://www.bhattiAcademy.com>

<https://www.youtube.com/bhattiAcademy>

<https://www.facebook.com/bhattiAcademy>

If above links are NOT WORKING contact us at

[yAsadBhatti@gmail.com](mailto:yAsadBhatti@gmail.com)



**Q.** In combustion analysis, why the %age of oxygen cannot be measured directly?

**Ans.** In combustion analysis, a known amount of organic compound is burnt in free supply of oxygen. The carbon and hydrogen of the organic compound is converted into  $\text{CO}_2$  and  $\text{H}_2\text{O}$  respectively. But as oxygen gas is also provided from the external source to burn the organic compound, so we cannot measure the %age of oxygen present in the compound directly. However it is determined by method of difference.

$$\% \text{ age of O} = 100 - (\% \text{ age of C} + \% \text{ age of H})$$

**Q.** How can we calculate the efficiency of a chemical reaction?

**Ans.** The efficiency of a chemical reaction is calculated in the term of % age yield. "Percentage yield is defined as: it is the ratio of the actual yield to the theoretical yield multiplied by 100."

$$\% \text{ age yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

Greater the % age yield of a chemical reaction, more amount of product is formed. Hence greater the efficiency of the reaction.

**Q.** What is a Compound?

**Ans.** A substance formed by the reaction of two or more chemical elements. The elements in a compound are present in fixed ratios. For example pure water is a compound made from two elements - hydrogen and oxygen. The fixed ratio of hydrogen to oxygen in water is 2:1. Each molecule of water contains two hydrogen atoms bonded to a single oxygen atom. Compounds can be decomposed chemically into their constituent elements.

**Q.** The Term formula mass is used for ionic compounds instead of molecular mass.

**Ans.** The term formula Unit is used for the ionic compounds as represented in their empirical formula. And the sum of atomic masses of elements in the formula unit is called formula mass e.g Formula Unit of sodium chloride is NaCl and formula mass is 58.5g/ mole. The term molecular mass is used for molecular compounds e.g.  $\text{H}_2\text{O}$ .

### Important Previous Board Questions

- Q.** How does a limiting reactant control the amounts of products formed?
- Q.** How can the efficiency of a chemical reaction be expressed?
- Q.** Concept of limiting reactant is not applicable to the reversible reactions. Explain it.
- Q.** The atomic masses may be in fractions. Why?
- Q.** What is the function of electrometer in mass spectrometer?
- Q.** Which laws are to be considered during stoichiometric calculations?
- Q.** What is the justification of two strong peaks in mass spectrum of bromine?
- Q.** Amount of products formed during a chemical reaction; depend upon the amount of limiting reactant. Justify.
- Q.** Explain formation of ions with respect to energy changes.
- Q.** Give the applications of limiting reactant.
- Q.** Why the isotopes have same chemical properties?

For Answers study Scholar's CHEMISTRY (Objective) XI

## Chapter 2

# EXPERIMENTAL TECHNIQUES IN CHEMISTRY

### Analytical Chemistry

"The branch of chemistry which deals with the chemical characterization (qualitative and quantitative analysis) of a compound is called analytical chemistry."

### Major steps involved in quantitative analysis of a compound

Following major steps are necessary for complete quantitative analysis of a compound

- i. Obtaining a sample for analysis,
- ii. Separation of the desired constituent.
- iii. Measurement and calculation of results.
- iv. Drawing conclusion from the analysis.

### SEPARATION TECHNIQUES

#### FILTRATION

"The process in which insoluble particles (suspended particles or precipitates) are separated from liquids is called filtration."

#### Filter media

Filtration can be performed with several types of filter media. Nature of the precipitate and other factors dictate which filter medium must be used. Following filter media are frequently used for filtration.

- |                |              |                    |
|----------------|--------------|--------------------|
| • Filter paper | • Paper pulp | • Filter crucibles |
| • Cloth        | • Cotton     | • Sand             |

#### Filtration through filter paper

1. Filtration by a glass funnel and filter paper is usually a slow process. As the mixture is poured onto the filter paper, the solvent (water) passes through leaving behind the suspended particles on the filter paper.
2. Filter papers are available in variety of porosities (pore sizes). Which pore size is to be used, depends upon the size of the particles in the precipitate.
3. The filter paper should be large enough so that it is one-fourth to one-half full of precipitate at the end of filtration. The funnel should be large enough for its rim to extend 1 to 2 cm above the top circumference of the paper.
4. If the process of filtration is to run smoothly, the stem of the funnel should remain continuously full of liquid as long as there is liquid in the conical portion.
5. The stem of funnel should be several inches long so that it can extend a few centimeters down into the receiving beaker and tip should touch the side of the beaker. In this way, the filtrate runs down the side of beaker without splashing.