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CHEMISTRY

4

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PREFACE

Chemistry is an interesting and fundamental branch of science because it gives us the chance to explain the secrets of nature.

What is water? What do we use in our cars as fuel? What is aspirin? What are perfumes made of? These kinds of questions and their answers are all part of the world of chemistry. Chemists work everyday to produce new compounds to make our lives easier with the help of this basic knowledge. All industries depend upon chemical substances, including the petroleum, pharmaceuticals, garment, aircraft, steel, and electronics industries, etc.

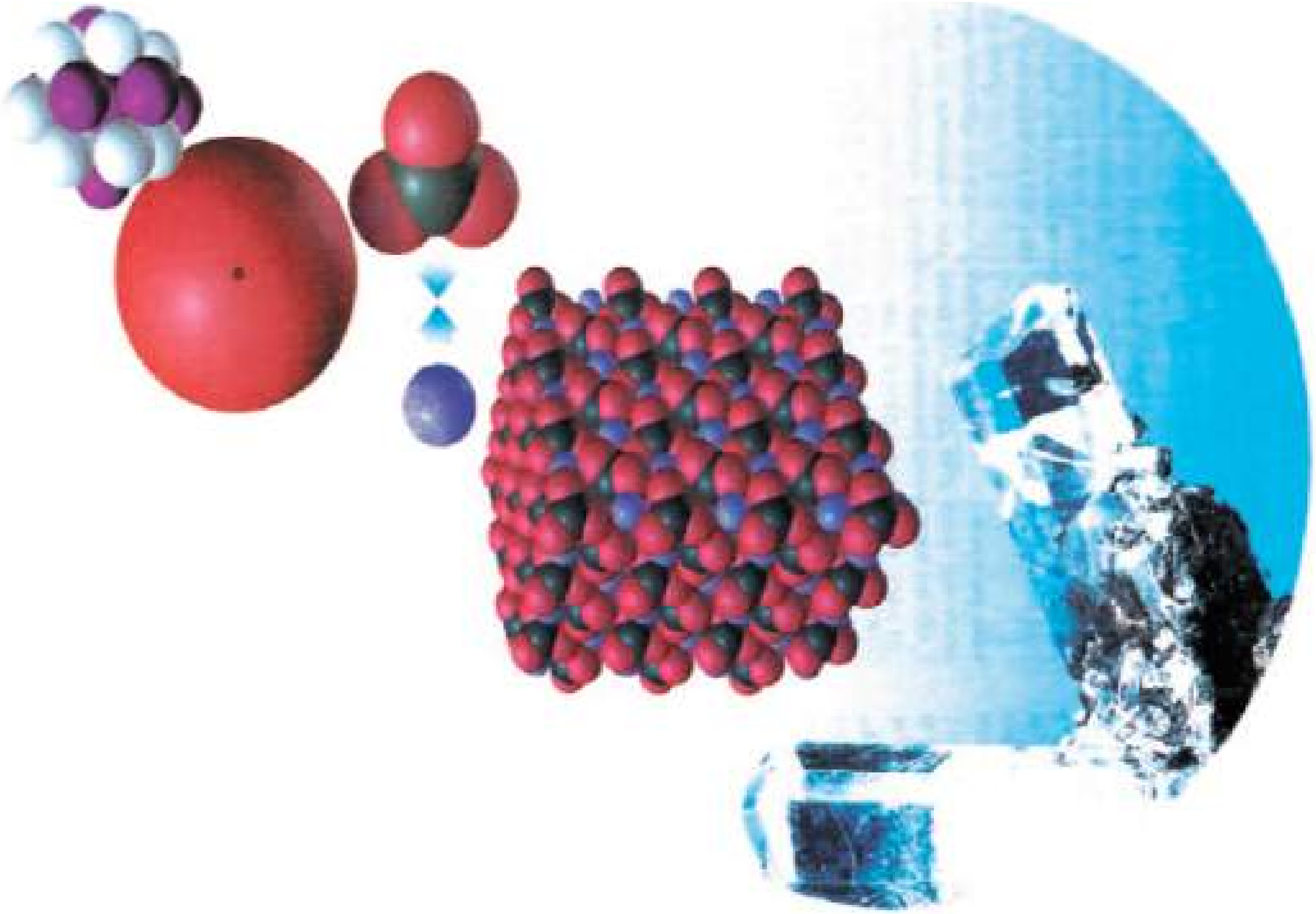
This book helps everyone to understand nature. However, one does not need to be a chemist or scientist to understand the simplicity within the complexity around us. The aim was to write a modern, up-to-date book where students and teachers can get concise information about chemical reactions and inorganic compounds. Sometimes reactions are given in detailed form, but, in general, excessive detail has been omitted. Throughout the book, different figures, colourful tables, important reactions are used to help explain ideas.

We hope that after studying this book, you will find chemistry in every part of your life.

The Author

BASIC CONCEPTS IN CHEMISTRY

CHAPTER-1



ACHIEVEMENTS

After studing this chapter, the student will be able to:

* Understand the atomic theory of Dalton and its hypotheses and use the laws of the Chemical Union to identify the formation of chemical compounds and the ratios of fixed elements in them.

*Understand the purpose of the Gay-Lussac's law for combining gas volume, and its relationship to the Avogadro's hypothesis.

*Define the basic terms: valance, atomic mass, equivalent mass and the relationship between them and define the concept of the mol, the molar mass, Avogadro's number and the relationship between them.

*Find out the chemical formula, Empirical and molecular formula for the compounds to find out them and how by percentages of elements in compounds.

John Dalton «1766 A.D.-1844 A.D.» Was born of Poor Weaver father. He has worked at an early age. However, he had spent much of his spare time studying Latin mathematics and the natural sciences. He took care of the meteorology study. From 1787 A.D. until his death, more than 200,000 notes were recorded, and his interest led him to investigate the properties of the gases, and he discovered the law of partial pressure, and also concluded that the dissolution of gas in a combination of gases proportional to the partial pressure and since 1803 A.D., he worked to develop his atomic theory, he added the double-descent law as well as the concept of relative atomic mass.

1-1-DALTON'S ATOMIC THEORY

Scientific and chemical researchs and discoveries, which took place at the end of the eighteenth century and the beginning of the 19th century, led to the knowledge that the substance consisted of atoms, and that the difference in type and number of the atom determines the properties and type of molecules that formate it, i.e. the specifications of the substance. This helped the English scientist John Dalton to declare the

atomic theory of the matter in 1803 A.D. and named the Dalton's atomic theory, which included the following hypotheses:

1-The substance consists of small, indivisible particles called "atoms" (which the scientists later managed to fragment

2-Atoms cannot be synthesized, within the human abilities.

3- Atoms of the same element are similar in all their physical and chemical properties and differ from the atoms of other elements.

4- Compound atoms (as called by Dalton) are formed by the union of atoms of the elements atoms with simple ratios.

Eight years later, some modifications were made, replacing the expression "compound atoms" with "molecules" by the Italian scientist Avogadro and our understanding of these hypotheses must be in the period Time that Dalton came in, in the first half of the 19th century, where scientific progress.

1-2-LAWS OF CHEMICAL UNION

As a result of the development of scientific knowledge and the new existing discoveries which based on practical experiments and scientific concualsions led to the interpretation of the composition of the material and the drafting of the laws of the Chemical Union in the second half of the eighteenth century and the beginning of the 19th century .

The first of these laws was the law of mass conservation, which answered the question of what happens to the substance during its chemical reaction? Can it be destroyed or created? is the mass of reactant substances differ from the mass of the material resulting from the reaction, or is it equal? To answer all those questions, the French scientist Lavoisier did the oxidation of tin in a closed receptacle, and he found that the mass of the closed receptacle remained unchanged, because a chemical reaction was made between tin and oxygen and new molecules were formed of tin oxide(II).

Since the mass of the atom does not depend on the nature of the other atoms with which it unites, it is axiomatic that all the oxygen and tin atoms involved in the chemical reaction maintain their mass unchanged, and the Lavoisier experiments indicated that:

"The mass of material cannot be destroyed or created during the chemical reaction" i.e., :-

The mass of reactants = the mass of products from the reaction.

The Arab scientist Abu Qasem Almajriti «950 A.D.-1007 A.D.» is considered first to prove the validity of this law, he noticed when heating a weighted amount of elemental mercury in a closed glass container with air presence will turn mercury into a fine red powder without a change inthe total mass of reactants within the vessel. He repeated Lavoisier and reached the same conclusion and then laid down a law of conservation of mass.

Example 1-1

73g of HCl gas was passed into a solution containing 158 g of sodium thiosulfate, as a result 117 g of table salt, 64 g of SO_2 gas, 32 g of sulfer and 18 g of water has been produced. Prove that these results support the law of conservation of mass.

Solution:

The sumation of the reactants' mass= 73+158=231 g

The summation of the products' mass=117+64+32+18=231 g

The summation of the reactants' mass=The summation of the products' mass which is support the law of conservation of mass

The other chemical union laws can be summed up with a single idea, the union of elements with constant weight ratios for compound formation. The scientist Prust was the first to develop the law of constant composition,

which stipulated that "all specimens of a particular compound have the same proportions of its constituent elements"

Let's take the following example:

If the water is dissociated we will find that 16 g of oxygen in the sample is present versus 2 g of hydrogen, or the ratio of oxygen mass to hydrogen:

$$\text{The ratio} = \frac{16 \text{ g (O)}}{2 \text{ g (H)}} = 8$$

This ratio will be found in each sample of pure water regardless of the source from which it was taken or in what way it was prepared (Fig. 1-1) The compounds have a fixed structure. The water in any way has a constant percentage of hydrogen to oxygen regardless of which source it came from.

Lavoisier 1743 A.D. -1794 A.D.
The law of Conservation of Mass

Ebu El Qasim Mesleme bin Ahmed El Majriti (950-1007).

He was born in Andalusia and died there. He was accepted as the pioneers of chemistry at his time.

One of his important is Ghayet El Hakim. He observed that when he put 1/4 pound of Mercury in closed container with air inside and heated it, a soft and red powder was formed. There occurred no change in mass. This substance is called mercury (II) oxide today.

The ratio of H and O is constant in water with regard to source of it

Exercise 1-1

Two carbon monoxide samples were analysed and obtained from two different sources. The first sample contained 4.3 g of oxygen and 3.2 g of carbon. The second sample contained 7.5 g oxygen and 5.6 g of carbon. Do these results achieve the law of constant compositions.

Example 1-2

Two carbon dioxide samples were obtained from two different sources. They were dissociated into components. The first sample contained 4.8 g of oxygen and 1.8 g of carbon, while the other sample contained 17.1 g of oxygen 6.4 g of carbon. Show that these results correspond to the law of constant compositions.

Solution:

The ratio of oxygen mass to carbon in the first sample

$$\text{Ratio} = \frac{4.8g(O)}{1.8g\ (C)} = 2.7$$

The ratio of oxygen mass to carbon in the second sample

$$\text{Ratio} = \frac{17.1\ g\ (O)}{6.4\ g\ (C)} = 2.7$$

Since the ratio is the same for two, it means that these results are consistent with the law of constant compositions.

1-3-GAY-LUSSAC'S LAW OF COMBINING GAS VOLUMES

The French scientist Gay-Lussac (Joseph Gay-Lussac) has worked very much in the reaction of gases and saw that there is a correlation between the volumes of gases that involved in the chemical reaction and resulting from it under the same conditions of pressure and temperature, and the measurement of the volumes of reactive gases resulting from the reaction formulated the results of his investigation in the year 1808 AC in the following announcement:

"The volumes of gases involved in or resulting from the chemical reaction correspond to each other by a simple numerical proportionality if measured under the same conditions of pressure and temperature."

For example;

1- One volume of hydrogen is combined with one volume of chlorine, and two volumes of hydrogen chloride are formed, the ratio between the two united gaseous volumes and the volume of gas produced is 1:1: 2, as in the following chemical equation:

$$H_2 + Cl_2 \rightarrow 2HCl$$

2- When the water is electrically analysed, the volume of the freed hydrogen is twice the volume of the oxygen, and it combines two volumes of hydrogen with one volume of oxygen and produces two volumes of water vapour.

$$2H_2 + O_2 \rightarrow 2H_2O$$

The ratio between the two volumes united gaseous sizes and the volume of water vapour produced is 2:1:2, which is a simple numerical ratio. The above can be represented in the following forms:

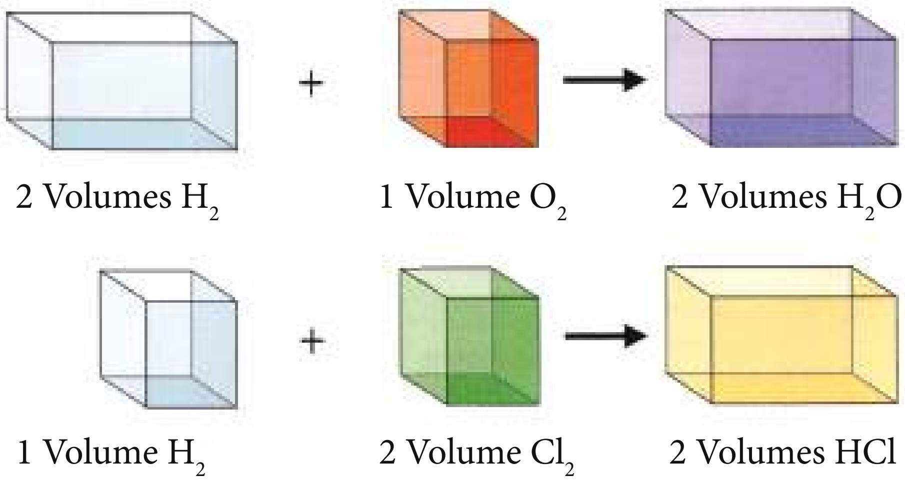


Figure 1-4
Reaction of gases with a constant ratio

1-4-AVOGADRO'S HYPOTHESIS

In 1811, Italian scientist Amedeo Avogadro found out that gas molecules can have more than one atom. In other words, gases can have diatomic molecules. Avogadro defined molecule as the smallest independent sample of matter. He also defined atom as the smallest part of an element which can be found in molecules of different compounds. The basic hypothesis of Avogadro is as follows:

(Equal volumes of gases contain equal number of molecules under the same pressure and temperature conditions.)

Avogadro's hypothesis explained the simple ratios between volumes of reactant and product gases and also provided significant results with the atomic numbers of gases with complex molecules. According to Avogadro, the number of atoms forming a molecule of an element is constant. This is also valid for compounds. Only the types of atoms are different which form compound molecules.

For example, if 1 volume of hydrogen gas reacts with 1 volume of chlorine gas, 2 volumes of HCl gas is formed.

$H_2 + Cl_2 \rightarrow 2HCl$

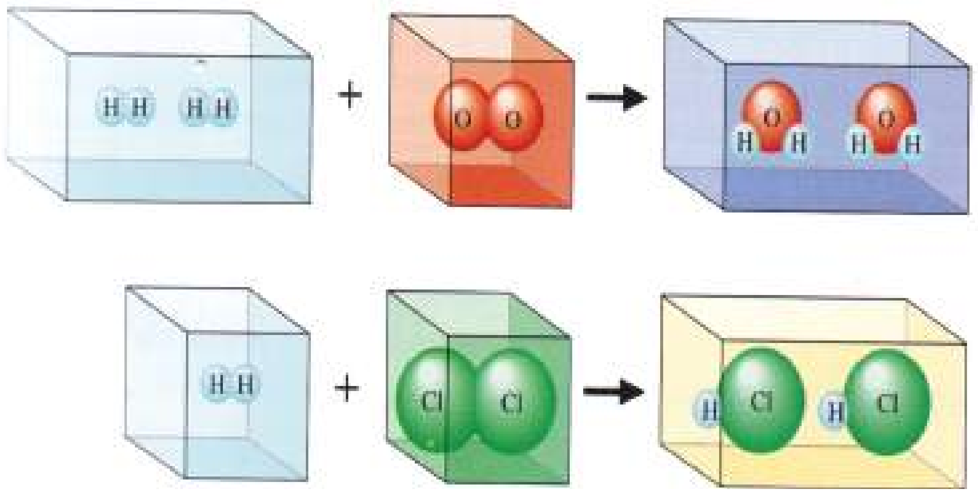
From the reaction of 2 volumes of hydrogen gas and 1 volume of oxygen gas, 2 volumes of H_2O are obtained.

$2H_2 + O_2 \rightarrow 2H_2O$

This isn't contradictory with Dalton's atom model. According to this, a hydrogen molecule consists of 2 atoms. Similarly, oxygen and chlorine molecules consist two atoms. HCl molecule consists of 1 chlorine and 1 hydrogen atom. Water molecule consists 2 hydrogen and 1 oxygen atom.

Figure 1-1

Balls represent atom in molecules formed from combination



1-5-VALANCE

Chemical formulas of compounds weren’t assigned arbitrarily but they were assigned according to bonding of atoms in molecules of compounds. An atom of element needs to have a capability in order to be bonded to atoms of other elements.

The capability of bonding of an atom or number of hydrogen atoms that an atom of an element can combine is called as valence. The bonding capabilities of elements are different. Valence concept was introduced to chemistry in 19th century. We can define valence of an element as follows:

“It is the number of electrons which an element loses, gains or shares during a chemical reaction.”

For example, the valence of hydrogen is 1. Because it has an electron in its outer shell which can be shared. Valence of oxygen in water is 2 because it has 6 electrons in its outer shell. Therefore it needs 2 electrons to satisfy its outer shell. Valence of Na is 1. It has an electron to lose in its outer shell. Valence of Mg is 2 because it can lose 2 electrons from its outer shell. Valence of Cl is 1 because it needs one electron to satisfy its outer shell.

Atomic mass unit (AMU or amu)

An atomic mass unit (symbolized AMU or amu) is defined as precisely 1/12 the mass of an atom of carbon-12. The carbon-12 (C-12) atom has six protons and six neutrons in its nucleus.

1-6-ATOMIC MASS

Atoms on the degree of precision and smallness so it was difficult to estimate their atomic masses, however it has been possible to assign their mass with great accuracy, for example, it has found that a hydrogen atom mass reach 1.64 × 10-24 g. it has also been possible to obtain relative masses of atoms by setting the mass of elements with another element, provided that the relative number of atoms in the compounds is known, so atomic mass used to express an element mass for another element atom mass and it is agreed to use in determining the relative masses of all the elements of the periodic table. In 1961 in Geneva held a Conference of the International Union of Pure and Applied Chemistry (IUPAC) which was agreed on the definition of the standard unit of atomic mass and Atomic Mass Unit named after (AMU) as equivalent to one of twelve part of carbon isotope atom mass 12 and that it's atomic mass considered equal to 12 units and therefore:-

Atomic mass unit (amu) = (mass of an atom of carbon isotope 12)/12

i.e.; 1 (amu)= (1/12) mass of an atom of carbon isotope 12 as the mass of carbon 12 isotope atom = $\dfrac{12}{\text{Avogadro's number}}$

$= \dfrac{12}{(6.023\times10^{23})}$

So

$1\ (\text{amu}) = \dfrac{1}{12} \times \dfrac{12}{6.023\times10^{23}}$

$1\ (\text{amu}) = 1.66\times10^{-24}$ g

Thus the atomic mass we use today and found in the periodic table are not actual masses, but relative masses show the relationship in terms of atomic mass between different atoms. For example, the atomic mass of

hydrogen isotope 1 is (1/12) the atomic mass of the isotope carbon-12 which is about 1 amu, either nucleus oxygen 16 have mass equal to(16/12) or (4/3) from the mass of carbon 12 isotope, so when atomic mass estimates in grams it called the gram-atomic mass.

The gram-atomic mass for oxygen = 16 g and for silver= 107.9 g and each mass of these masses has Avogadro's number of atoms which is equal to 6.023×10^{23} atoms, for example:

1 g of hydrogen contains 6.023×10^{23} atoms of hydrogen

39 g of potassium contain 6.023×10^{23} atoms of potassium

207 g of lead contain 6.023×10^{23} atoms of lead

The absolute atomic mass is the mass of one atom of the element. i.e.;

The absolute atomic mass of an element=gram-atomic mass of the element/ Avogadro's number

Example 1-3:

Calculate the absolute atomic mass of the oxygen given that it's atomic mass is equal to 16.

Solution:

The absolute atomic mass of an element=gram-atomic mass of the element/ Avogadro's number

$= \dfrac{16}{(6.023\times10^{23})}$

$=2.656\times10^{-23}$g

1-7-EQUIVALENT MASS

The study of the law of mass ratio in which the various elements combined led to find out the equivalent masses, where Dalton was the first one who calculate these masses, he assumed that the element mass that combined with one atom of hydrogen's mass is the mass equivalent to the element, because the hydrogen component failure in the formation of compounds with the most other elements, or to the fact that most elements do not combine directly with hydrogen, and oxygen combine with it directly. Oxygen has been adopted as a basis for the calculation of equivalent masses, and it's combined mass considered equal to «eight 8». This is the elements are not limited to the combination with each other by equivalent quantities only, but replace each other in their compounds with equivalent masses. Thus the definition of equivalent mass for an element is:-

"The mass of that element which combine with eight mass parts of oxygen or pull these quantities from their compounds"

The equivalence mass concept has enabled drafting of the following law, called the law of equivalent masses:-
" The elements combine together in quantities correspond with their equivalent masses."
When the equivalent mass estimated in grams it is called then the (Gram Equivalent), for example, the gram equivalent for oxygen = 8 g , chlorine = 35.5 g, hydrogen = 1 g and silver = 107.9 g, and so on. The equivalent masses can be estimated from the data on the analysis of different compounds, or replace an element, and it is not necessary to specify the equivalent masses to start from compounds containing oxygen or with other equivalent mass element information depending on the following:
(Mass of 1st element/it's equivalent mass)=(mass of 2nd element/it's equivalent mass)

Exercise 1-2

1.31 g of Cu was obtained from reduction of 1.64 g of copper oxide with hydrogen. Calculate equivalent mass of Cu. (equivalent mass of O = 8g)

Exercise 1-3

What is the equivalent mass of an element if its atomic mass is 55.85 and its valence is 3?

Example 1-4

3.5 g of iron combined with sulphur to form 5.5 g of iron sulphide (II). Calculate the equivalent mass of iron given that the equivalent mass of sulphur=16 g.

Solution:

Mass of S = 5.5 - 3.5 = 2 g

$$\frac{\text{mass of } 1^{st} \text{ element}}{\text{equivalent mass of } 1^{st} \text{ element}} = \frac{\text{mass of } 2^{nd} \text{ element}}{\text{equivalent mass of } 2^{nd} \text{ element}}$$

$$\frac{2}{16} = \frac{3.5}{\text{equivalent mass of Fe}}$$

Equivalent mass of Fe = 28 g

1-8-THE RELATIONSHIP BETWEEN ATOMIC MASS, EQUIVALENT MASS AND VALENCE

While measuring atomic masses, O atom was accepted as 16 units. While measuring equivalent masses, O atom was accepted as 8 units. According to this, there is a mathematical relationship between the two. In order to calculate equivalent mass of an element, mass of the element is divided to the number of atoms which the element can be bonded to or replace.

Here the common denominator is the valence of the element. Equivalent mass (the result) is obtained by dividing mass of element to its valence.

$$\text{equivalent mass of element} = \frac{\text{atomic mass of element}}{\text{valence of element}}$$

Example 1-5

What is the valence of Al if its atomic mass is 27 and equivalent mass is 9?

Solution:

$$\text{valence of element} = \frac{\text{atomic mass of Al}}{\text{equivalent mass Al}} = \frac{27}{9} = 3$$

1-9-DENSITY OF GASES

The density can be defined by the following relationship:

Density (kg/m^3) = [the mass (kg)/the volume (m^3)]

ρ (kg/m^3) = m (kg) / V (m^3)

The density unit can be(g/cm^3), or (g/mL) for solids and liquids as for gases, the mass of 1 ml be too small to deal in practice, so it had taken a litter «L» as unit volume for gas density measurement. The volumes of gases are greatly influenced by pressure and temperature, so the circumstances must be specified under which measure the density of gases, and the conditions that gas is measured in the temperature of zero

degrees Celsius (0 °C) and pressure (1 atm)are called standard conditions or (Standard Temperature and Pressure) (STP).

Example 1-6

If you know that a density of a gas is equal to 0.7 g/L and it occupies a volume of 490 cm3 at STP. What is the mass of this gas?

Solution:

Firstly, we need to convert cm^3 to L.

$$V\,(L) = 490\,\cancel{cm^3} \times \frac{1\,L}{1000\,\cancel{cm^3}} = 0.490\,L$$

We use the following relationship to calculate mass of gas.

$$m\,(g) = \rho\,(g/L) \times V\,(L)$$

$$m\,(g) = 0.7\,(g/\cancel{L}) \times 0.490\,(\cancel{L}) = 0.343\,g$$

1-10-MOLE CONCEPT

Chemical reactions occur between large number of particles, and these particles may be in the form of atoms, molecules or ions and each of these particles has their own relative mass, but there is no fit in between the amounts of each matter and its mass. For example, if a student was asked to compare 1 g of H_2 gas and 1 g of N_2 gas and 1 g of O_2 gas in terms of the number of particles it cannot solved for the simple reason that the molecular masses of these elements vary from each other, the

molecular mass of hydrogen gas 2, nitrogen 28 and for oxygen 32. If the mass of 1 g is divide for each item on the molecular mass.

$$0.031 = \frac{1}{32} : O_2 \qquad 0.036 = \frac{1}{28} : N_2 \qquad 0.5 = \frac{1}{2} : H_2$$

Then values obtained can be used for comparison. So, years ago, the need arose for independent units to express the amount of substance and this unit had found general acceptance, this unit is mole (mole) and symbolizes (n) and it is one of the basic units in the international system of units, the mole defined as the quantity of material containing the same number of particles (molecules, atoms or ions) containing in 12 g of carbon isotope 12 (12C) (where this isotope is used also as a measure to calculate atomic mass as previously mentioned) and this number of particles called the Avogadro's Number) which equals to

Exercise 1-4

The mass of a gas is 0.4 g and its volume is ¼ L. What is its density under standard conditions?

6.023×10^{23} and symbolizes N_A. It must be emphasized that Mole is the actual amount of the matter and is not the mass.
The Mole is one of the most important basic concepts in general chemistry and which adopted by scientists to consolidate their theory to many important issues in chemistry. And we can apply the concept of mole on atoms, molecules, ions or electrons so it is always necessary to specify the type of particle that we deal with, for example,
Mass of one mole of atoms of the isotope carbon-12 is 12 g
Mass of one mole of atoms of silver is 107.868 g
Mass of one mole of H_2 molecules are 2 g
Mass of one mole of ions SO_4^{2-} is 96 g
And number of moles (n) is calculated using the following relationship:

$$\mathbf{n\ (mol)} = \frac{\mathbf{mass\ (m)\ (g)}}{\mathbf{molar\ mass\ (M)\ (g/mol)}}$$

1-10-1-Molar Mass

Since the particles are a group of atoms chemically joined together the final mass for these molecules known from the masses of their atoms i.e., we use relative mass to compare various molecules in terms of mass. Meaning that:-

Molar mass of matter =Total relativistic masses of atoms which make up matter

Molar mass is the mass of 1 mole of a compound in gram unit. (Molar mass used to be called as molecular mass formerly.)

For example, if we want to calculate the molar mass of methane gas (CH_4) or if we ask what is the mass of 1 mole of methane gas, as each methane molecule has 1 C and 4 H atoms, 1 mole of methane molecule contains 1 mole of C atom and 4 moles of H atoms. According to this, we can calculate the mass of 1 mole of methane gas as follows:

Mass of 1 mole of C = $1 \times 12 = 12$ g/mol

Mass of 4 moles of H = $4 \times 1 = 4$ g/mol

Mass of 1 mole of CH_4 = 16 g/mol

In the same manner, we can calculate molar mass of H_2SO_4 acid.

Mass of 2 moles of H = $2 \times 1 = 2$ g/mol

Mass of 1 mole of S = $1 \times 32 = 32$ g/mol

Mass of 4 moles of O = $4 \times 16 = 64$ g/mol

Mass of 1 mole of H_2SO_4= 98 g/mol

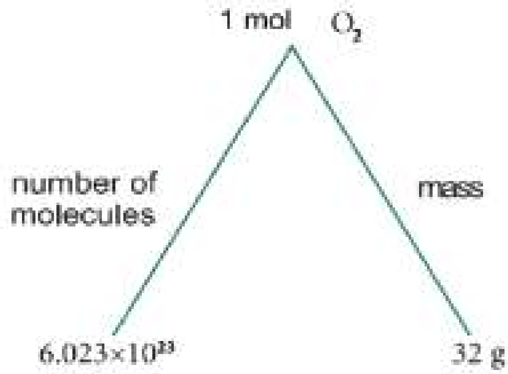
If 16g of methane gas makes up 1 mole of CH_4 and 98 g of sulfuric acid makes up 1 mole of H_2SO_4, logically we can call those masses as molar mass. As we understand from previous calculations we have made, in chemistry, all atoms and molecules are bound with the following :

a) Atomic mass and molar mass

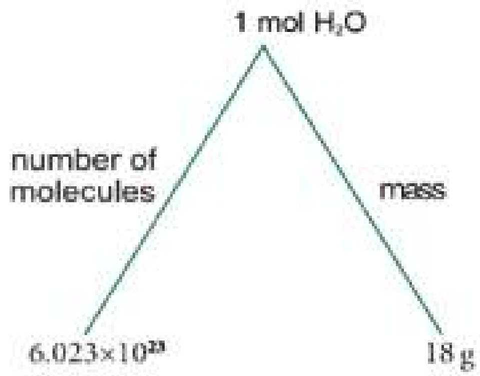
b) Mole

c) Avogadro’s number

Avogadro's number or Avogadro's constant is the number of atoms in 1 mole of an element or number of molecules in 1 mole of a molecule. It is shown with N_A. Previously, we have mentioned that atoms, molecules or ions as many as Avogadro's number are called as **a mole**.



Oxygen molecules



Water Molecules

Example 1-7

Find the molar mass for the following compounds:-

a) Aqueous sodium sulfate $Na_2SO_4 . 7H_2O$

b) Juglone $C_{10}H_6O_3$

c) Sulfur dioxide SO_2

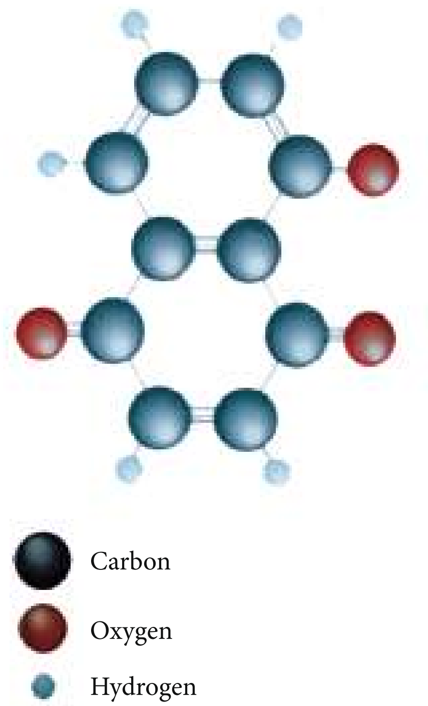
Solution:

a) $M(Na_2SO_4 . 7H_2O) = (2\times23) + (1\times32) + (4\times16) + 7(2\times1+1\times16) = 268$ g/mol

b) $M(C_{10}H_6O_3) = (10\times12) + (6\times1) + (3\times16) = 174$ g/mol

c) $M(SO_2) = (1\times32) + (2\times16) = 64$ g/mol

*You can refer to Table 3 at the end of the book to learn atomic masses.



Juglone is an organic compound. It is used as a pesticide. It is also a natural dye.

Example 1-8

Find out the mole numbers of the following compounds.

a) 9.6 g of SO_2
b) 85 g of NH_3

Solution:

a) Molar mass of SO_2:

$$M(SO_2)=(1\times32)+(2\times16)=64 \text{ g/mol}$$

$$n\,(\text{mol}) = \frac{m\,(g)}{M\,(g/mol)} = \frac{9.6\,\cancel{(g)}}{64\,\cancel{(g/mol)}} = 0.15 \text{ mol } SO_2$$

b) Molar mass of NH_3:

Exercise 1-5

Perform the following conversions.

a) 3.2 g Cu = ……….mol Cu
(Cu = 64 g/mol)

b) 0.2 mol Ag=……..g Ag
(Ag = 108g/mol)

c) 1 silver atom = …… g

d) 3.01×10^{22} Fe atoms =………..mol Fe =…………g Fe
(Fe = 56 g/mol)

e) 6.4 g S =…….mol S =…….S atoms (S = 32 g/mol)

Exercise 1-6

a) What is the mass of 0.04 mole of N_2?
b) What is the number of moles in 5.6 g of PCl_5?
c) Calculate the molar mass of the gas which has 22.54 g in 0.23 mole.

$$M(NH_3) = (1\times14)+(3\times1)=17 \text{ g/mol}$$

$$n(\text{mol}) = \frac{m(\text{g})}{M(\text{g/mol})} = \frac{85\ (\cancel{\text{g}})}{17\ (\cancel{\text{g}}/\text{mol})} = 5 \text{ mol } NH_3$$

Example 1-9

Calculate the mass of 0.7 mol of manganese dioxide (MnO_2).

Solution:

$$M(MnO_2) = (1\times55)+(2\times16) = 87 \text{ g/mol}$$

$$m(\text{g}) = n(\text{mol})\times M(\text{g/mol})$$

$$m(\text{g}) = 0.7\ (\cancel{\text{mol}})\times 87\ (\text{g}/\cancel{\text{mol}}) = 60.9 \text{ g } MnO_2$$

1-14-2-Applications of Mole Concept

As we have told before, the mass of 1 mole of carbon atom is 12 grams. According to this, it is possible to calculate the mass of 1 carbon atom.

$$= \frac{12 \text{ g}}{\text{Avogadro's number}} = \frac{12 \text{ g}}{6.023\times10^{23}} = 1.995\times10^{-23} \text{ g/atom}$$

We can write the following equation:

$$\text{Number of moles} = \frac{\text{Number of particles (atoms, ions or molecules)}}{\text{Avogadro's number}}$$

Example 1-10

a) Calculate the number of moles of 3.01×10^{25} water molecules.
b) Calculate the number of molecules in 0.02 mole of CO_2.

Solution:

a) $$\text{Number of moles} = \frac{\text{number of molecules}}{\text{Avogadro's number } (N_A)}$$

$$= \frac{3.01\times10^{25}}{6.023\times10^{23}} = 50 \text{ moles } H_2O$$

b) Number of molecules = number of moles x Avogadro's number

$$= 0.02\times6.023\times10^{23} = 1.2\times10^{22}\ CO_2 \text{ molecules}$$

Example 1-11

Calculate the number of molecules in 170 g of H_2S gas.

Solution:

$M(H_2S) = (2\times1) + (1\times32) = 34$ g/mol

$$n\ (\text{mol}) = \frac{m\ (\text{g})}{M\ (\text{g/mol})} = \frac{170\ (\text{g})}{34\ (\text{g/mol})} = 5\ \text{mol}$$

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

$$= 5 \times 6.023 \times 10^{23} = 3.01 \times 10^{24}\ H_2S \text{ molecules}$$

1-11-MASS PERCENTAGES OF ELEMENTS IN A COMPOUND

There are two ways to describe molecular structures for compounds, first; knowing how many atoms of each element involved in the composition of the compound and second; know percentages in terms of mass of elements in this structure, i.e., element in g 100 grams of the compound, the percentage can be found for each element in the composition of compound as follows:

A- Find the molar mass of a compound of molecular form.

B- Set and find the mass of each element in a molecule, i.e., product of atomic mass for each element x number of its atoms

C- Find percentage for the item in the compound by the following relationship:

$$\text{Mass percentage of an element in a compound} = \frac{\text{Molar mass of element} \times \text{number of atoms}}{\text{Molar mass of the compound}} \times 100$$

Example 1-12

Calculate the percentage for each carbon, hydrogen and oxygen in the compound of isopentyl acetate ($C_7H_{14}O_2$) (substance released by the bee insect).

Solution:

Molar mass of $C_7H_{14}O_2$:

$M(C_7H_{14}O_2) = (7\times12) + (14\times1) + (2\times16) = 130$ g/mol

According to the equation above, we can find out mass percentages of elements.

$$\%C = \frac{7\times12}{130} = \frac{84}{130}\times100\% = 64.61\%$$

$$\%H = \frac{14\times1}{130} = \frac{14}{130}\times100\% = 10.77\%$$

$$\%O = \frac{2\times16}{130} = \frac{32}{130}\times100\% = 24.62\%$$

Sum of mass percentages of elements is equal to 100%.

Exercise 1-7

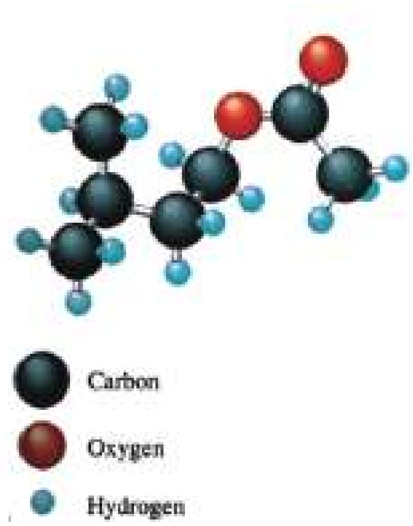
Calculate the number of molecules of SiO_2 found in 1mg of dust. (Assume that dust particles are made up of 100% SiO_2.)

Exercise 1-8

Calculate the number of moles of each of the following.

a) 3.01×10^{22} N_2 molecules

b) 4.82×10^{24} iron atoms



Isopentyl is a substance made by bees.

Exercise 1-9

Find out mass percentages of elements in CH_3COOH acid.

Exercise: 1-10

50% of by mass of a 2.2 moles mixture of He and Ar gasses is He. Calculate the mole number of He in the mixture? He = 4, Ar = 40

Example 1-13

a) Find out mass percentages of elements in $C_2H_2O_4$ acid.

b) What is the mass percentage of crystal water in $C_2H_2O_4.2H_2O$ compound?

Solution:

a) Molar mass of $C_2H_2O_4$

$M(H_2C_2O_4) = (2\times1) + (2\times12) + (4\times16) = 90$ g/mol

$$C\% = \frac{2\times12}{90} = \frac{24}{90} \times100\% = 26.67\%$$

$$H\% = \frac{2\times1}{90} = \frac{2}{90} \times 100\% = 2.22\%$$

$$O\% = \frac{4\times16}{90} = \frac{64}{90} \times100\% = 71.11\%$$

b) For $C_2H_2O_4 . 2H_2O$:

$M(C_2H_2O_4 . 2H_2O) = 2\times12 + 2\times1 + 4\times16 + 2(2\times1 + 1\times16) = 126$ g/mol

For crystal water:

$$H_2O\% = \frac{2\times18}{126} \times 100\% = 28.57\%$$

In order to find out the mass of any element in a compound with a known mass, molar mass of compound and atomic mass of elements are used. This calculation can be expressed with the following formula.

$$\text{Mass of element} = \frac{\text{Atomic mass of element in the compound} \times \text{number of element in the compound} \times \text{mass of compound}}{\text{Molar mass of compound}}$$

Example 1-14

Find out the mass of Ca in 20 g of $Ca_3(PO_4)_2$.

Solution:

Molar mass of $Ca_3(PO_4)_2$:

$M(Ca_3(PO_4)_2) = (3\times40) + 2(31 + 4\times16) = 310$ g/mol

If we insert the values in the formula above;

$$\text{Mass of Ca} = \left(\frac{3\times40}{310}\right) \times 20 = 7.74 \text{ g}$$

Example 1-15

In 10 g of $CuSO_4.5H_2O$,

a) Find out mass of Cu.

b) Find out mass percentage of crystal water.

Solution:

a) Molar mass of $CuSO_4.5H_2O$:

$$M(CuSO_4.5H_2O) = 1 \times 64 + 1 \times 32 + 4 \times 16 + 5(2 \times 1 + 1 \times 16)$$
$$= 250 \text{ g/mol}$$

$$\text{Mass of Cu} = \left(\frac{64}{250}\right) \times 10 = 2.56 \text{ g}$$

b) Using the same equation;

$$\text{Mass of crystal water} = \left(\frac{18 \times 5}{250}\right) \times 10 = 3.6 \text{ g}$$

Exercise 1-11

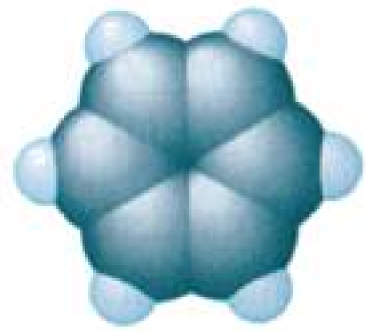
Calculate the mass of Na in 25 g of $Na_2CO_3.10H_2O$.

1-12-CHEMICAL FORMULA

The chemical composition of compounds represented with formulas, which are its constituent elements symbols set with the number of atoms of those elements in the molecule and the composition of known chemical matter can be expressed in different formats including:

1-12-1 Empirical Formula

It is the simplest formula that gives the minimum limit of absolute information about the compound, it sets the rational number of the atoms included in the compound. For example, one molecule of benzene consist of 6 carbon atoms and 6 hydrogen atoms, so the greatest common divisor for these numbers is 6, and by dividing the number of atoms over 6 we will get the empirical formula for benzene which is CH.



$C_6H_6 = (CH)_6$

Benzene molecule

The same for the water molecule which is consist of two atoms of hydrogen and one atom of oxygen, so the empirical formula will be H_2O, and so for one molecule of glucose which consist of 6 carbon atoms, 12 hydrogen atoms and 6 oxygen atoms, so the greatest common division for these numbers is 6, and by dividing he atom numbers by 6 the empirical formula will be CH_2O.

$C_6H_{12}O_6 = (CH_2O)_6$

Glucose

How to find the empirical formula for compounds?

We should follow the steps below to find an empirical formula of a compound:

a) The elements forming a compound are analyzed.

b) The ratio of number of atoms is found by dividing mass of element or its percent ratio with atomic mass.

$$\text{Ratio of no. atoms of element} = \frac{\text{Mass of element or percent ratio}}{\text{Atomic mass}}$$

c) In order to get the simplest ratio of number of atoms, we need to divide number of atoms of element to the number of atoms with lowest ratio and round it up to nearest integer.

$$\text{Simplest ratio of number of atoms} = \frac{\text{Number of atoms of each element}}{\text{The lowest value of number of atoms}}$$

We get the empirical formula this way.

Example 1-16

It is found that one of the gases consist of 20% of hydrogen and 80% of carbon. Find the empirical formula for this gas.

Solution:

1) Divide each percent of element on its atomic mass, i.e.,

$$\text{Number of atoms of element} = \frac{\text{Percent ratio of element}}{\text{Atomic mass}}$$

$$\text{Number of atoms of H} = \frac{20}{1} = 20$$

$$\text{Number of atoms of C} = \frac{80}{12} = 6.60$$

2) Dividing the above ratios over the smallest one and round to the nearest integer

$$\text{Simplest ratio of number of atoms} = \frac{\text{Number of atoms of each element}}{\text{The lowest value of number of atoms}}$$

$$\text{The simplest number of atoms of H} = \frac{20}{6.60} = 3$$

$$\text{The simplest number of atoms of C} = \frac{6.60}{6.60} = 1$$

The empirical formula of gas : CH_3

Example 1-17

Cholesterol is an organic compound and it is found in almost every tissue of human body. Therefore it causes occlusion in veins. It contains 83.87% C, 11.99% H and 4.14% O. Find out the empirical formula of cholesterol.

Solution:

1) Divide each percent of element over its atomic mass, i.e.

$$\text{Number of atoms of element} = \frac{\text{Percent ratio of element}}{\text{Atomic mass}}$$

,

$$\text{Number of atoms of H} = \frac{11.99}{1} = 11.99$$

Number of atoms of C = $\frac{83.87}{12}$ = 6.989

Number of atoms of O = $\frac{4.14}{16}$ = 0.258

2) We divide the results by the smallest value and get it close to the nearest ….. number.

$$\text{The simplest number of atoms} = \frac{\text{Number of atoms of each element}}{\text{The lowest value of number of atoms}}$$

The simplest number of atoms of H= $\frac{11.99}{0.258}$ = 46

The simplest number of atoms of C= $\frac{6.989}{0.258}$ = 27

The simplest number of atoms of O = $\frac{0.258}{0.258}$ =1

The empirical formula of cholesterol: $C_{27}H_{46}O$

1-12-2-Molecular Formula

It is the chemical formula that gives the real number of the atoms of elements that involve in the composition of one molecule of the matter. For example the molecule of ethane consist of 2 carbon atoms and 6 hydrogen atoms so the molecular formula will be (C_2H_6), according to that, the molecular formula is twice the empirical formula (CH_3). The same for the molecular formula of the water which is (H_2O) that means the water composes by the bonding of two hydrogen atoms with one oxygen atom, and it is the same of its empirical formula of water (H_2O) so:

Molecular formula = empirical formula x empirical formula unit

To find the molecular formula of a substance, we follow these steps:

a) Find out the empirical formula as mentioned before.

b) Calculate the molar mass for the empirical formula by summation of atomic masses of their elements.

c) Find the molar mass of the substance (molecular formula).

d) Divide the molar mass of the molecular formula over the molar mass of the empirical formula to obtain the units of the empirical formula.

$$\text{Empirical formula unit} = \frac{\text{Molar mass of molecular formula}}{\text{Molar mass of empirical formula}}$$

e) The result of division is multiplied with empirical formula and molecular formula is determined.

Exercise 1-12

After burning a medicine, it was found to contain 74.27% carbon, 7.47% hydrogen, 12.99% nitrogen and 4.95% oxygen. Write down the empirical formula of this medicine.

Exercise 1-13

What is the empirical formula of a compound consisting of 7.8 g of K, 7.1 g of Cl and 9.6 g of O? K = 39 Cl = 35.5 O = 16

Example 1-18

The molar mass of an organic acid = 60 g/mol and contains 0% carbon, 6.7% hydrogen and the rest is oxygen. Find the molecular formula of the organic acid.

Solution:

The percentage of oxygen is 100-(40+6.7) = 53.3%

1) For each element, percent ratio is divided to atomic mass:

$$\text{Number of atoms of element} = \frac{\text{Percent ratio of element}}{\text{Atomic mass}}$$

Number of atoms of H= $\frac{6.7}{1} = 6.7$

Number of atoms of C= $\frac{40}{12} = 3.3$

Number of atoms of O = $\frac{53.3}{16} = 3.3$

2) These values are divided to the smallest among them and rounded up to the nearest integer.

$$\text{Simplest ratio of number of atoms} = \frac{\text{Number of atoms of each element}}{\text{The lowest value of number of atoms}}$$

The simplest no. of atoms of H = $\frac{6.7}{3.3} = 2$

The simplest no. of atoms of C = $\frac{3.3}{3.3} = 1$

The simplest no. of atoms of O = $\frac{3.3}{3.3} = 1$

For the empirical formula CH_2O

Thus, molar mass of CH_2O:

M (CH_2O) = (1 × 12) + (2 × 1) + (1 × 16) = 30 g/mol

$$\text{Empirical formula unit} = \frac{\text{Molar mass of molecular formula}}{\text{Molar mass of empirical formula}}$$

$$= \frac{60}{30} = 2$$

$(CH_2O)x\ 2$

Molecular formula = $C_2H_4O_2$

Example 1-19

The empirical formula of an organic compound is known as C_2H_4O.As molar mass of this compound is 88g/mol, find out its molecular formula.

Solution:

C_2H_4O

$M(C_2H_4O) = (2 \times 12) + (4 \times 1) + (1 \times 16) = 44$ g/mol

$$\text{Empirical formula unit} = \frac{\text{Molar mass of molecular formula}}{\text{Molar mass of empirical formula}} = \frac{88}{44} = 2$$

Molecular formula: $C_4H_8O_2$

Exercise 1-14

Caffeine is found in tea, coffee and chocolate. It contains 49.48% C, 5.15% H, 16.49% O and 28.88% N. As its molar mass is 194 g/mol, find out its molecular formula.

Caffeine

BASIC CONCEPTS

Law of Conservation of Mass: Matter is neither created nor destroyed

Law of constant composition: The composition of elements are constant in all samples of a compound.

Law of Combining Gas Volumes: There is a simple ratio between the volumes of reactant and product gases in a chemical reaction under the same temperature and pressure.

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

Valence: It is the number of electrons which an element loses, gains or shares during a chemical reaction.

Atomic Mass Unit:The mass of C-12 isotope is accepted as 12.0000 and 1/12 of this is called as 1 atomic mass unit (amu).

Temperature:Degree of temperature is used to tell to express that some objects have high or low temperatues; temperature tells which way heat flows. Heat flow occure from a cold object to a hot object spontaneously. Temperature has got three units. Centigrade (C˚), Kelvin (K) and Fahrenheit (F˚)

Empirical Formula: It is the simplest formula of chemical compounds. It shows relative numbers of elements which form a compound.

Molecular Formula: It is the formula which gives detaied and clear information about elements which form a chemical compound.

Avogadro's Number: It is either number of atoms in 1 mole of an element or number of molecules in 1 mole of a compound. Its value is 6.023×1023.

Mlar mass: It is the unit of mass of 1 mole of atom or 1 mole of molecule in gram unit. Its unit is g/ mol.

Mole (n): The amount of substance which contains, molecules or ions as many as Avogadr's number is called as 1 mole of substance.

QUESTIONS OF CHAPTER -1

1.1) What is Dalton' atomic model? What is the relationship of it to the law of mass conservation?

1.2) It was observed that when H_2 and Cl_2 reacted, there was a simple ratio between the resulting gas (HCl) and the elements forming it. How do you explain the results according to the law of constant composition?

1.3)When two samples of CS_2 samples were decomposed to its components, the first sample was found to have 8.68g S and 1.51g C whereas the second sample was found to have 31.3g S and 3.85g C. Verify if these results are consistent with the law of constant composition.

1.4) When two samples of sodium chloride were decomposed to their constituents , the first sample was found to have 4.65g Na and 7.16g Cl whereas the second sample was found to have 7.45g Na and 11.5g Cl. Verify if these results are consistent with the law of constant composition.

1.5) Na/F ratio in sodium fluoride is 1.21. When a sample of NaF was decomposed, it was found to have 34.5g of Na. Calculate the amount of fluoride that the sample contained in grams.

1.6) When a sample of magnesium fluoride was decomposed, it was found to have 1.65kg of magnesium and 2.57kg of fluoride. Another sample contained 1.32kg of magnesium. Calculate the amount of fluoride inside it.

1.7) When two samples of carbon tetrachloride (CCl_4) were decomposed, the first sample was found to have 32.4g of C and 373g of Cl whereas the second sample contained 12.3 g of C and 112 g of Cl. Verify if these results are consistent with the law of constant composition.

1.8) Define the following terms
Valence, atomic mass unit (amu), equivalent mass, atomic mass, Avogadro's theory

1.9) 2 L of gas was obtained from the reaction of 1L of Cl_2 and 3L of F_2 gases. As the volumes of these gases were measured at constant pressure and temperature, what is the formula of the product gas?

1.10) When 1.55 g of silver was heated with chlorine gas, 2.05 g of silver chloride (AgCl) was formed. As the equivalent mass of chlorine is 35.5 g, calculate the equivalent mass of silver.

1.11) When 0.72g of zinc was added to lead acetate, lead precipitated. After this precipitate was washed and dried, its mass was measured as 2.29g. What is the equivalent mass of lead? (Equivalent mass of zinc = 32.5g)

1.12) If the valence of an element is 2 and its equivalent mass is 32.7g, find out the atomic mass of this element.

1.13) If the atomic mass of an element is 55.85g and its valence is 3, find out its equivalent mass.

1.14) Calculate the number of moles of the following.

a) 7 g of $NaHCO_3$
b) 10 mg of Fe
c) 16g of CO_2

1.15) a) Calculate the number of moles of 5 g of silver and its number of atoms.

b) A piece of diamond contains 5.0×10^{21} C atoms. Calculate the no. of moles of C and its mass in grams.

1.16) Calculate the amount of the following.

a) The mass of 3.8×10^{20} NO_2 molecules.

b) The number of atoms of Cl in 0.0425g of $C_2H_4Cl_2$.

1.17) Calculate the molar mass percentages of the elements forming the following compounds.

a- $NaClO_3$ b- $CuSO_4.5H_2O$ c- $(NH_4)_3PO_4$ d- $Al_2(SO_4)_3$ e- $Ca(C_2H_3O_2)_2$

1.18) Calculate the molar mass of magnesium and crystal water in $MgSO_4.7H_2O$ compound.

1.19) A sample of urea contains 1.121 g of N, 0.161 g of H, 0.4808 g of C and 0.640 g of O. Find out the empirical formula of this compound.

1.20) A compound contains C, H and N elements. When 35mg of this compound was burnt, 33.5mg of CO_2 and 41.1mg of H_2O were obtained. Find out the empirical formula of this compound.

1.21) The mass of some white dust was known to be 31.9% K, 39.2% O and 28.9% Cl. Find out the empirical formula and molecular formula of this compound. (Molar mass of the compound is 122.5g.)

1.22) A compound is made up of 24.27% C, 4.07% H and 71.65% Cl and its molar mass is known to be 99g/mol. What is the molecular formula of this compound?

1.23) 52.2% of a compound is C and 13.1% H and the remaining is O. Find out the molecular formula of this compound. (Molar mass of the compound is 46 g/mol)

1.24) Make the following calculations.

a) The number of moles of O in 7.2 moles of H_2SO_4 ,

b) The number of atoms of Zn with a mass of 48.3 g,

c) The mass of 6.73 moles of Al in grams,

d) The mass of Fe in 79.2 g of Fe_2O_3.

GASES **CHAPTER-2**

ACHIEVEMENTS

After studying this chapter, students will be able to.

*Learn gas state of matter and its properties.

*Learn the factors affecting gas state of matter.

*Learn gas laws.

*Can explain diffusion of gases.

*Can differentiate ideal gases and real gases.

*Can explain the effect of pressure on liquid vapor and boiling point.

2-1-PREFACE

We live in the layer of the atmosphere in which gases are found most densely is called as troposphere. It consists of 78% N_2, 21% O_2 and 1% other gases. Most of this 1% is CO_2.In nature,some substances are found in gas form at 25 °C temperature and 1 atm pressure conditions. In the following table 2-1, those gases and their symbols are given.

Table 2-1 Some Elements and Compound in gas form at room temperature

Chemical Symbol	Element	Chemical Formula	Compound
H_2	Hydrogen	HF	Hydrogen flouride
N_2	Nitrogen	HCl	Hydrogen Chloride
O_2	Oxygen	HBr	Hydrogen Bromide
F_2	Fluorine	HI	Hydrogen Iodide
Cl_2	Clorine	CO	Caron monoxide
Ne	Neon	CO_2	Carbon dioxide
Ar	Argon	NH_3	Ammonia
Kr	Krypton	NO	Nitrogen Monoxide
Xe	Xenon	NO_2	Nitrogen Dioxide
Rn	radon	N_2O	Dinitrogen monoxide
		SO_2	Sulfur dioxide
		H_2S	Hydrogen Sulfide

Molecules of matter can be studied most easily when they are in gas form. Gas molecules occupy 0.1% of the medium in which they are found; the remaining is space. Therefore, gas molecules move independently. As gases have compressibility, their volumes can be reduced, they can be liquefied by pressure or cooling.

2-2-VOLUME

Volume is the amount of space matter takes up. The volume of gases is as much as the volume of the container they are inside. Volume is shown with V and measured with l, ml or cm^3.

To convert units of volume;

1 L = 1000 cm^3

1 L = 1000 mL

1 cm^3 = 1 mL conversion are used

Exercise 2-1

What is the volume of a 0.125 L sample of O_2 gas in millilitres(ml)?

Do you know that?

The highest temperature was recorded in Mexico as +85°C whereas the lowest temperature was recorded in Antarctica as -88°C.

Exercise 2-2

a) -100°C b)1°C c)127°C

Write down the temperatures above in Kelvin scale.

Do you know that?

According to football rules, mass of a football must be between 410 and 450 grams and pressure inside must be between 0.6 and 1.1 atm. If pressure is higher than those, there is a risk of blast.

Exercise 2-3

What is the equivalent of 1.5 atm pressure in Torr?

Example 2-1

What is the volume of an 800 cm^3 sample of NO_2 gas in liters?

Solution:

$$V(L) = V\,cm^3 \times \frac{1L}{1000\ cm^3} = 800\ \cancel{cm^3} \times \frac{1L}{1000\ \cancel{cm^3}} = 0.8\ L\ NO_2$$

2-3-TEMPERATURE

In Chapter 1, when we have mentioned temperature, we have told that there are different temperature scales, the Celsius (°C) and Kelvin (K) following formula to convert Celsius (°C) scale to Kelvin (K).

$$T(K) = t(°C) + 273$$

Example 2-2

The temperature of water in a container is 80°C whereas the temperature of water in another container is -13°C. What are their temperatures in Kelvin scale?

Solution:

First container:

$$T(K) = t(°C) + 273$$

$$T(K) = 80 + 273 = 353\ K$$

Second container:

$$T(K) = t(°C) + 273$$

$$T(K) = (-13) + 273 = 260\ K$$

2-4-PRESSURE

Pressure is the force (F) applied to per unit area (A). symbolizes it(P). While pressure in open air is measured with a barometer, pressure of a closed container is measured with a manometer.

$$= \frac{F\ (Force)}{A\ (area)} = Pa\ (Pascal)$$

$$P = \frac{1N(Newton)}{m^2}$$

Units of pressure are Pascal (Pa), atmosphere (atm) and Torr. The relationship between them is shown below.

1 atm = 101325 Pa
1 atm = 760 mmHg
1 atm = 760 Torr
1 Torr = 1 mmHg
1 atm = 76 cmHg = 760 mmHg = 760 Torr

Example 2-3

What is the equivalent of 688 Torr pressure in atm?

Solution:

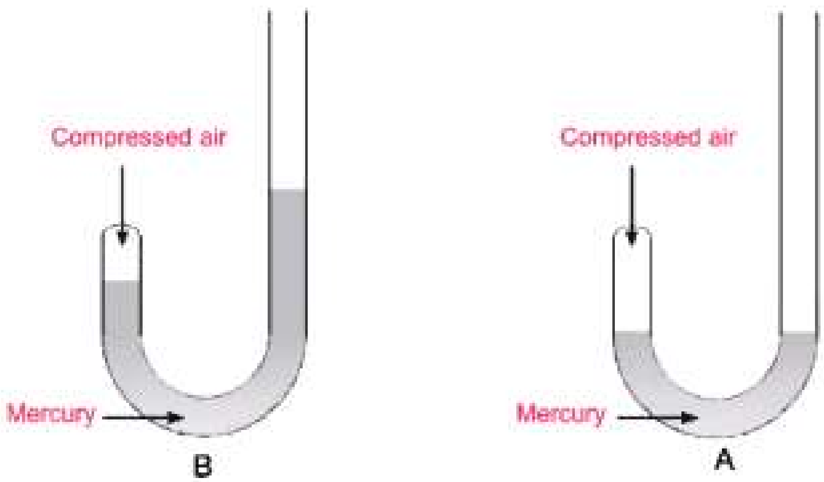
$$P\,(atm) = P\,Torr \times \frac{1\,atm}{760\,Torr} = 688\,\cancel{Torr} \times \frac{1\,atm}{760\,\cancel{Torr}}$$

$$P\,atm = 0.905\ \ atm$$

2-5-THE GAS LAWS

2-5-1-Volume-Pressure Relationship (Boyle's Law)

British scientist Robert Boyle made the first experiment on pressure and volume relationship. for gases He used a U-shaped glass tube with one arm shorter. The short arm is closed and filled with some gases and when mercury is added from the open and long arm, it starts to apply pressure on the gas. Until the gas is compressed to some volume, the amount of mercury increases and as shown in Fig. 2-1, the volume of gas has decreased.



Boyle observed that as pressure increased, volume of air decreased at constant temperature and the gas amount. Thus, Boyle's Law is stated "as long as temperature and amount of gas are kept constant, volume of gas is inversely proportional to pressure applied to it." This relationship is shown mathematically as follows:

$$V \propto \frac{1}{P} \quad \text{and} \quad V = k\,\frac{1}{P}$$

That means the product of pressure and volume is equal to a constant. According to this law, at constant temperature, for a certain amount of sample gas under P_1 pressure and V_1 volume is taken and if we would like to change P_2 pressure or V_2 volume of this gas, we can use the following equation:

P V = k (k is constant)

Do you know that?

Air bags are actually inflated by the equivalent of a solid rocket booster. Sodium azide (NaN_3) and potassium nitrate (KNO_3) react very quickly to produce a large pulse of hot nitrogen gas. This gas inflates the bag, which literally bursts out of the steering wheel or dashboard as it expands.

Figure-2-1

A) Pay attention to the air when some mercury is added.
B) We can observe that some more mercury is added ,volume of air reduced to half

Do you know that?

The pressure on us is the pressure of air surrounding the Earth. The thickness of the atmosphere is 500 miles. In other words, we live under the ocean of atmosphere.

Exercise 2-4

If the pressure of 50 L of gas in a balloon was reduced from 1 atm to 0.9 atm, what will be the final volume of the gas?

$P_1 \times V_1 = P_2 \times V_2$ (constant temperature and constant amount of gas)

Example 2-4

The pressure of perfume in a 0.5 L bottle is 3atm. If applied pressure were 4 atm, what would be its volume?

Solution:

$$P_1 V_1 = P_2 V_2 \qquad P_1, V_1, P_2 \longrightarrow V_2$$

$$V_2 = \frac{P_1 V_1}{P_2} = \frac{(3\ \cancel{atm})(\frac{1}{2}\ L)}{(4\ \cancel{atm})} = 0.375\ L$$

Example 2-5

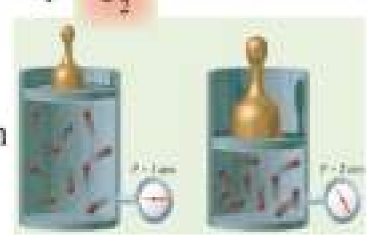
a)1 L of a gas was filled in a cylinder under 1 atm pressure. When a weight was put on it, the volume decreased to 0.5 L. Assuming the temperature was constant, calculate the final pressure.

b) Divers are subject to 1 atm pressure over the surface of water. At constant temperature, in 20 m depth, how much pressure is applied to divers? (Due to the weight of water, assume that pressure increases by 1 atm at every 10m)

Solution:

a) $P_1 V_1 = P_2 V_2 \qquad P_1, V_1, V_2 \longrightarrow P_2$

$$P_2 = \frac{P_1 V_1}{V_2} = \frac{(1\ atm)(1\ \cancel{L})}{(0.5\ \cancel{L})} = 2\ atm$$



b) As the pressure on a diver increases by 1 atm at every 10 m, the pressure in 20 meter depth is 2 atm. Therefore, total pressure applied to the diver is 3 atm to take the pressure over the surface into account.

2-5-2-Temperature and Volume Relationship (Charles' Law)

When temperature of all gases increase, also their volumes increase. To measure volume increase by rise in temperature, a gas with a certain mass is filled into a cylinder with a moving piston as in Figure 2-2.

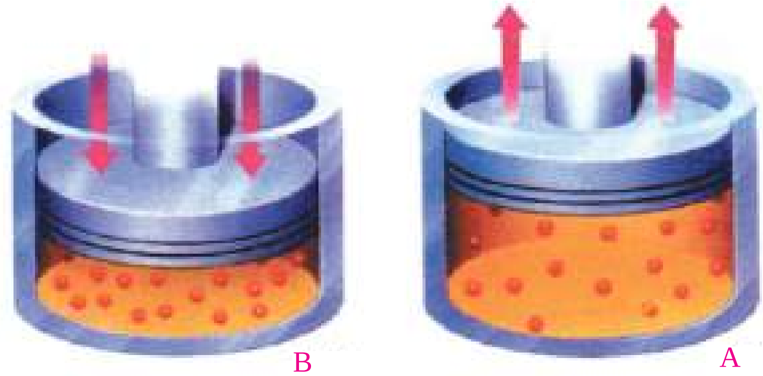


Figure 2-2

A) Volume of heated gas increase and piston moves up.

B) Volume of cooled gas decreases and piston moves down.

When the mass over the piston is constant, the gas is under constant pressure. When the gas is heated, the piston moves upward and the volume of the gas increases. According to Charles' Law, temperature (in Kelvin) and volume of gases change in direct proportion when mass and pressure are constant. Mathematically,

$$V \propto T$$
$$V = k \times T$$

$$\frac{V}{T} = k$$

(k is constant)

Generally, for a certain mass of gas, at T_1 and T_2 temperatures, V_1 and V_2 volumes are considered. In this case, at constant temperature, temperature-volume relationship is as follows:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

At constant pressure and amount of gas

Example 2-6

A balloon was filled with air until its volume reached 4 L at 27°C. What will be its volume when we put it in a fridge at 0°C? (Pressure is the same for both temperatures.)

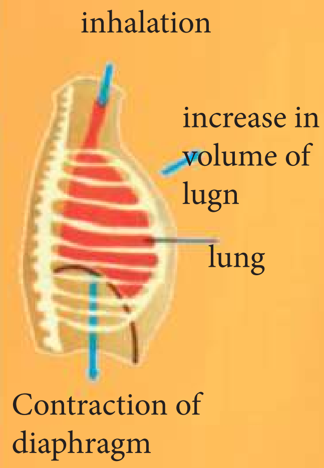
Solution:

We convert temperature from °C to K.

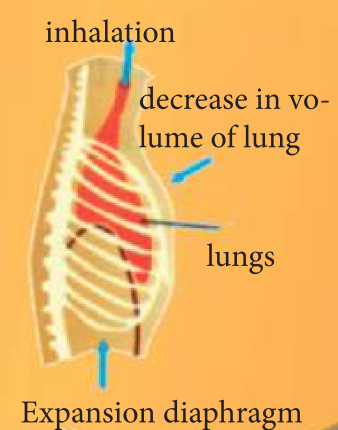
$$T\ (K) = t\,°C + 273$$
$$T_1\ (K) = 27 + 273 = 300\ K$$
$$T_2\ (K) = 0 + 273 = 273\ K$$

Do you know that?

Respiration in humans occurs according to Boyle's Law. Contraction of diaphragm causes expansion of lungs and decrease in pressure in them. Therefore, they let air in and inhaling occurs.



Expansion of diaphragm causes decrease in volume of lungs and increase of pressure in them. By exit of air from lungs, exhaling occurs.



Do you know that?

Bicycle pumps are examples for Charles' Law. When we use them, we notice that they get warm because air molecules in a pump are forced to fit into a smaller area. Therefore, they collide with the walls of the pump more frequently and they warm the surroundings.

After conversion, we find the volume by Charles' Law.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \qquad T_1, V_1, T_2 \longrightarrow V_2$$

$$V_2 = \frac{V_1 T_2}{T_1} = \frac{(4\ \text{L})(273\ \text{K})}{(300\ \text{K})}$$

$$V_2 = 3.64\ \text{L}$$

2-5-3-Pressure-Temperature Relationship (Gay-Lussac's Law)

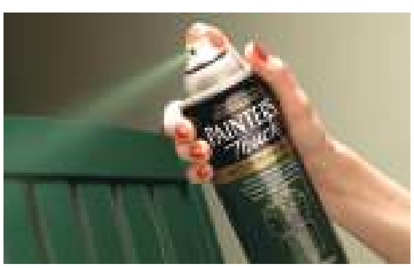
Do you know that?

When an inflated balloon is immersed in cold water, rates of crash of air molecules in the balloon decrease and therefore volume of the balloon decreases.

When with constant volume is heated , its pressure increases

When with constant volume is cooled , its pressure decrease

According to Gay-Lussac's Law, for an ideal gas with constant volume and mass, its pressure is directly proportional to its temperature in Kelvin scale.



The principle of a spray mechanism is that the difference between air pressure and pressure in spray can makes the spray work.

This expression is shown as follows:

$$P \propto T$$

$$P = k \times T$$

$$\frac{P}{T} = k \quad \text{(k : constant)}$$

For gases at two different temperatures (T_1 and T_2) and two different pressures (P_1 and P_2), the following equation are used:

$$\frac{P_1}{T_1}=\frac{P_2}{T_2} \text{ or } \frac{P_1}{P_2}=\frac{T_1}{T_2}$$

(Volume and amount of gas constant)

Example 2-7

Why shouldn't deodorant and spray cans be disposed of in fire? We can understand the reason better with the following question.
The pressure of a spray can is 3 atm at 17°C. What will be its pressure at 187°C?

Solution:

When temperature increases, kinetic energy of molecules in the can increases. Therefore, number of collisions of gas molecules with the walls of the can and also pressure increase. This causes can to burst.
The following steps are applied for solution:
We convert °C to K.

$$T(K) = t(°C) + 273$$

$$T_1(K) = 17 + 273 = 290\ K \qquad P_1 = 3\ atm$$

$$T_2(K) = 187 + 273 = 460\ K \qquad P_2 = ?\ atm$$

We find out P_2 through Gay-Lussac Law.

$$\frac{P_1}{T_1}=\frac{P_2}{T_2} \qquad T_1, P_1, T_2 \longrightarrow P_2$$

$$P_2 = \frac{P_1 T_2}{T_1} = \frac{(3\ atm)(460\ \cancel{K})}{(290\ \cancel{K})}$$

$$P_2 = 4.75\ atm$$

2-6-THE COMBINED GAS LAW

If we re-write mathematical expressions of gas laws;
Boyle's Law ....P.V = k

Charles' Law .... $\frac{V}{T} = k'$

Gay-Lussac's Law................................ $\frac{P}{T} = k''$

A combined law can be written from those 3 laws:

$$\frac{PV}{T} = K \qquad K(\text{Constant})$$

We can write the change in a gas with constant amount from P_1, V_1 and T_1 conditions to P_2, V_2 and T_2 conditions.

Exercise 2-5

1 L of CO_2 is found in a balloon at 27°C. What happens when we put the balloon in a pool at -3°C?

Exercise 2-6

The air pressure of tires of a car is 1.8 atm at 20°C. The car's owner wants to go to Basra. When he arrives at Basra, what will be the air pressure of tires at 36°C?

Do you know that?

Inside a refrigerator, cooling liquid is constantly cycled in the tubes. It rapidly vaporizes and turns into gas while passing through a thin tube. When it is transformed into gas, it absorbs heat in the refrigerator and it cools the surroundings. The gas goes to condenser and it is transformed into liquid again. This condensation is made through pressure.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

The equation above is also called as equation of state for a constant amount of gas.

Example 2-8

A bubble of 2.1 mL in volume is going up from the bottom of a lake at 8°C temperature and 6.4 atm pressure. At the surface, the temperature is 25°C and the pressure is 1 atm. Calculate the volume of the bubble at the surface.

Solution:

We convert °C to K:

$$T\,(K) = t\,(°C) + 273$$

$$T_1\,(K) = 8 + 273 = 281\,K$$

$$T_2\,(K) = 25 + 273 = 298\,K$$

$$\frac{PV}{T} = k$$ k: constant

Do you know that?

We need to check air pressure of tires to avoid unequal corrosion on the surfaces of them.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \qquad P_1, T_1, V_1, P_2, T_2 \longrightarrow V_2$$

$$V_2 = \frac{V_1\,P_1\,T_2}{P_2\,T_1}$$

$$V_2 = \frac{2.1\ mL \times 6.4\ \cancel{atm} \times 298\ \cancel{K}}{1\ \cancel{atm} \times 281\ \cancel{K}}$$

V_2 =14.25 mL (The volume of the bubble at the surface)

2-7-THE RELATIONSHIP BETWEEN AMOUNT OF GAS AND VOLUME (AVOGADRO'S LAW)

Italian scientist Avogadro discovered that volume of a gas and its amount is in direct proportion under constant temperature and pressure. The amount of gas is expressed with number of moles (n). According to this:

$$V \propto n$$

$$V = k\,n$$

For two different amounts of gases (n_1, n_2) and two different volumes (V_1,V_2) of them, we can use the following equation.

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$ (at constant temperature and pressure)

Avogadro's Law is expressed as "equal volumes of all gases, at the same temperature and pressure, have the same number of moles."

We can observe this in Fig 2-3.

Figure-2-3

When we withdraw gas from the environment, so, number of moles decreases. When we add some more gas, number of moles increases and gas volume increases.

2-7-1-MOLAR VOLUME

The value obtained by dividing volume to no. of moles is called as molar volume. As no. of moles is shown with (n) and volume with (V), V_m expresses molar volume.

$$V_m = \frac{V\,(L)}{n\,(mol)} = L\,/\,mol$$

Molar volume of a gas is equal to 22.4 L (22414cm^3) at standard temperature and pressure (STP). (1atm = 760 Torr, 0°C = 273 K).

Molar mass (M) is found by dividing mass (m) to no. of moles (n) as we have studied in Chapter 1.0.0.

$$M = \frac{m\,(g)}{n\,(mol)} = g\,/\,mol$$

Example 2-9

If the volume of 1 mole of hydrogen gas is 22.4 L at standard temperature and pressure, what is the volume of 3 moles of hydrogen gas at the same conditions?

Solution:

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} \qquad V_1, n_1, n_2 \longrightarrow V_2$$

$$V_2 = \frac{V_1 n_2}{n_1} = \frac{(22.4\ L)(3\cancel{mol})}{(1\cancel{mol})}$$

$$V_2 = 67.2\ L$$

Exercise 2-7

The volume of a sample of CO_2 is 4 L at 66°C temperature and 1.2 atm pressure. At 42°C, the volume becomes 1.7 L. Calculate the pressure at this temperature. (The no. of moles hasn't changed.)

Do you know that?

Gay-Lussac discovered the relationship between chemical reactions and volume. But other chemists didn't accept his theory then and they didn't know composition of substances. Therefore, chemical equations couldn't be written at that time. The studies of Gay-Lussac and Avogadro enabled writing of chemical equations.

2-8-IDEAL GAS LAW

We can obtain a single equation by combining 4 gas laws:

$PV = k$ .... Boyle`s Law

$\frac{V}{T} = k'$ .... Charles`s Law

$\frac{P}{T} = k''$ .... Gay-Lussac`s Law

$\frac{V}{n} = k'''$ .... Avogadro Law

$$V \propto n \frac{T}{P}$$

We can get a single equation from the equations above:

When we convert the ratio to an equation:

$$V = (\text{constant})\ n \frac{T}{P}$$

By inserting gas constant (R), we get the final equation will be:

$PV = nRT$ ideal gas law

We can apply the equation above can be applied to gases which are applied 4 laws. This kind of gases are called as ideal gases. R is the general constant for ideal gases(R) (0.082). This equation must be used along with P(atm) , V (L), n (number of moles) and T(K) units in mathematical calculations.

In order to calculate R value, we take 1 mole of an ideal gas with 22.414 L volume at standard temperature and pressure (0°C and 1 atm).

$$PV = nRT$$

$$R = \frac{PV}{nT} = \frac{1\ atm \times 22.414\ L}{1\ mol \times 273\ K}$$

$$R = 0.082 \frac{atm \cdot L}{mol \cdot K}$$

In order to find R value in International System of Units, we take pressure as 101325 Pa (Pascal), volume as 22.4×10^{-3} m^3, no. of moles as 1 mole and temperature as 273K.

$$PV = nRT$$

$$R = \frac{PV}{nT} = \frac{101325\ Pa \times 22.4 \times 10^{-3}\ m^3}{1\ mol \times 273\ K}$$

$$R = 8.314\ Pa.m^3/mol \cdot K$$

Exercise 2-8

Calculate the molar volume of 3 moles of a gas with a volume of 36 L.

Exercise 2-9

If 0.5 mole of a gas has 11.2 L of volume at standard temperature and pressure (STP), what is the no. of mole when it has a volume of 16.8 L?

Pa unit = $\dfrac{Kg}{m.s^2}$

When we use it in the equation above,

$$\dfrac{\dfrac{kg}{\cancel{m}\ .\ s^2}\ m^2 . \cancel{m}}{mol\ .\ K}\ )$$

Unit of energy (J) $\dfrac{kg.m^2}{s^2}$

$$R = 8.314\ J / mol\ .\ K$$

Example 2-10

The volume of NO gas is 5.6 L. Find out the number of moles under standard temperature and pressure.

Solution:

Standard pressure and temperature conditions are 1 atm and 273K.

$PV = nRT$ $\quad P, V, T \longrightarrow n$

$$n = \frac{PV}{RT}$$

$$n = \frac{1\ atm \times 5.6\ \cancel{L}}{0.082\ \cancel{atm}.\cancel{L} / mol.\cancel{K} \times 273\ \cancel{K}}$$

$n = 0.25$ mol

2-8-1-Calculation of Gas Density

We can use ideal gas law to calculate gas density. According to the following equation;

$PV = nRT$........ ....(1)

As , we can insert n in (1).

$$n = \frac{m\ (g)}{M\ (g / mol)}$$

Do you know that?

If we know pressure, temperature and volume of any gas, we can find out its number of moles by making use of ideal gas law.

$$PV = \left(\frac{m}{M}\right)RT \quad \dots\dots\dots \quad (2)$$

or;

$$PM = \left(\frac{m}{V}\right)RT \quad \dots\dots\dots \quad (3)$$

$$\rho = \frac{m(g)}{V(L)}$$ from density formula

Exercise 2-10

Calculate the number of moles of 10 L of oxygen under standard conditions.

ρ shows density and by using it in (3)

We get

$$PM = \rho RT \quad \dots\dots\dots \quad (4)$$

By arranging Equation (4), we can calculate density of a gas with known molar mass and pressure at a certain temperature.

$$\rho = \frac{PM}{RT}$$

We can find out the mass of gas or molar mass from Equation (3):

$$PM = \left(\frac{m}{V}\right)RT$$

By re-arranging

$$m = \frac{PMV}{RT}$$ mass of gas

Example 2-11

Hydrazine (N_2H_4) is used as rocket propellant. Calculate the density of this substance at standard temperature and pressure (STP).

Solution:

$$\rho = \frac{PM}{RT} \qquad P, T, M \longrightarrow \rho$$

Molar mass of hydrazine

$$M(N_2H_4) = (2 \times 14) + (4 \times 1) = 32 \text{ g/mol}$$

At standard conditions, the pressure is 1 atm and the temperature is 273K

$$\rho = \frac{1\,(\cancel{atm}) \times 32\,(g/\cancel{mol})}{0.082\,(L\,.\,\cancel{atm}/\cancel{mol}\,.\,\cancel{K}) \times 273\,(\cancel{K})}$$

$$\rho = 1.43\ g/L$$

Example 2-12

Find out the number of moles of a 700 mL gas at 27°C temperature and 0.8 atm pressure.

Solution:

The following steps are followed for solution:
We convert the volume of the gas from mL to L.

$$V\,(L) = V\ mL \times \frac{1\ L}{1000\ mL} = 700\ \cancel{mL} \times \frac{1\ L}{1000\ \cancel{mL}} = 0.7\ L$$

We convert the temperature from °C to K.

$$T\,(K) = t\,(°C) + 273$$

$$T = 27 + 273 = 300\ K$$

From ideal gas law;

$$PV = nRT$$

$$n = \frac{PV}{RT} \qquad P, V, T \longrightarrow n$$

$$n = \frac{0.8\ \cancel{atm} \times 0.7\ \cancel{L}}{0.082\ \cancel{atm}.\cancel{L}\,/\,mol.\cancel{K} \times 300\ \cancel{K}}$$

$$n = 0.023\ mol$$

Example 2-13

The pressure of a gas in a 3 litre container was found as 5.46 atm at 27°C temperature. As the molar mass of the gas is 44 g/mol, find out the mass and mole number of the gas.

Solution:

The following steps are applied for solution:

1. We convert the temperature from °C to K.

$$T\,(K) = t\,(°C) + 273$$

$$T = 27 + 273 = 300\ K$$

2. Using the following equation;

$$m = \frac{MPV}{RT} \qquad P, V, T, M \longrightarrow m$$

$$m = \frac{44\ g/\cancel{mol} \times 5.46\ \cancel{atm} \times 3\ \cancel{L}}{0.082\ \cancel{atm}.\cancel{L}\ /\ \cancel{mol}.\cancel{K} \times 300\ \cancel{K}}$$

$m = 29.3$ g

3. Calculate the mole number;

$$n = \frac{m\ (g)}{M\ (g/mol)}$$

$$n = \frac{29.3\ \cancel{g}}{44\ \cancel{g}/mol} = 0.67\ mol$$

Exercise 2-11

Calculate the density of oxygen gas (O_2) at 373K temperature and 5 atm pressure in g/L unit.

Example 2-14

At 227°C temperature and 748 Torr pressure, the mass of a gas is 0.6 g in a 500 mL container. Calculate the molar mass of the gas.

Solution:

The following steps are applied for solution:

1) We convert the volume from mL to L.

$$V\ (L) = V\ mL \times \frac{1\ L}{1000\ mL} = 500\ \cancel{mL} \times \frac{1\ L}{1000\ \cancel{mL}} = 0.5\ L$$

Exercise 2-12

Methane is a gas obtained by oil refining. Find out the volume of 0.5 mole of a sample from methane gas at 27°C temperature and 3 atm pressure.

We convert the temperature from °C to K.

$$T\ (K) = t\ (°C) + 273$$

$$T = 227 + 273 = 500\ K$$

2) We convert the pressure from Torr to atm.

$$P\ atm = P\ Torr \times \frac{1atm}{760Torr} = 748\ \cancel{Torr} \times \frac{1atm}{760\cancel{Torr}} = 0.984\ atm$$

3) From Equation $M = \dfrac{mRT}{PV}$

$$M = \frac{0.6\ g \times 0.082\ \cancel{atm}.\cancel{L}\ /\ mol.\cancel{K} \times 500\cancel{K}}{0.984\ \cancel{atm} \times 0.5\ \cancel{L}}$$

$M = 50$ g /mol

Example 2-15

The molar mass of 0.31 g of a gas is 32 g/mol and it has a pressure of 1.17 atm. What is the temperature when this gas has a volume of 0.23 L?

Solution:

$$PV = \frac{m}{M} RT$$

$$T = \frac{PVM}{mR} \qquad P, V, m, M \longrightarrow T$$

$$T = \frac{1.17\ (\text{atm}) \times 0.23\ (\text{L}) \times 32\ (\text{g/mol})}{0.31\ (\text{g}) \times 0.082\ (\text{L . atm/mol . K})}$$

$$T = 339\ \text{K}$$

2-9-DALTON’S LAW OF PARTIAL PRESSURES

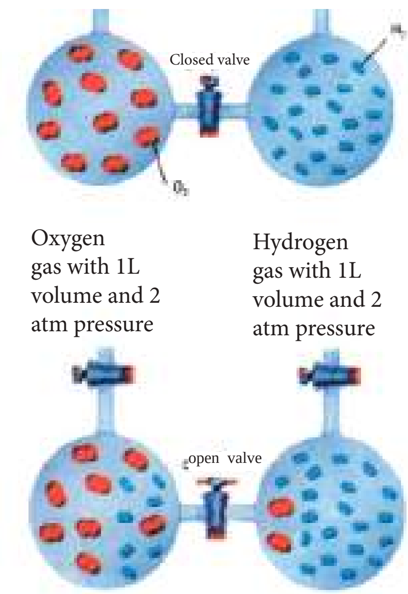
So far, we have dealt with only one type of gas. But what happens when we handle a gas mixture which doesn’t interact such as air? When Dalton studied a sample of air, he found out that the total gas pressure is equal to the sum of partial pressures of gases. Partial pressure is the pressure of a gas in a gas mixture.

Dalton’s law of partial pressures is expressed as “Total pressure of a gas mixture in which gases don’t react is equal to sum of partial pressures

$$P_T = P_1 + P_2 + P_3 + \ldots\ldots$$

of all gases.” Mathematically,

P_T show total gas pressure whereas P_1 , P_2 and P_3 …. show partial pressures of gases in the mixture.



When we open the valve , gases diffuse in both bottles. According to Dalton’s Law ,pressure of both gases change.

2-9-1-The Relationship between Total Pressure, Total Number of Moles and Mole Fraction

Assume that there is a gas mixture of two gases with known volume and pressure. We find the partial pressures using the ideal gas law as follows.

$$P_1 = \frac{n_1 R T}{V} \quad \ldots\ldots\ldots\ \ldots\ldots\ldots\ \ldots\ldots\ldots\ \ldots\ldots\ldots \quad (1)$$

$$P_2 = \frac{n_2 R T}{V} \quad \ldots\ldots\ldots\ \ldots\ldots\ldots\ \ldots\ldots\ldots\ \ldots\ldots\ldots \quad (2)$$

From Dalton’s Law

$$P_T = P_1 + P_2 \ldots\ldots\ldots\ldots\ldots\ldots\ldots\ldots\ldots\ldots\ldots 3$$

We insert Equation 1 and 2 in Equation 3.

$$P_T = \frac{n_1 R T}{V} + \frac{n_2 R T}{V} = (n_1 + n_2)\frac{R T}{V} \quad \ldots\ldots\ldots\ \ldots\ldots\ldots \quad (4)$$

By dividing Eq. 1 to Eq. 4;

$$\frac{P_1}{P_T} = \frac{n_1 \frac{R T}{V}}{(n_1 + n_2)\frac{R T}{V}} \quad \ldots\ldots\ldots\ \ldots\ldots\ldots\ \ldots\ldots\ldots\ \ldots\ldots\ldots \quad (5)$$

Exercise 2-13

At 127°C temperature and 3.65 atm pressure, find out the molar mass of a gas with 900 mL volume and 4.41 g mass.

Exercise 2-14

Calculate the molar mass of a gas with 0.4 g mass and 280 mL of volume at standard temperature and pressure (STP).

Do you know that?

What prevents molecules in the atmosphere from escaping to space? Gravity is applied on molecules as it is applied on us. Molecules need to have a velocity of 1.1x10^3 m/s to escape from this gravitational force. As molecules of helium gas are much faster than this limit, they can escape to space. Therefore, there isn't much helium gas in the atmosphere.

By eliminating similar terms;

$$\frac{P_1}{P_T} = \frac{n_1}{n_1+n_2} \quad \text{.................................... (6)}$$

We get the equation

n_T shows total no. of moles in the mixture. It is found by adding no. of moles of gases in the mixture.

$$n_T = n_1 + n_2$$

Thus, Equation 6 becomes as follows:

$$\frac{P_1}{P_T} = \frac{n_1}{n_T} \text{.............................(7)}$$

Similarly, for the second gas in the mixture;

$$\frac{P_2}{P_T} = \frac{n_2}{n_T} \text{.............................(8)}$$ equation is obtained.

Mole Fraction: The ratio of no. of moles (for 1st or 2nd) of each component in a mixture to total no. of moles of components gives mole fraction. Mole fraction of 1st component is shown as (x_1).

$$x_1 = \frac{n_1}{n_1+n_2} = \frac{n_1}{n_T}$$

Mole fraction of 2nd component is shown as (x_2).

$$x_2 = \frac{n_2}{n_T}$$

When we insert mole fraction in Equations 7 and 8;

$$\frac{P_1}{P_T} = x_1 \quad , \quad \frac{P_2}{P_T} = x_2 \quad \text{.................................... (9)}$$

we get these equations:

$$P_1 = x_1 \times P_T \quad , \quad P_2 = x_2 \times P_T \quad \text{.................. (10)}$$

We can re-write Eq. 10 as follows:

$$P_i = x_i \times P_T$$

x_i shows mole fraction of a component and P_i shows its partial pressure. Besides, sum of mole fractions of all gases in a mixture is equal to 1. We can apply this to the previous two-gas mixture:

$$x_1 + x_2 = \frac{n_1}{n_1+n_2} + \frac{n_2}{n_1+n_2} = \frac{n_1+n_2}{n_1+n_2} = 1$$

Generally, we can write the following equation:

$$x_1 + x_2 + x_3 + x_4 \ldots = 1$$

Example 2-16

In a mixture of noble gases, there are 4.46 moles of Ne, 0.74 mole of Ar and 2.15 moles of Xe. Calculate the partial pressures of each gas. (Temperature is constant and total pressure is 2 atm.)

Solution:

The following steps are applied for solution:

1) Total number of moles is found.

$$n_T = n_{Ne} + n_{Ar} + n_{Xe}$$

$$n_T = 4.46 \text{ mol} + 0.74 \text{ mol} + 2.15 \text{ mol}$$

$$n_T = 7.35 \text{ mol}$$

2) Mole fractions of noble gases are found.

$$x_{Ne} = \frac{n_{Ne}}{n_T} = \frac{4.46 \text{ mol}}{7.35 \text{ mol}} = 0.607$$

$$x_{Ar} = \frac{n_{Ar}}{n_T} = \frac{0.74 \text{ mol}}{7.35 \text{ mol}} = 0.100$$

$$x_{Xe} = \frac{n_{Xe}}{n_T} = \frac{2.15 \text{ mol}}{7.35 \text{ mol}} = 0.293$$

3) Using the following formula, partial pressure of each gas is found.

$$P_i = x_i \times P_T$$ As

We write symbol of each gas in place of i in the formula.

$$P_{Ne} = P_T \ x_{Ne} = 2 \text{ atm} \times 0.607 = 1.214 \text{ atm}$$

$$P_{Ar} = P_T \ x_{Ar} = 2 \text{ atm} \times 0.100 = 0.200 \text{ atm}$$

$$P_{Xe} = P_T \ x_{Xe} = 2 \text{ atm} \times 0.293 = 0..586 \text{ atm}$$

To make sure, we add up partial pressures. (The sum must be 2.)

$$1.214 \text{ atm} + 0.200 \text{ atm} + 0.586 \text{ atm} = 2.000 \text{ atm}$$

Example 2-17

Potassium chloride releases oxygen gas under extreme heat through MnO_2 catalyst. In order to collect oxygen gas, water is used. Calculate the mass of gas in grams with 128 mL of volume at 24°C and 762.4 mmHg pressure. (At 24°C, vapor pressure of water is 22.4 mmHg and molar mass of oxygen is 32g/mol.)

Exercise 2-15

Gases obtained by oil refining were placed into a container. There are 6 moles of methane, 4 moles of ethane and 2 moles of propane in it. As the total pressure is 6 atm, calculate partial pressures of each gas.

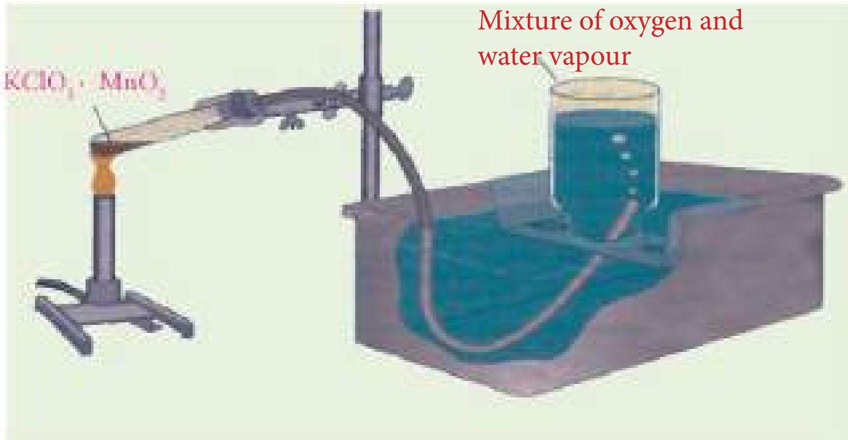
Exercise 2-16

When Ca and water react at 30°C and 988 mmHg pressure, H_2 gas is released. Water is used to collect the gas. Calculate the mass of H_2 gas which is 200 mL in grams. (Molar mass of H_2 gas is 2 g/mol. Water vapor pressure at 30°C is 32.82 mmHg.)

Solution:

The following steps are applied for solution;

1) Collected gases were mixed with water vapor. Therefore, in the tube, there is a mixture of gas and water vapor.



The total pressure of oxygen and water vapor mixture is

764.2 mmHg. We need to find out the partial pressure of oxygen gas.

$$P_T = P_{O_2} + P_{H_2O}$$

$$762.4 = P_{O_2} + 22.4$$

$$P_{O_2} = 740 \text{ mmHg}$$

2) We convert pressure from mmHg to atm.

$$P \text{ atm} = P \text{ mmHg} \times \frac{1\text{atm}}{760 \text{ mmHg}} = 740 \ \cancel{\text{mmHg}} \times \frac{1\text{atm}}{760 \ \cancel{\text{mmHg}}}$$

$$= 0.974 \text{ atm}$$

3) We convert volume from mL to L.

$$V(L) = V \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 128 \ \cancel{\text{mL}} \times \frac{1 \text{ L}}{1000 \ \cancel{\text{mL}}} = 0.128 \text{ L}$$

4) We convert °C to K.

T(K) =t(°C) +273 = 24+273 =297K

5) By using ideal gas law;

$$PV = \frac{m}{M} RT$$

mass of gas

$$m = \frac{PVM}{RT}$$

Exercise 2-17

The partial pressures of gases in a sample of air are as follows: Nitrogen = 569 Torr, Oxygen = 116 Torr, Carbon dioxide = 28 Torr and water vapor = 0.47 Torr. Calculate the percentages of these gases in air through mole fractions.

$$m = \frac{0.974\ (\cancel{atm}) \times 0.128\ (\cancel{L}) \times 32\ (g/\cancel{mol})}{0.082\ (\cancel{L}.\ \cancel{atm}/\cancel{mol}\ .\ \cancel{K}) \times 297\ \cancel{K}}$$

m =0.164g

Example 2-18

At 7 °C, there are 3.2 g of oxygen, 0.4 g of helium and 14 g of nitrogen gases mixture in a 2 L test tube. Calculate the total pressure of this mixture. (Molar masses of gases: Oxygen = 32 g/mol, nitrogen = 28 g/mol, helium = 4g/mol)

Solution:

The following steps are applied for solution:

1)
$$n_{He} = \frac{m(g)}{M\ (g/mol)} = \frac{0.4\ \cancel{g}}{4\ \cancel{g}/mol} = 0.1 \quad mol$$

$$n_{O_2} = \frac{3.2\ \cancel{g}}{32\ \cancel{g}/mol} = 0.1 \quad mol$$

$$n_{N_2} = \frac{14\ \cancel{g}}{28\ \cancel{g}/mol} = 0.5 \quad mol$$

2) We convert °C to K.

$$T\ (K) = t\ (°C) + 273 = 7 + 273 = 280\ \ K$$

3) We find out the total number of moles of the mixture.

$$n_T = n_{He} + n_{O_2} + n_{N_2}$$

$$n_T = 0.1 + 0.1 + 0.5 = 0.7\ mol$$

$$P_T = \frac{n_T RT}{V}$$

4) By using ideal gas law;

$$P_T = \frac{0.7\ \cancel{mol} \times 0.082\ atm.\cancel{L}/\cancel{mol}.\cancel{K} \times 280\ \cancel{K}}{2\ \cancel{L}} = 8.036\ atm$$

Example 2-19

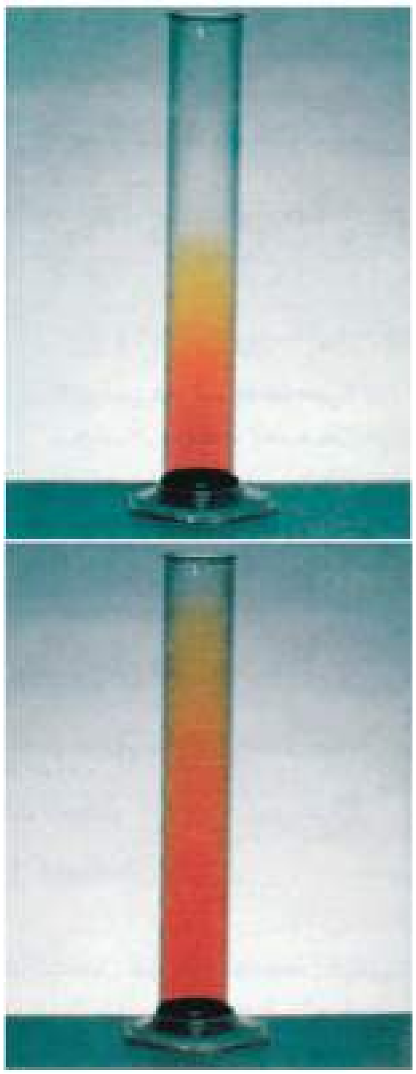
Two containers in different volumes were connected via a valve. The volume of the first container was 1 L and it had CO_2 gas under 720 Torr pressure. The volume of the second container was 2 L and it had nitrogen as under 540 Torr pressure. After the valve was switched on and equilibrium established, what was the total pressure in Torr? (Temperature was constant).

Do you know that?

Under oceans, very important studies have been carried out and experienced divers perform these studies. Divers are subject to some amount of pressure underwater. Even there were a pipe immersed from surface to underwater for their breathing, this wouldn't be of any use. The reason is that lungs of divers cannot expand therefore they can't breathe. Thus, divers carry special tubes on their backs. These tubes carry an automatic instrument. Tubes have a special gas mixture to balance air going into lungs. This mixture consists of helium, oxygen and a little amount of nitrogen and it is used in deep water operations.

Exercise 2-18

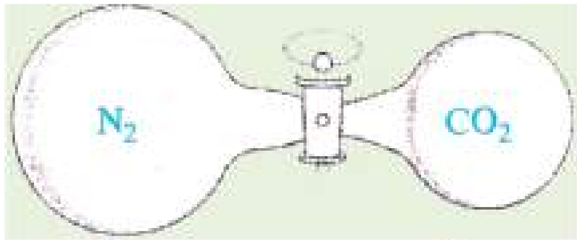
20 mL of N_2 under 740 Torr pressure was mixed with 30 mL of O_2 under 640 Torr pressure in a 50 mL container. Assuming the temperature was constant, find out the total pressure of the mixture.



Do you know that ?

Bromine gas diffuse inside the tube a few hours later after it is placed in the tube.

Solution:



The following steps are applied for solution:

1) After the valve was switched on, the gases were mixed and total volume was equal to the sum of two volumes.

$$V_2 = V_{N_2} + V_{CO_2} = 2 + 1 = 3\ L$$

2) Through Boyle's Law, we find out partial pressures of each gas.

$$P_1 V_1 = P_2 V_2 \Rightarrow P_2 = \frac{P_1 V_1}{V_2}$$

Partial pressure of CO_2 gas

$$P_1 = \frac{720\ \text{Torr} \times 1\ \cancel{L}}{3\ \cancel{L}} = 240\ \text{Torr} = P_{CO_2}$$

Partial pressure of N_2gas

$$P_2 = \frac{540\ \text{Torr} \times 2\ \cancel{L}}{3\ \cancel{L}} = 360\ \text{Torr} = P_{N_2}$$

3) Total pressure of the mixture;

$$P_T = P_{CO_2} + P_{N_2}$$

$$P_T = 240\ \text{Torr} + 360\ \text{Torr} = 600\ \text{Torr}$$

2-10-GRAHAM'S LAW OF DIFFUSION

Graham discovered during his studies that the diffusion rates of gases which pass through tiny holes are inversely proportional to square root of gas densities. At the same time he found out that diffusion rates of gases which pass through tiny holes are inversely proportional to square root of molar masses (M).

Let the diffusion rates of two different gases through the same hole be r_1 and r_2, their densities be ρ_1 and ρ_2. According to Graham's law;

$$\frac{r_1}{r_2} = \sqrt{\frac{\rho_2}{\rho_1}} = \sqrt{\frac{M_2}{M_1}}$$

M_1 and M_2 are molar masses of both gases respectively.

If we apply diffusion law to diffusion of hydrogen and oxygen gases through the hole;

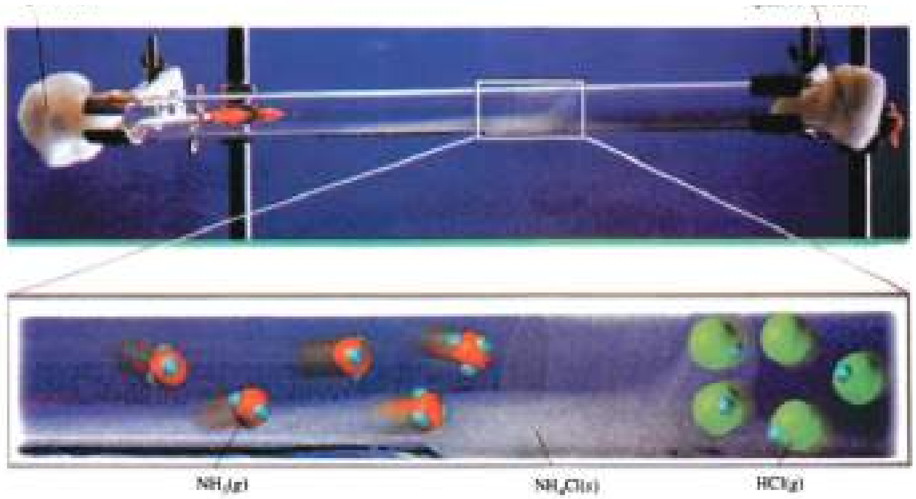
$$\frac{r_{H_2}}{r_{O_2}} = \sqrt{\frac{\rho_{O_2}}{\rho_{H_2}}} = \sqrt{\frac{M_{O_2}}{M_{H_2}}}$$

M_{H_2} , ρ_{H_2} , r_{H_2} show rate of hydrogen gas, its density and molar mass respectively.

show rate of oxygen gas, its density and molar mass respectively.

Ammonia NH_3
Cotton cork

Hydrogen Chloride HCl
Cotton cork



Exercise 2-19

The diffusion rate of O_2 gas is 8 mL/s and the diffusion rate of H_2 gas is 32 mL/s. As molar mass of O_2 is 32 g/mol, find out the molar mass of H_2 gas.

The molar mass of hydrogen chloride is 36.5 g/mol whereas the molar mass of ammonia is 17 g/mol. Therefore, ammonia has a higher diffusion rate than hydrogen chloride .

Example 2-20

If the diffusion rate of oxygen gas through a porous obstacle is 8 mL/s, Find out the diffusion rate of hydrogen gas through the same obstacle. (Densities of gases at the same temperature and pressure $\rho(O_2)$=144g/L, $\rho(H_2)$ = 0.09 g/L).

$$\frac{r_{H_2}}{r_{O_2}} = \sqrt{\frac{\rho_{O_2}}{\rho_{H_2}}}$$

$$\frac{r_{H_2}}{8\ mL/s} = \sqrt{\frac{1.44\ g/L}{0.09\ g/L}}$$

$$\frac{r_{H_2}}{8\ mL/s} = \sqrt{16}$$

$$r_{H_2} = 32\ mL/s$$

Diffusion rate of hydrogen gas

Example 2-21

If diffusion rate of nitrogen gas through a tiny hole is 2.65 mL/s, find out the diffusion rate of NH_3 gas through the same hole. (Molar mass of N_2 is 28 g/mol, molar mass of NH_3 is 17 g/mol.)

Solution:

$$\frac{r_{N_2}}{r_{NH_3}} = \sqrt{\frac{M_{NH_3}}{M_{N_2}}}$$

$$\frac{2.65\ mL/s}{r_{NH_3}} = \sqrt{\frac{17\ g/mol}{28\ g/mol}}$$

Taking the square root of both sides;

$$\frac{7.0225}{r^2_{NH_3}} = \frac{17}{28}$$

$$r^2_{NH_3} = 11.56\ (ml/s)^2$$

$$r_{NH_3} = 3.40\ ml/s$$ diffusion rate of ammonia

2-10-1-The Relationship between Diffusion Rate and Diffusion Time

Diffusion rate of a gas is inversely proportional to diffusion time for a given temperature and pressure. That means as diffusion rate of a gas increases, diffusion time decreases.

$$\frac{r_1}{r_2} = \frac{t_2}{t_1}$$

t_1 and t_2 are diffusion times of the 1st and 2nd gases respectively. If we combine Graham's laws in one single law;

$$\frac{r_1}{r_2} = \frac{t_2}{t_1} = \sqrt{\frac{\rho_2}{\rho_1}} = \sqrt{\frac{M_2}{M_1}}$$ we get this equation:

Example 2-22

A sample of hydrogen gas diffuses in 5 seconds through a hole. At the same conditions, another sample of a gas diffuses in 20 seconds through the same hole. As the molar mass of hydrogen is 2 g/mol, find out the molar mass of the other gas.

Solution:

$$\frac{r_1}{r_2} = \frac{t_2}{t_1}$$

$$\frac{t_2}{t_1} = \sqrt{\frac{M_2}{M_1}}$$

$$\frac{20\ s}{5\ s} = \sqrt{\frac{M_2}{2\ g/mol}}$$

Squaring both sides; $4 = \sqrt{\frac{M_2}{2\ g/mol}}$

$$16 = \frac{M_2}{2\ g/mol}$$

$$M_2 = 32\ g/mol$$

2-11-KINETIC THEORY OF GASES

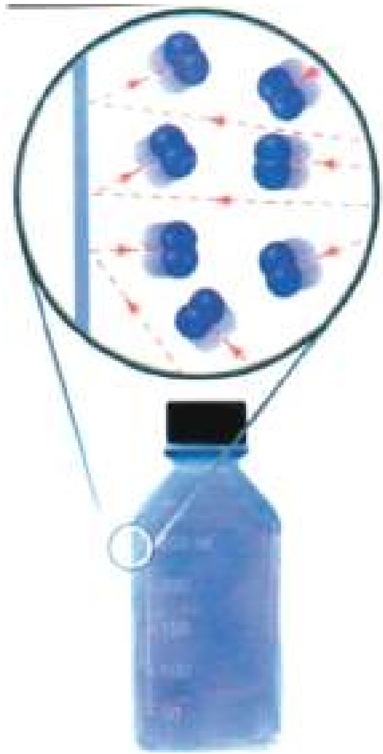
Gas laws which we have studied previously were formed as a result of experiments by scientists at their times. These laws weren't derived from a theory or depended on one. Countless studies had been made in order to explain behaviors of gases theoretically. As a result, in the light of some assumptions, relationships between theory and experimental knowledge were established. The results obtained from those assumptions are as follows;

1) Gases are made up of molecules in great numbers. With respect to volume that a gas occupies, volume of a single molecule can be neglected. In other words, intermolecular distance is too big.

2) The movements of gases are fast, random and linear. Therefore, gas molecules collide either with themselves or with the walls of the container they are inside.

3) There are no attraction or repulsion forces between gas molecules.
Pressure applied by gas molecules arise from collisions with the walls of container they are inside.

4) Gas molecules move at different speeds. Average speed of molecules is directly proportional to absolute temperature of gas.



Do you know that?

Molecules of gases move linearly until they collide with walls of the container they are inside.

2-12-REAL AND IDEAL GASES

Gases which obey gas laws or general gas equation under any temperature and pressure conditions are called as ideal gases. In fact, gases show distinct diversions from ideal behavior. Ideal state is only observed under certain temperature and pressure values.
Gases which deviate from ideal behavior are called as real or non-ideal gases. Diversions by real gases arise from two wrong hypotheses of kinetic theory of gases (This can be possibly true at low pressures.)

1) With respect to total volume of a gas, the volume occupied one of its molecules is almost zero. But at high pressures, gas molecules must occupy a certain volume. Otherwise gas cannot be turned into liquid or solid.

2) There is an attraction force between gas molecules. But in fact, there is a general attraction force. Otherwise molecules would not approach each other and gas wouldn't turn into liquid or solid.
Despite those above, ideal gas laws can be applied on real gases by considering

Exercise 2-20

Xenon gas needs 1 min. 8.3 sec. to pass through a hole. Find out the molar mass of a gas which passes through this hole in 57 sec. (Molar mass of Xe is 131.3 g/mol)

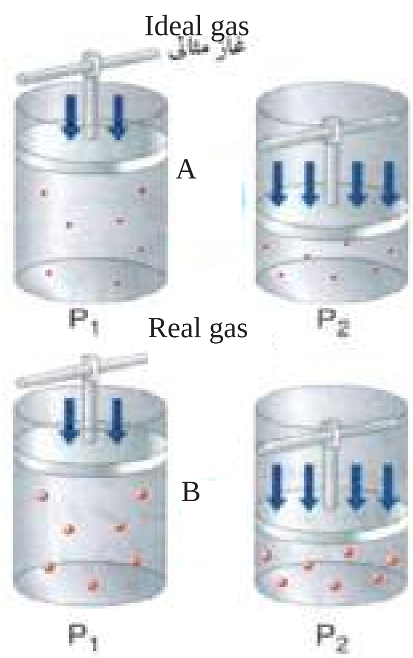
Exercise 2-21

Explain the reason why ammonia molecules diffuse faster than perfume molecules.

Exercise 2-22

Explain what the following statement means.

A gas doesn't show an ideal gas behavior at all temperature and pressure values. Under what conditions do real gases show ideal gas behavior and what is the reason for that?



A) Ideal gas volumes molecules are small cab be neglected
B) Real – gas volumes of gas molecules cannot be neglected

Do you know that ?

Cryostats are containers which were designed to store liquefied gases. Due to their design, heat cannot reach extremely cold liquid inside them .Most widely used type of cryostats are Dewar flask which were named after Scottish Scientists Sir James Dewar who designed them in 1892. Those are like thermoses used to carry hot and cold drinks. They consists of two walls with vacuum between them. Cryostats are very lightweight compared with vacuum gas tubes. Under high pressure ,liquid form any substance has much little volume than its gas form. Therfore, many gases are stored and carried in liquid form.

different accuracy ratios. This ratio increases by rise in temperature and drop in pressure but decreases by drop in temperature and rise in pressure.

Do you know that?

At sea level, average of number of collisions between air molecules is 7 billion in a square metre per second. But at an altitude of 600 km from sea level, it is 1 collision per minute.

2-13-CRITICAL PRESSURE-CRITICAL TEMPERATURE AND LIQUEFACTION OF GASES

According to kinetic theory of gases, gases move randomly. At high temperature and low pressure values, a gas molecule moves freely and isn't affected from other molecules. When temperature decreases, kinetic energy of gas molecules also decreases. Therefore, molecular movement slows down.

When a certain low temperature is reached, gas turns into liquid. On the other hand, with increasing pressure, gas molecules approach each other more. Therefore, as a result of a continuous increase in pressure and decrease in volume, gas turns into liquid. The temperature at which gas turns into liquid is called as critical temperature and the pressure is called as critical pressure. In other words, critical temperature is the temperature at which gas cannot turn into liquid no matter how much pressure is applied. Critical pressure is the pressure necessary to be applied on a gas at critical temperature for it to turn into liquid. Besides, the volume of 1 mole of a gas under critical temperature and pressure is called as critical volume. In Table 2-2, critical temperature and pressures of some gases are given.

Table 2-2 Critical Temperature and pressure values of some Gases

Gas	Critical Temperature (^0C)	Critical pressure (atm)
Helium	-267.9	2.26
Hydrogen	-239.9	12.8
Nitrogen	-147	33.5
Oxygen	-118.4	50.1
Carbon dioxide	+31	50.1

2-14-LIQUID VAPOR PRESSURE

In a closed container at constant temperature, some molecules of a liquid leave from liquid surface to outer medium. As the medium is closed, vapor molecules cannot get outside, but they collide with each other and with the walls of the container. Therefore, energies of some molecules are transferred to other molecules and they turn into liquid again. The first case is called as evaporation whereas the second case is called as condensation. Molecules in vapor form apply a type of pressure. When equilibrium is reached, this pressure is called as vapor pressure as a property of the liquid. Vapor pressure is the pressure of vapor molecules which are at equilibrium with liquid molecules at a certain temperature.

Figure 2-4

In an open container , molecules of liquid leave the container when heated (vaporization). But in a closed container , liquid molecules remain in the container (condensation)

Do you know that?

Absolute zero is the temperature which is equal to 0K (-273°C). Theoretically, volume of a gas becomes zero at this temperature. But practically this isn't possible because gases liquefy before reaching this temperature or sometimes just as in carbon dioxide case, they solidify. The lowest temperature which can be reached by scientists is -269°C for helium.

2-15-BOILING POINTS OF LIQUIDS

Increase in temperature causes a rise in kinetic energy of liquid molecules. This rise decreases attraction force between molecules. As a result, number of gas molecules over the liquid surface increases and also vapor pressure goes up. When vapor pressure of the liquid and atmospheric pressure are equal, the liquid starts to boil. Therefore, boiling point of a liquid is the temperature at which vapor pressure of the liquid is equal to atmospheric pressure.

CHEMICAL IMPACT

READING PART

The Chemistry of Air Bags

Most experts agree that airbags represent a very important advance in autumobile safety. These bags, which are stored in the auto's steering wheel or dash, are designed to inflate rapidly (within about 40 ms) in the event of crash, cushioning in the front-seat occupants against impact. The bags then the deflate immediately to allow vision and movement after the crash. Airbags are activated when a severe deceleration (an impact) causes a steel ball to compress a spring and electrically egnite a detonator cap, which, in turn, causes sodium azide (NaN_3) to decompose explosively, forming sodium and nitrogen gas:

$$2NaN_3(s) \longrightarrow 2Na(s) + 3N_2(g)$$

This system works very well and requires a relatively small amount of sodium azide (100 g yields 56 L $N_2(g)$ at 25°C and 1.0 atm).

When a vehicle containing airbags reaches the end of its useful life, the sodium azide present in the activators must be given proper disposal. Sodium azide, besides being explosive, has a toxicity roughly equal to that of sodium cyanide. It also forms hydrazoic acid (HN_3), a toxic and exclusive liquid, when treated with acid.

The airbag represents an application of chemistry that has already saved thousands of lives.

Inflated airbags

BASIC CONCEPTS

Volume: Volume is the area which matter occupies. Volume of a gas is as much as the volume of the container it is inside.

Pressure: Pressure is the force (F) applied to per unit area (A)

Boyle's Law: At constant temperature, the product of volume and pressure of some amount of ideal gas is constant.

Charles' Law: At constant pressure, change in volume of any amount of ideal gas is directly proportional to change in its temperature in Kelvin.

Gay Lussac's Law: Pressure of a constant amount of an ideal gas with a constant volume is directly proportional to its temperature in Kelvin.

Avogadro's Law: Equal volumes of different gases contain equal numbers of particles at constant pressure and temperature.

Dalton' Law of Partial Pressures: Total pressure of a gas mixture with gases which don't react with each other is equal to sum of partial pressures of all gases in the mixture.

Mole Fraction: It is the ratio of number of moles of components in a gas mixture over total number of moles.

Graham's Law of Diffusion: Diffusion rate of gases through small holes is inversely proportional to square root of its density.

Ideal Gas: It is a gas model in which volumes of molecules are negligibly small when compared to total volume; intermolecular attraction and repulsion forces don't exist and intermolecular collisions are elastic.

Real Gas: A gas which deviates from ideal behavior is called as a real gas or a non-ideal gas.

Critical Temperature: It is the temperature at which a gas cannot turn into liquid no matter how much pressure is applied.

Critical Pressure: It is the pressure necessary for a gas at critical temperature to turn into liquid.

Critical Volume: Volume of 1 mole of a gas at critical pressure and temperature.

Vapor Pressure: It is the pressure of vapor molecules which are at equilibrium with liquid molecules at a certain temperature.

Boiling Temperature: It is the temperature at which vapor pressure of a liquid and atmospheric pressure are equal.

QUESTIONS OF CHAPTER- 2

Reminder: In some questions, you may need to know atomic masses. You can refer to charts at the end of the book for information.

2.1) 0.5 L cylinder of a car engine (combustion chamber) was filled with gasoline and air mixture at 1 atm. What is the necessary pressure to be applied to have a volume of 57 L before ignition by the spark plug? (Assume gasoline vapor and air as a single gas.)

2.2) The volume of a balloon filled with helium gas is 50 L at 25°C under 1.08 atm pressure. What is the volume of it at 10°C and 0.0885 atm pressure conditions?

2.3) Calculate the volume of a sulfur hexafluoride (SF_6) sample under standard conditions which has a volume of 200 mL at 27°C and 570 atm pressure.

2.4) What is the volume of 5 g of acetylene gas (C_2H_2) (one of the components of oxyacetylene flame) at 50 °C and 740 Torr pressure?

2.5) The volumes of 3.7 g of a gas sample at 25 °C and 0.184 g of hydrogen gas at 17 °C are the same. As the gases are under the same pressure, what is the molar mass of the gas sample?

2.6) 45 mL of hydrogen gas was collected at 25°C and 754 Torr pressure as a result of an experiment to obtain hydrogen gas through the reaction of magnesium element and hydrochloric acid. As vapor pressure of water is 23.8 Torr at 25°C, calculate the number of moles of H_2 gas obtained from this experiment.

2-7) Calculate mole fraction of each component as total pressure of the gas mixture which has 78% mole of nitrogen and 22% mole of oxygen is 1.12 atm. What are the partial pressures of components?

2.8) Average rate of fluorine molecules under certain temperature and pressure is 0.038 m/s. What is the average rate of sulfur dioxide (SO_2) gas molecules under the same conditions?

2.9) Circle the correct answer for the following questions.

1. Which of the following can be the sample with a density of 1.6g/L at 26°C temperature and 680.2 mmHg pressure?

A) CH_4 B) C_2H_6 C) CO_2 D) SF_6

2. At which of the following temperature and pressure conditions does helium gas have a molar volume of 51.4 L/mol?

A) 25°C temperature, 0. 25 atm pressure

B) 0°C temperature, 0.5 atm pressure

C) 300°C temperature, 1.00 atm pressure

D) 40°C temperature, 0.50 atm pressure

3. Which of the following is the diffusion rate of O_2 gas at the same temperature?

A) 4 times greater than the diffusion rate of He gas.

B) 2.08 times greater than the diffusion rate of He gas.

C) 0.35 times greater than the diffusion rate of He gas.

D) 0.125 times greater than the diffusion rate of He gas.

4. Which of the following is the number of moles of He gas with 22.4 litres volume at 1 atm pressure and 30°C temperature?

A) 0.11 mol B) 1.00 mol C) 0.90 mol D) 1.11 mol

5. A gas occupies a volume of 430 mL at 754.2 Torr pressure and 28.2°C. If this gas is cooled to 20°C, what will be the pressure in Torr?

A) 534.9 Torr B) 733.7 Torr C) 775.3 Torr D) 842.3 Torr

6. In which of the following cases can Charles' Law be applied?

A) Pressure change

B) Constant temperature

C) A temperature range

D) Very low pressure

7. In a 1 liter container, there are 3 non-reacting gases. According to this, what is the pressure of the first gas?

A) 1/3 of total pressure

B) Total pressure-partial pressures of other gases

C) Number of its molecules

D) Always atmospheric pressure

8. If we assume that some amount of a gas has a volume of 117 cm^3 at 39 °C, at which temperature will the volume of the gas be 213 cm^3?

A) 39 ^{0}C B) 78 ^{0}C C) 395 ^{0}C D) 295 ^{0}C

9. A gas has a volume of 20 litres at 760 mmHg pressure. Which of the following is the volume at 38 mmHg pressure?

A) 10 L B) 20 L C) 40 L D) 400 L

10. In which of the following does 2 g of hydrogen gas have the largest volume?

A) 0 ^{0}C, 1 atm B) 273 ^{0}C, 380 mmHg C) 273 ^{0}C, 120 mm Hg D) 17 ^{0}C, 700 mmHg

2.10) What is the mass of Cl_2 gas found in a 10 L container at 27°C under 3.05 atm pressure? (Atomic mass of Cl is 35.5 g/mol)

2.11) What is the molar mass of a gas sample which has a mass of 1.25g and 1 L volume at 27 °C and 0.961 atm pressure?

2.12) Helium gas in a meteorology balloon has a volume of 250 L at 740 mmHg pressure and 22°C temperature. When the balloon reaches 400 L volume and a pressure of 0.475 atm, it bursts. At which temperature will this balloon burst?

2.13) Explain the following:

1. When you go cycling in summer, do you expect that the pressure in tires to be higher at the beginning or end of cycling?

2. If you check the fully inflated tires of your bicycle after you use the bike for some time, will the air in tires be cold or warm?

3. If gases behaved ideally at different temperature and pressure conditions, matter wouldn't have solid or liquid forms. Why?

4. At the same temperature, average diffusion rates of carbon monoxide and nitrogen are practically the same. Why?

2.14) In a can, there is a mixture of some gases at 20°C and 4.5 atm pressure. When this can was left on sand on a hot day, pressure of gases inside the can rose to 4.8 atm. What is the temperature of the sand?

2.15) What would be the final pressure if the temperature of a sample of oxygen gas was increased from 21°C to 68°C under 0.97 atm constant pressure?

2.16) What is the volume of a gas which has a volume of 94 mL at 25°C and 0.6 atm pressure when it is at 66°C temperature and 0.85 atm pressure?

2.17) A gas mixture in a container consists of CO_2 with a partial pressure of 289 mmHg, O_2 with a partial pressure of 342 mmHg and N_2 with a partial pressure of 122 mmHg. What is the total pressure of gases in this container and mole fraction of each gas?

CHEMICAL EQUATIONS AND CALCULATIONS

CHAPTER-3

ACHIEVEMENTS

After studying this chapter, students will be able to.

*Learn chemical equation concept.

*Write a chemical equation and add symbols showing properties of substances and reaction conditions to the equation.

*Know information given by a balanced equation.

*Calculate number of moles of substances from a reaction equation.

*Calculate masses of substances from a reaction equation.

*Calculate volumes of gases from a reaction equation.

*Identify limiting reagent which determines product and substance formed in bigger amount.

*Calculate percent yields.

3-1-PREFACE

Chemical calculations have a great significance in our lives. Through calculations, ratios of reactants and products are determined. Also knowing reaction ratios of different reactants enables us to calculate amounts of products. Besides, we can calculate how much substance is needed for a reaction with another substance.

Chemists use balanced chemical equations while calculating amounts of reactants and products. In this chapter, we will mention how we can make use of balanced chemical equations to calculate correct amounts of reactants and products.

3-2 –CHEMICAL EQUATION AND CONCEPTS

A chemical equation is a short way to show a chemical reaction using chemical symbols and formulas. Table 3-1 shows additional information and symbols used in writing chemical equations.

Symbol	Use
$\longrightarrow$	*separate between reactants and products
(s)	*indication for solid/abbreviation of solid
(l)	*indication for liquid/ abbreviation of liquid
(g)	*indication for gas/ abbreviation of gas
(aq)	*indication for aqueous solution/ abbreviation of aqueous
$\xrightarrow{\Delta}$ $\xrightarrow{heat}$	*shows reactants are heated
$\xrightarrow{pt}$	*tells that platinum is used as catalyst. Catalyst can be written under the arrow.

3-3-INFORMATION GIVEN BY BALANCED EQUATIONS

A lot of information can be obtained from a balanced chemical equation. For example, we can study the formation reaction of ammonia gas.

$$N_{2(g)} \quad + \quad 3H_{2(g)} \quad \rightarrow \quad 2NH_{3(g)}$$

We can get the information from this equation given in Table 3-2.

Table 3-2 Information given by chemical equations

1.Properties of reactants and products	Nitrogen gas $N_2(g)$	Hydrogen gas $3H_2(g)$	Ammonia gas $2NH_3(g)$
2.Relative numbers of molecules	1 molecule	3 molecules	2 molecules
3.Number of moles	1 mole	3 moles	2 moles
4.Ratio between masses of substances	28 g	6 g	34 g
5.Ratios between volumes of gases if measured under the same conditions	1 volume	3 volumes	2 volumes
Ratios between volumes of gases if measured under standard conditions (STP)	22.4 L	67.2 L	44.8 L

Explanation of titles given in Table 3-2

1) Determining properties of reactants and products

Gas state of matter is shown with (g) symbol. In other words, the equation shows the reaction between nitrogen and hydrogen gases to form ammonia gas.

2) Determining relative numbers of molecules

A molecular formula shows one single molecule of a substance. In other words, the reaction occurs between 1 molecule of nitrogen (N_2) and 3 molecules of hydrogen (H_2) to form 2 molecules of ammonia (NH_3).

Ratio between numbers of molecules of N_2 and H_2 is 1/3

Ratio between numbers of molecules of N_2 and NH_3 is 1/2

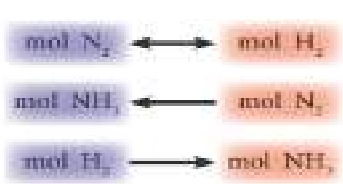
Ratio between numbers of molecules of H_2 and NH_3 is 3/2

Similarly, other ratios can also be written.

3) Determining relative number of moles

Ratio between numbers of molecules is equal to ratio between numbers of moles.

According to this:



Ratio between numbers of moles of N_2 and H_2 1:3 = $\frac{1 \text{ mol } N_2}{3 \text{ mol} H_2}$

Ratio between numbers of moles of NH_3 and N_2 2:3 = $\frac{2 \text{ mol } NH_3}{1 \text{mol } N_2}$

Ratio between numbers of moles of H_2 and NH_3 3:2 = $\frac{3 \text{ mol } H_2}{2 \text{ mol } NH_3}$

Similarly, other ratios can also be written.

Using moles instead of molecules can be explained this way:

If we study the reaction, notes that 1 molecule of nitrogen (N_2) and 3 molecules of hydrogen (H_2) react to form 2 molecules of ammonia (NH_3).

Multiply two sides of the equation with Avogadro's number.

Number of molecules of any substance as many as Avogadro's number is equivalent to 1 mole of that substance.

1 mole of (N_2) reacts with 3 moles of (H_2) to form 2 moles of (NH_3).

4) Determining ratio between masses of substances

To calculate masses of substances in equations, we need to know number of moles and molar masses. Thus, we can calculate mass via the following formula.

$m\ (g) = n\ (mol) \times M\ (g/mol)$

$n\ (mol)$ = number of moles

$m\ (g)$ = mass in gram

$M(g/mol)$= molar mass

$M\ (N_2) = 2 \times 14 = 28$ g/mol

Calculation of mass of N_2 in the reaction equation: We check the molar mass of N from the table of atomic masses.

Then we use the following equation to calculate the mass of N_2 in the reaction

$$m\ (g) = n\ (mol) \times M\ (g/mol)$$

$$m\ (g) = 1\ mol \times 28 \frac{g}{mol} = 28\ g\ N_2$$

To calculate mass of H_2 in the reaction equation, we follow the same method as in calculation of N_2. Firstly, we check molar mass of H from the chart of atomic masses.

$M(H_2) = 2 \times 1 = 2$ g/mol

$$m\ (g) = 3\ mol \times 2 \frac{g}{mol} = 6\ g\ H_2$$

Then we calculate the mass of H_2 in the reaction equation.

By using the same method, we calculate the mass of NH_3. For this, first we calculate molecular mass of NH_3.

$M\ (NH_3) = (1 \times 14) + (3 \times 1) = 17$ g/mol

Then we calculate the mass of NH_3 in the reaction equation.

$$m\ (g) = 2\ mol \times 17 \frac{g}{mol} = 34\ g\ NH_3$$

Ratio of NH_3 to N_2.

$$\frac{34\ g\ NH_3}{28\ g\ N_2}$$

Exercise 3-1

Calculate the masses of the following in grams.

A- 2 moles of H_2O

B- 10 moles of H_2SO_4

Exercise 3-2

What are the volumes of samples of 3 moles of CO_2 and 2 moles of CO_2 under standard conditions (STP)?

And total masses of reactants,

$28 \text{ g } N_2 + 6 \text{ g } H_2 = 34 \text{ g}$

Thus, total mass of products must be 34 g (NH_3).

Therefore, total mass of reactants is equal to total mass of products. This is consistent with law of mass conservation.

5) Determining ratio between volumes of gases

We can express volume of gas in litre (L), milliliter (mL), and cubic centimeter (cm^3).

Ratio of volume of H_2 to volume of $N_2 = \frac{3L\ H_2}{1L\ N_2}$

Thus, other ratios are the same.

As we have studied in Chapter 1&2, 1 mole of any gas occupies 22.4 L (22.400 mL) of volume under standard conditions. Therefore, volume of a gas under standard conditions can be calculated through the following equation.

$$V(L) = n\ (\text{mol}) \times 22.4\ (\frac{L}{\text{mol}})$$

Calculate the volume of N_2.

$$V_{N_2} = 1\ (\text{mol}) \times 22.4\ (\frac{L}{\text{mol}}) = 22.4\ L$$

Calculate the volume of H_2.

$$V_{H_2} = 3\ (\text{mol}) \times 22.4\ (\frac{L}{\text{mol}}) = 67.2\ L$$

Calculate the volume of NH_3.

$$V_{NH_3} = 2\ (\text{mol}) \times 22.4\ (\frac{L}{\text{mol}}) = 44.8\ L$$

Total volume of reactant gases

$$22.4\ L\ N_2 + 67.2\ L\ H_2 = 89.6\ L$$

The volume of product gas (NH_3) is 44.8 L. As densities of gases are different, volumes of reactants and products need not to be equal.

3-4-CHEMICAL CALCULATIONS

3-4-1-Calculation of number of moles

In a reaction equation, we can calculate unknown number of moles of any of reactants or products in a balanced equation through known no. of moles of a substance or from ratio of no. of moles of substances.

$$\frac{\text{Number of moles of unknown substance}}{\text{Number of moles of known substance}}$$

Example 3-1

For the following reaction:

$$2Na_{(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)}$$

Calculate those;

A- Number of moles of H_2 produced by the reaction of 0.145 mole of Na

B- Number of moles of H_2O necessary to produce 0.75 mole of NaOH

Sodium reacts with water vigorously and produces hydrogen gas which burns instantly

Solution:

Known: 0.75 mole of NaOH and 0.145 moles of Na

Unknown: mole of H_2O and mole of H_2

A) Conversion factor

$$\text{mol Na} \longrightarrow \text{mol } H_2$$

$$\frac{1 \text{ mol } H_2}{2 \text{ mol Na}}$$

Number of moles of H_2 produced by the reaction of 0.145 mole of Na:

$$= 0.145 \text{ mol Na} \times \frac{1 \text{ mol } H_2}{2 \text{ mol Na}} = 0.072 \text{ mol } H_2$$

B) Conversion factor

$$\text{mol } H_2O \longrightarrow \text{mol NaOH}$$

$$\frac{2 \text{ mol } H_2O}{2 \text{ mol NaOH}}$$

Number of moles of H_2O necessary to produce 0.75 moles of NaOH

$$= 0.75 \text{ mol NaOH} \times \frac{2 \text{ mol } H_2O}{2 \text{ mol NaOH}} = 0.75 \text{ mol } H_2O$$

Example 3-2

Sodium chloride (NaCl) is formed from the reaction of sodium metal and chlorine gas as in the following equation.

$$2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$

If 3.4 moles of Cl_2 reacts with enough sodium, what is the no. of moles of produced NaCl?

Solution:

Known: 3.4 moles of Cl_2

Unknown: Mole of NaCl

$$\text{mol } Cl_2 \longrightarrow \text{mol } NaCl$$

Factor: $\dfrac{2 \text{ mol } NaCl}{1 \text{ mol } Cl_2}$

Number of moles of NaCl is as follows:

$$= 3.4 \text{ mol } Cl_2 \times \frac{2 \text{ mol } NaCl}{1 \text{ mol } Cl_2} = 6.8 \text{ mol } NaCl$$

Example 3-3

How many moles of Na_2O can be obtained from the reaction of 4.8 moles of sodium according to the following reaction?

$$4Na_{(s)} + O_{2\,(g)} \longrightarrow 2Na_2O_{(s)}$$

Solution:

Known: 4.8 moles of Na

Unknown: mole of Na_2O

$$\text{mol } Na \longrightarrow \text{mol } Na_2O$$

Factor: $\dfrac{2 \text{ mol } Na_2O}{4 \text{ mol } Na}$

Number of moles of Na_2O produced from the reaction of 4.8 moles of Na

$$= 4.8 \text{ mol } Na \times \frac{2 \text{ mol } Na_2O}{4 \text{ mol } Na} = 2.4 \text{ mol } Na_2O$$

Exercise 3-3

$$4Al_{(s)} + 3O_{2\,(g)} \longrightarrow 2Al_2O_{3\,(s)}$$

The equation above shows the oxidation of aluminum in air and formation of a layer on aluminum which protects it from continuous oxidation.

A- Write down 3 relationships expressing ratio between moles of two substances in the equation.

B- Calculate number of moles of Al necessary to form 6 moles of Al_2O_3.

C-Calculate number of moles of O_2 which reacts to form 12 moles of Al_2O_3.

3-4-2-Calculation of Masses of Substances

This is applied in 3 steps:

Step 1:

Number of moles of the substance with known mass is calculated. If this substance is A, firstly we calculate molar mass from periodic table, then we use the following equation to calculate number of moles.

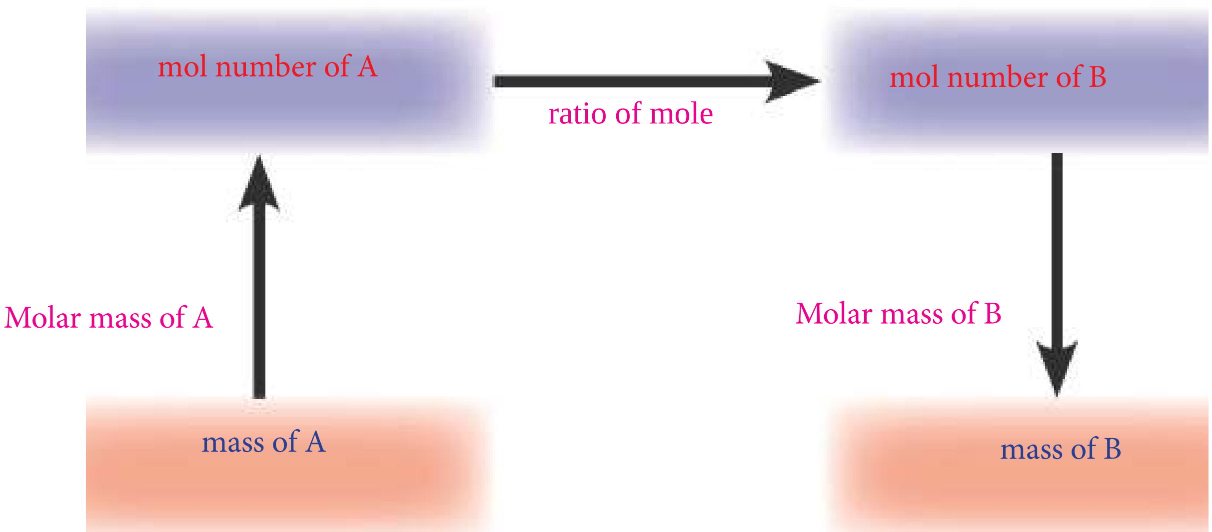
$$n\,(\text{mol}) = \frac{m\,(\text{g})}{M\,(\text{g/mol})}$$

Step 2:

Calculate unknown no. of moles of the other substance. If we assume this substance as B, using no. of moles of A, we use the following equation.

Mol number of substance = Mol number of known substance *x* mol ratio

Or the following diagram is followed.



Step 3:

We start from calculation of molar mass to find out unknown mass of substance B. Firstly, by using periodic table as in Step 2, we apply mole calculation in the following equation.

m (g) = n (mol) × M (g/mol)

Note:

1) We can directly start from Step 2 if number of moles of substance is known instead of mass.

2) If number of moles is unknown in the equation instead of mass, we can avoid Step 3.

Example 3-4

Calculate the mass of released CO_2 in grams from burning of 500g of C_8H_{18} according to the following equation.

$$2\,C_8H_{18\,(g)} + 25\,O_{2\,(g)} \xrightarrow{\Delta} 16\,CO_{2\,(g)} + 18\,H_2O_{(g)}$$

Solution:

Mass of C_8H_{18}=500g

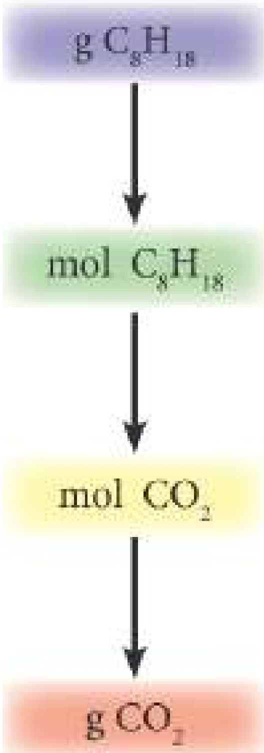
Mole ratio = $\dfrac{16 \text{ mol } CO_2}{2 \text{ mol } C_8H_{18}}$

Unknown: Mass of CO_2

Step 1: Firstly, we calculate molar mass of C_8H_{18} from periodic table.

$M(C_8H_{18}) = (8 \times 12) + (18 \times 1) = 114$ g/mol

Then we apply the following formula to calculate number of moles of C_8H_{18}.



$$n\,(mol) = \frac{m\,(g)}{M\,(g/mol)}$$

$$n\,(mol) = \frac{500\,g}{114\,g/mol} = 4.39\ mol\ C_8H_{18}$$

Step 2: We apply the following formula to calculate number of moles of released CO_2 from the reaction of 4.39 moles of C_8H_{18}.

mol number of CO_2 = mol number of C_8H_{18} *x* ratio of number moles both substance

$$n\,(mol) = 4.39\ mol\ C_8H_{18} \times \frac{16\ mol\ CO_2}{2\ mol\ C_8H_{18}} = 35.12\ mol\ CO_2$$

Step 3: Firstly, we calculate molar mass of CO_2 from periodic table.
$M\,(CO_2) = (1\times12) + (2\times16) = 44$ g/mol
Then we calculate mass of CO_2 by applying the following formula.

$m(g) = n\,(mol) \times M(g/mol)$

$$m\,(g) = 35.12\ mol\ CO_2 \times 44\frac{g}{mol\ CO_2} = 1545\ g\ CO_2$$

Exercise 3-4

CS_2 burns with oxygen as in the following equation:

$$CS_{2(l)} + 3O_{2(g)} \longrightarrow CO_{2(g)} + 2SO_{2(g)}$$

What is the number of moles of each product produced from the reaction of 48.0 g of O_2?

Example 3-5

One of the components of acid rain is nitric acid which is formed from the reaction of NO_2 with oxygen and rainwater according to the following equation.

$$4\,NO_{2\,(g)} + O_{2\,(g)} + 2\,H_2O_{(l)} \longrightarrow 4\,HNO_{3\,(aq)}$$

What is the amount of HNO_3 produced from the reaction of 1500g of NO_2 with enough water and oxygen?

Solution:

Known: Mass of NO_2 1500g

Mole ratio $= \dfrac{4\ mol\ HNO_3}{4\ mol\ NO_2}$

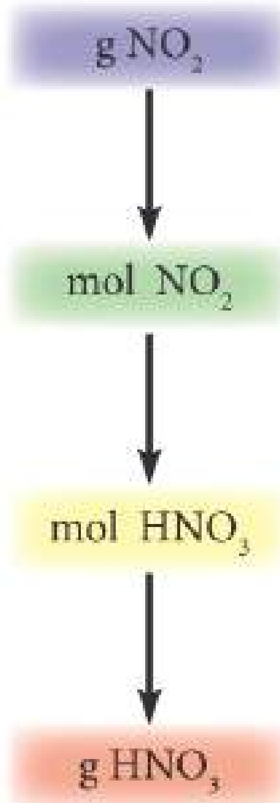
Unknown: Mass of HNO_3.

Step 1: Firstly, we calculate mass of NO_2 from periodic table.

$M\,(NO_2) = (1 \times 14) + (2 \times 16) = 46$ g/mol. Then, to calculate number of moles of NO_2, we use the following equation.

$$n\ (mol) = \frac{m\ (g)}{M\ (g/mol)}$$

$$n\ (mol) = \frac{1500\ g}{46\ \ g/mol} = \ 32.6\ \ mol\ NO_2$$



Step 2: We use the following equation to calculate number of moles of HNO_3 which reacts with 32.6 moles of NO_2.

Mole number of HNO_3= mole number of NO_2 × ratio of number of moles of both substances

$$n\ (mol) = 32.6\ \ mol\ NO_2 \times \frac{4\ \ mol\ \ HNO_3}{4\ \ mol\ \ NO_2} = 32.6\ \ mol\ HN0_3$$

Step 3: We calculate mass of HNO_3 from its molar mass (63 g/mol) and number of moles as follows.

$$m\ \ (g) = 32.6\ mol\ \ HNO_3 \times 63\ \frac{g}{mol\ \ HNO_3} = 2054\ g\ \ HNO_3$$

Example 3-6

Calculate number of moles of released O_2 from heating of 1.65 g of $KClO_3$ according to the following equation.

$$2KClO_{3(s)} \xrightarrow{\Delta} 2KCl_{(s)} + \ 3O_{2(g)}$$

Solution:

1-Known: Mass of $KClO_3$:1.65g

$$\text{2-Mole ratio} = \frac{3\ \ mol\ O_2}{2\ \ mol\ KClO_3}$$

Unknown: Number of moles of O_2.

Step 1: Firstly, we calculate molar mass of $KClO_3$ from periodic table.

M ($KClO_3$) = (1 × 39) + (1 × 35.5) + (3 × 16) = 122.5 g/mol

Then we use the following formula to calculate number of moles of $KClO_3$.

$$n\ (mol) = \frac{m\ (g)}{M\ (g/mol)}$$

$$n\ (mol) = \frac{1.65\ (g)}{122.50\ \ (g/mol)} = 0.013\ mol\ KClO_3$$

Step 2: We use the following formula to calculate number of moles of O_2.

Exercise 3-5

Phosphorus is prepared synthetically from calcium phosphate, silicon dioxide and coal in electric oven as in the following equation.

$$2Ca_3(PO_4)_2 + 6SiO_2 + 10C \xrightarrow{\Delta} 6CaSiO_3 + 10CO + P_4$$

Make the following calculations:

A-What is the mass of P_4 in grams produced from the reaction of 1.0 mole of $Ca_3(PO_4)_2$?

B- Calculate number of moles of P_4 produced from the reaction of 62.0 g of $Ca_3(PO_4)_2$.

Exercise 3-6

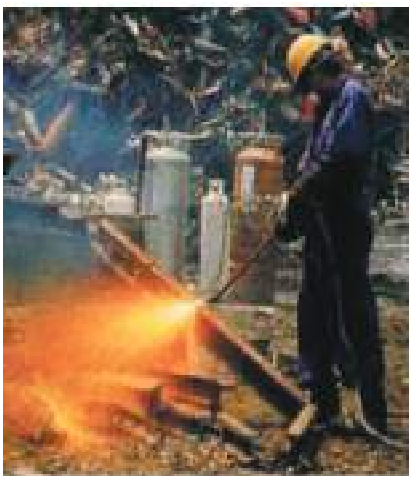
Acetylene gas is prepared through the following equation by adding water to calcium carbide (CaC_2).

$$CaC_{2(s)} + 2H_2O_{(l)} \longrightarrow C_2H_{2(g)} + Ca(OH)_{2(s)}$$

Calculate the following:

A-What is the mass of released acetylene from the reaction of 32 g of CaC_2?

B- What is number of moles of CaC_2 which reacts with 36 g of H_2O?



Acetylene is used as a raw material in the production of rubber, plastics and acetic acid and produce the oxyacetylene for cutting or welding metals.

Number of moles of O_2= Number of moles of $KClO_3$ *x* ratio of number of moles

$$n\ (mol) = 0.013\ mol\ KClO_3 \times \frac{3mol\ O_2}{2\ mol\ KClO_3} = 0.02\ mol\ O_2$$

3-4-3-The Specific reactant for product

As we have learnt before, a balanced equation gives ratio between number of moles of reactants and products as in the following equation.

$$N_{2\ (g)} + 3H_{2\ (g)} \longrightarrow 2\ NH_{3\ (g)}$$

In order to get 2 moles of NH_3, a reaction occurs between 1 mole of N_2 and 3 moles of H_2. According to this, when 1 mole of N_2 and 3 moles of H_2 are mixed, both substances react completely, because ratios of no. of moles are consistent with the ratios in the balanced equation. At the end of the reaction, 2 moles of NH_3 is formed.

But if we use more N_2 in the reaction, for example when we use 2 moles of N_2 and 3 moles of H_2, only 1 mole of N_2 and 3 moles of H_2 reacts and 1 mole of N_2 (substance in excess) doesn't react. Therefore, N_2 is in excess and doesn't participate in the reaction. H_2 is called as specific reactant. The reason is that all of this substance participates fully in the reaction. Number of moles of limiting reagent component determines number of moles of products.

A) We determine specific reactant through the following methods:

We calculate ratio of shown substance in the reaction and number of moles of all reactants.

B) According to balanced chemical reaction, each of these ratios is multiplied with no. of moles of each substance.

C) According to the result, the substance giving the smallest value is the specific reactant which determines the product.

If we apply those steps above to the previous reaction:

2 moles of N_2 and 3 moles of H_2,

A- Ratio of number of moles of NH_3 to number of moles of N_2 according to reaction equation $\frac{2\ mol\ NH_3}{1\ mol\ N_2}$

Ratio of number of moles of NH_3 to number of moles of H_2 according to reaction equation $\frac{2\ mol\ NH_3}{3\ mol\ H_2}$

B- We multiply ratios with number of moles of N_2 and H_2.

$$2\ mol\ N_2 \times \frac{2\ mol\ NH_3}{1\ mol\ N_2} = 4\ mol\ NH_3$$

$$3\ mol\ H_2 \times \frac{2\ mol\ NH_3}{3\ mol\ H_2} = 2\ mol\ NH_3$$

C- As number of moles of NH_3 formed from the reaction of H_2 is less than that of from the reaction of N_2, specific reactant of ammonia product is **hydrogen**.

Example 3-7

According to the following reaction,

$$Ti_{(s)} + 2Cl_{2(g)} \longrightarrow TiCl_{4(s)}$$

When 1.8 moles of titanium (Ti) and 3.2 moles of chlorine (Cl_2) are mixed, which is the specific reactant that determines the product?

Solution:

Unknown: Specific reactant which determines the product

The ratio of number of moles of $TiCl_4$ to number of moles of Ti in the reaction

$$\frac{1 \text{ mol } TiCl_4}{1 \text{ mol } Ti}$$

The ratio of number of moles of $TiCl_4$ to number of moles of Cl_2 in the reaction

$$\frac{1 \text{ mol } TiCl_4}{2 \text{ mol } Cl_2}$$

Number of moles of $TiCl_4$ from the reaction of Ti

$$1.8 \text{ mol } Ti \times \frac{1 \text{ mol } TiCl_4}{1 \text{ mol } Ti} = 1.8 \text{ mol } TiCl_4$$

Number of moles of $TiCl_4$ from the reaction of Cl_2.

$$3.2 \text{ mol } Cl_2 \times \frac{1 \text{ mol } TiCl_4}{2 \text{ mol } Cl_2} = 1.6 \text{ mol } TiCl_4$$

As number of moles of $TiCl_4$ formed from the reaction of 3.2 moles of Cl_2 is less than that of from the reaction of 1.8 moles of Ti, limiting reagent that determines the product is Cl_2.

Example 3-8

Sodium chloride (NaCl) is obtained through the reaction of sodium (Na) and chlorine (Cl_2) according to the equation below:

$$2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$

Which is the specific reactant in the reaction of 11.2 moles of Na and 3.2 moles of Cl_2?
How many number of moles of NaCl is formed?

Exercise 3-7

Magnesium nitrite (Mg_3N_2) is obtained from the reaction magnesium and nitrogen according to the following equation.

$$3\ Mg_{(s)} + N_{2(g)} \xrightarrow{\Delta} Mg_3N_{2(s)}$$

When the reaction occurs at a certain temperature by mixing 4.0 moles of N_2 and 6.0 moles of Mg, which mixture is present in the reaction vessel that fits the following answers?

A) 4.0 moles of Mg_3N_2 and 1.0 mole of Mg in excess
B) 2.0 moles of Mg_3N_2 and 2.0 moles of N_2 in excess
C) 6.0 moles of Mg_3N_2 and 3.0 moles of N_2 in excess

$$\text{mol Na} \longrightarrow \text{mol NaCl}$$

$$\text{mol } Cl_2 \longrightarrow \text{mol NaCl}$$

Solution:

Known:

Number of moles of Na is 11.2

Number of moles of Cl_2 is 3.2.

Unknown:

1) Limiting reagent

2) Number of moles of formed NaCl

Ratio between number of moles of NaCl and number of moles of Na

$$\frac{2\ \text{mol NaCl}}{2\ \text{mol Na}}$$

Ratio between number of moles of NaCl and number of moles of Cl_2.

$$\frac{2\ \text{mol NaCl}}{1\ \text{mol Cl}_2}$$

Number of moles of NaCl obtained from 11.2 moles of Na

$$11.2\ \text{mol Na} \times \frac{2\ \text{mol NaCl}}{2\ \text{mol Na}} = 11.2\ \text{mol NaCl}$$

Number of moles of NaCl obtained from 3.2 moles of Cl_2.

$$3.2\ \text{mol Cl}_2 \times \frac{2\ \text{mol NaCl}}{1\ \text{mol Cl}_2} = 6.4\ \text{mol NaCl}$$

As number of moles of NaCl obtained from Cl_2 interaction is less than that of Na interaction, the limiting reagent is Cl_2.
As the limiting reagent is Cl_2, number of moles of formed NaCl is calculated from the interaction of Cl_2 as follows:

$$3.2\ \text{mol Cl}_2 \times \frac{2\ \text{mol NaCl}}{1\ \text{mol Cl}_2} = 6.4\ \text{mol NaCl}$$

3-4-4-Calculation of Gas Volume

The following steps are followed to calculate volumes of gases in chemical reactions:

Step 1: We calculate number of moles of substance with known mass by the following formula:

$$n\ (\text{mol}) = \frac{m\ (g)}{M\ (g/\text{mol})}$$

If the substance is in gas state, starting from its volume given in the question, we apply ideal gas law to calculate no. of moles of the substance.

$$n\,(mol) = \frac{PV}{RT}$$

If the volume of the gas is measured under standard conditions (STP), we use the following formula to calculate number of moles:

$$n\,(mol) = \frac{V\,(L)\,(\text{at STP})}{22.4\,(L/mol)}$$

Step 2: Number of moles of the substance asked in the question is calculated by the method in Step 1.

Step 3: From the number of moles calculated in Step 2, mass of the asked substance is calculated by using the following formula:

$$m\,(g) = n\,(mol) \times M\,(g/mol)$$

To calculate gas volume from number of moles calculated in Step 2, we can apply the following formula.

$$V(L) = \frac{nRT}{P}$$

We apply the following formula to calculate volume of a gas under standard conditions (STP).

$$V\,(L) = n\,(mol) \times 22.4\,(L/mol)$$

Besides, we can calculate unknown gas volume in the question through known volume of a gas in the balanced reaction equation. But volumes must be measured under the same temperature and pressure values.

Example 3-9

According to the following equation, nitrogen monoxide gas changes into brownish nitrogen dioxide gas through reaction with oxygen.

$$2NO_{(g)} + O_{2\,(g)} \longrightarrow 2\,NO_{2\,(g)}$$

Calculate the volume of formed NO_2 from the reaction of 34 L of oxygen with enough NO. Volumes measured under (STP)

Solution:

Known: The volume of O_2 under (STP) is 34 L.

Unknown: The volume of formed NO_2

Exercise 3-8

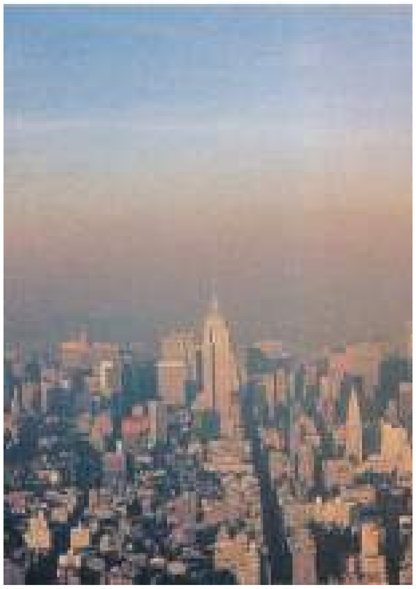
Silicone oxide (quartz) is generally an inactive substance. But as it is seen below, it reacts vigorously with hydrogen fluoride.

$$SiO_{2(s)} + 4HF_{(g)} \longrightarrow SiF_{4(g)} + 2H_2O_{(l)}$$

As 2.0 moles of HF and 4.5 moles of SiO_2 were put in the container,

A) Which is the limiting reagent?

B) How many moles of SiF_4 were formed?



Air pollution from brownish nitrogen dioxide (NO_2) gas poses danger to human life.

The ratio between number of moles of NO_2 and O_2.

$$\frac{2\ \text{mol}\ NO_2}{1\ \text{mol}\ O_2}$$

If measurements were made under standard conditions (STP), volumes will be directly proportional with moles. Thus, ratio between volumes:

$$\frac{2L\ NO_2}{1L\ O_2}$$

Therefore, the volume of NO_2 will be as follows:

$$V(L) = 34\ L\ O_2 \times \frac{2L\ NO_2}{1L\ O_2} = 68\ L\ NO_2$$

Example 3-10

Calculate the volume of obtained O_2 from heating 3.5 moles of KNO_3 according to the following equation under standard conditions (STP).

$$2KNO_{3(s)} \xrightarrow{\Delta} 2KNO_{2(s)} + O_{2(g)}$$

Solution:

Known: 3.5 moles of KNO_3.

Unknown: Volume of obtained O_2 under STP

Ratio of number of moles of O_2 to number of moles of KNO_3 in the reaction:

$$\frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ KNO_3}$$

Number of moles of O_2 released from decomposition of 3.5 moles of KNO_3:

$$n(\text{mol}) = 3.5\ \text{mol}\ KNO_3 \times \frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ KNO_3} = 1.75\ \text{mol}\ O_2$$

Calculate the volume of O_2 under standard conditions by using the following formula:

$$V(L) = n(\text{mol}) \times \frac{22.4\ L}{1\ \text{mol}} = 1.75\ \text{mol} \times \frac{22.4\ L}{1\ \text{mol}} = 39.2\ L$$

Exercise 3-9

Phosphine gas is formed from the reaction of phosphorus and hydrogen gas according to the following equation.

$$P_{4(s)} + 6H_{2(g)} \longrightarrow 4PH_{3(g)}$$

Calculate the volume of PH_3 formed from the reaction of 0.42 L of H_2.

Example 3-11

Calculate number of moles of copper atoms formed from the reaction of 4250 mL of H_2 with enough CuO under standard conditions (STP).

$$CuO_{(s)} + H_{2(g)} \xrightarrow{\Delta} Cu_{(s)} + H_2O_{(l)}$$

Solution:

Known: Volume of H_2 under standard conditions is 4250 mL.

Unknown: Number of moles of Cu

1- We convert volume of H_2 from mL to L.

$$V(L) = 4250\ mL\ H2 \times \frac{1L}{1000\ mL} = 4.250\ L\ H_2$$

2- We calculate number of moles of H_2 through the following formula under normal conditions:

$$n(mol) = \frac{V(L)\ \text{at STP}}{22.4\ (L/mol)}$$

$$n(mol) = \frac{4.250\ L}{22.4\ (L/mol)} = 0.19\ mol\ H_2$$

3- We calculate number of moles of Cu formed from the reaction of 0.19 moles of H_2 through ratio of number of moles.

Number of moles = Known number of moles *x* ratio of number of moles

$$n(mol) = 0.19\ mol\ H_2 \times \frac{1\ mol\ Cu}{1\ mol\ H_2} = 0.19\ mol\ Cu$$

Example 3-12

0.4 mole of potassium chlorate decomposes with heat according to the following equation.

$$2KClO_{3\,(s)} \xrightarrow{\Delta} 2KCl_{(s)} + 3O_{2(g)}$$

Calculate the volume of O_2 released at 27 °C temperature and 760 Torr pressure.

Solution:

Known: Number of moles of $KClO_3$ is 0.4.

Unknown: volume of O_2 released at 27 °C temperature 760 Torr pressure

Hydrazine is used as rocket fuel

Exercise 3-10

Hydrazine (N_2H_4) which is used as rocket propellant burns according to the following equation.

$$N_2H_{4(l)} + O_{2(g)} \rightarrow N_{2(g)} + 2H_2O_{(g)}$$

Calculate the volume of formed N_2 through reaction of 3.2 kg of N_2H_4 with enough O_2 under standard conditions (STP).

Exercise 3-11

Ammonium nitrate (NH_4NO_3) decomposes through high heat according to the following equation:

$$2NH_4NO_{3(s)} \rightarrow 2N_{2(g)} + 4H_2O_{(g)} + O_{2(g)}$$

Calculate total volume of gases released from decomposition of 40 grams of NH_4NO_3 under standard conditions (STP).

Exercise 3-12

$$Mg_3N_{2(s)} + 6H_2O_{(l)} \longrightarrow 3Mg(OH)_{2(s)} + 2NH_{3(g)}$$

Make the following calculations according to the equation above.

A-What is the mass of magnesium nitride (Mg_3N_2) necessary to produce 22.4 L of ammonia under standard conditions (STP)?

B-What is the number of moles of formed $Mg(OH)_2$?

1- We convert 27 °C to Kelvin.

T (K) = 27 + 273 = 300 K

2- We calculate number of moles of O_2 released from 0.4 mole of decomposition of $KClO_3$ and mole ratio.

$$n(mol) = 0.4 \text{ mol } KClO_3 \times \frac{3 \text{ mol } O_2}{2 \text{ mol } KClO_3} = 0.6 \text{ mol } O_2$$

3- By applying general gas law, we calculate the volume of O_2 at 300 K temperature and 1 atm pressure (1 atm = 760 Torr).

$$V = \frac{nRT}{p}$$

$$V(L) = \frac{0.6 \text{ mol} \times 0.082 \frac{L.atm}{K.mol} \times 300 \text{ K}}{1 \text{ atm}} = 14.76 \text{ L } O_2$$

3-4-5-Percent Yield

Theoretical yield is obtained through calculation of masses of products and reactants from a balanced chemical equation. Actual yield is obtained from experimental results by the same components.

Actual yield is always less than theoretical yield. The reasons are given below:

1) Interaction between reactants cannot be completed.

2) Usage of impure substances. As a result, unwanted products are formed through side reactions.

3) Loss of some product while filtering or transferring to another container.

4) Lack of attention during weighing reactants and products.

$$\text{percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

Example 3-13

For the following reaction;

$$2AL(s) + 3CI(g) \longrightarrow 2AlCl_3(s)$$

Calculate the percentage yield of aluminum chloride obtained from the reaction of 1.5 moles of Al. Mass of actual yield is 139 grams.

Solution:

Percent yield is calculated through the following formula.

$$\text{percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

Known: 1.5 moles of Al

Unknown: Percent yield of $AlCl_3$.

For percent yield, we calculate number of moles of $AlCl_3$ using ratio of mole numbers from the balanced equation.

$$n(mol) = 1.5 \text{ mol Al} \times \frac{2 \text{ mol AlCl}_3}{2 \text{ mol Al}} = 1.5 \text{ mol AlCl}_3$$

We calculate mass of formed $AlCl_3$ from molar mass of $AlCl_3$ and known number of moles theoretically.

$M(AlCl_3) = 133.5$ g/mol $AlCl_3$

$$m(g) = n(mol) \times M(g/mol) = 1.5 \text{ mol} \times 133.5 \text{ g/mol} = 200.3 \text{ g}$$

Thus, percent yield:

$$\% AlCl_3 = \frac{139 \text{ g}}{200.3 \text{ g}} \times 100\% = 69.4\%$$

Example 3-14

$CaCO_3$ decomposes with heat according to the following equation.

$$CaCO_{3\,(s)} \xrightarrow{\Delta} CaO_{(s)} + CO_{2\,(g)}$$

1- What is the mass of CaO obtained theoretically through heating of 24.8 g of $CaCO_3$?

2- As molar mass of $CaCO_3$ is 100 g/mol, molar mass of CaO is 56 g/mol and actual yield is 13.1 g, calculate percent yield of CaO product.

Solution:

Known:

Mass of $CaCO_3$= 24.8 g

Actual mass of CaO = 13.1 g

molar mass of $CaCO_3$ is 100 g/mol, molar mass of CaO is 56 g/mol

Unknown:

1) Theoretical yield of CaO

2) Percent yield of CaO

1- Calculate number of moles of $CaCO_3$.

$$n(mol) = \frac{24.8 \text{ (g)}}{100 \text{ (g/mol)}} = 0.25 \text{ mol}$$

2- Calculate no. of moles of CaO produced from 0.25 mole of $CaCO_3$ using mole ratio in the balanced equation.

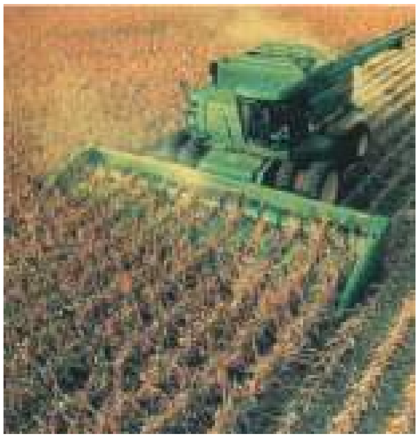
$$0.25 \text{ mol CaCO}_3 \times \frac{1 \text{mol CaO}}{1 \text{mol CaCO}_3} = 0.25 \text{ mol CaO}$$

3- Calculate mass of 0.25 mole of CaO (theoretical product).

$$m(g) = 0.25 \text{ (mol)} \times 56 \text{ (g/mol)} = 14 \text{ g CaO}$$

4- Calculate percent yield using the following formula to produce CaO.

$$\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\% \Rightarrow \% CaO = \frac{13.1\text{g}}{14.0\text{g}} \times 100\% = 93.6\%$$



Production in field is calculated through percent yield .The reason is that production changes from year to year, actual yields is different from expected (theoretical) yield

Exercise 3-13

56 g of cadmium and dilute hydrochloric acid react according to the following equation.

$$Cd_{(s)} + 2HCl_{(aq)} \longrightarrow CdCl_{2(s)} + H_{2(g)}$$

1- Calculate mass of produced hydrogen.

2- As actual yield of hydrogen is 0.75g, calculate percent yield.

Exercise 3-14

56 g of iron reacts with 4 moles of dilute hydrochloric acid according to the following equation.

$$Fe_{(s)} + 2HCl_{(aq)} \longrightarrow FeCl_{2\,(s)} + H_{2\,(g)}$$

1- Calculate mass of produced hydrogen.

2- As actual yield of hydrogen is 1.75 g, calculate percent yield.

3- Which one is the limiting reagent?

BASIC CONCEPTS

***Chemical Equation**: It is a way of showing a chemical reaction via symbols and numbers of atoms.

The following formula is used to calculate number of moles from mass of substance and its molar mass.

$$n\ (mol) = \frac{m\ (g)}{M\ (g/mol)}$$

*The following formula is used to calculate mass of substance from its number of moles and its molar mass.

$$m\ (g) = n\ (mol) \times M\ (g/mol)$$

*The following formula is used to calculate number of moles of a gas from general gas law.

$$n\ (mol) = \frac{P\ (atm) \times V\ (L)}{R\ (atm.\ L/mol.K) \times T\ (K)}$$

*The following formula is used to calculate number of moles of a gas from gas volume measured under standard conditions.

$$n\ (mol) = \frac{V\ (L)\ \text{at}\ (STP)}{22.4\ (L/mol)}$$

*The following formula is used to calculate volume of a gas from its number of moles by applying general gas law.

$$V\ (L) = \frac{n\ (mol) \times R\ (atm.\ L/mol.K) \times T\ (K)}{P\ (atm)}$$

-The following equation is used to calculate volume of a gas from its number of moles under standard conditions (STP).

$$V\ (L) = n\ (mol) \times 22.4\ (L/mol)$$

*In an equation, from a known number of moles of a substance, we can calculate unknown number of moles of another substance as follows:

Unknown number of moles is found through ratio of number of moles of both substances in the balanced reaction equation.

***Specific Reactant for product** which determines products is the substance which is fully used up in a reaction. Number of moles of reactants determine number of moles of products. At the same time, it is the substance which has a lower number of moles than the product substance.

***Theoretical Yield** : Mass of a substance which is calculated from a known mass of a substance which participates in the balanced reaction.

***Actual Yield**: Mass of a substance which can be measured through an experiment. Actual yield is always less than theoretical yield.

QUESTIONS OF CHAPTER 3

Reminder: You may need to know atomic numbers for some questions. You may refer to the Table at the end of the book.

3.1) 1.26 moles of copper reacts with 0.8 mole of sulfur and forms copper sulfide compound according to the following equation. Thus,

$$2Cu_{(s)} + S_{(s)} \longrightarrow Cu_2S(s)$$

A- Which is the spesific reactant which determines the product?

B- What is the number of moles of the substance in excess?

3.2) Iron is obtained from reduction of iron (III) oxide (Fe_2O_3) with carbon monoxide (CO) according to the following equation.

$$Fe_2O_{3\ (s)} + 3CO_{(g)} \longrightarrow 2Fe_{(s)} + 3CO_{2(g)}$$

A- What is the highest amount of iron which can be obtained from reduction of 400 g of iron (III) oxide?

B -What is the mass of CO necessary for the reduction process?

C- As actual yield is 265 g, what is the percent yield of iron production?

3.3) 60 g silicon dioxide (SiO_2) reacts with enough carbon (C) according to the following equation.

$$SiO_{2(s)} + 3C_{(s)} \longrightarrow SiC_{(s)} + 2CO_{(g)}$$

Calculate the following.

A- What is the mass of formed silicon carbide (SiC)?

B- What is the mass of formed carbon monoxide (CO)?

3.4) Hydrogen gas is formed through the reaction of magnesium with dilute hydrochloric acid according to the following equation.

$$Mg_{(s)} + 2HCl_{(aq)} \longrightarrow M_gCl_{2(aq)} + H_{2(g)}$$

A- Which is the spesific reactant for product when 5.84 g of HCl and 4.8 g of Mg react?

B- What is the number of moles of formed $MgCl_2$?

C- What is the volume of formed H_2 under standard conditions (STP)?

D- Calculate the mass of formed $MgCl_2$.

E- What is the volume of H_2 formed at 27°C and 2.5 atm pressure?

3.5) Oxyacetylene flame is produced through the reaction of acetylene with oxygen according to the following equation.

$$2C_2H_{2(g)} + 5O_{2(g)} \xrightarrow{\Delta} 4CO_{2(g)} + 2H_2O_{(g)}$$

Calculate the following.

A- What is the volume of O_2 which reacts with 52 g of acetylene at STP?

B- What is the number of molecules of CO_2 formed at the end of the reaction?

3.6) For the following reaction:

$$P_{4(s)} + 6F_{2(g)} \rightarrow 4PF_{3(g)}$$

Which of the following is the mass of enough F_2 that reacts with 6.2 g of P_4?

A) 2.85 g B) 5.70 g C) 11.4 g D) 37.2 g

3.7) Methyl alcohol is produced industrially from carbon monoxide gas and hydrogen gas under high pressure and with chromium (III) oxide and zinc oxide catalysts according to the following equation.

$$CO_{(g)} + 2H_{2(g)} \longrightarrow CH_3OH_{(g)}$$

A) Calculate the volume of produced CH_3OH from the reaction of 60.0 L of CO and 80.0 L of H_2.

B) What is the volume of CO or H_2 gas in excess?

3.8) For the following reaction;

$$3Cl_{2(g)} + 6KOH_{(aq)} \longrightarrow 5KCl_{(aq)} + KClO_{3(aq)} + 3H_2O_{(1)}$$

Calculate the number of moles of $KClO_3$ produced from the reaction of 6.72 L of chlorine (Cl_2) under standard conditions (STP).

3.9) Write down 3 equations expressing the ratios between number of moles of substances in the following equation.

$$2Fe_{(s)} + 3S_{(s)} \longrightarrow Fe_2S_{3\,(s)}$$

A) Write down mole ratios between reactants and product in the following reaction.

$$2Na_{(s)} + Cl_{2\,(g)} \longrightarrow 2NaCl_{(s)}$$

B) Correct the mistakes in the following terms.

2 g of Na reacts with 1 g of Cl_2 to produce 2 g of NaCl.

3.10) The reaction between solid lead (Pb) and silver nitrate ($AgNO_3$) is as follows:

$$Pb_{(s)} + 2AgNO_{3\,(aq)} \longrightarrow Pb(NO_3)_{2\,(aq)} + 2Ag_{(s)}$$

A- Calculate the number of moles of silver nitrate necessary for 9 moles of lead to react completely.
B- Calculate the number of moles of produced Ag from complete reaction of 2 moles of lead.

3-11) Calculate the mass of each product as a result of complete reaction of each substance.

$$2Al_{(s)} + Fe_2O_{3\,(s)} \longrightarrow Al_2O_{3(s)} + 2Fe_{(s)}$$

A) 4.70 g Al
B) 4.79 g Fe_2O_3

3-12) In the following reaction,

$$2A + 3B \longrightarrow C$$

When substance B is used in the following amounts, which is the specific reactant for product?

A) 2 moles of A and 3 moles of B

B) 24 moles of A and 75 moles of B

3-13) A chemist mixed 14.4 g of CaO with 13.8 g of CO_2. At the end of the reaction, he obtained 19.4 g $CaCO_3$. Find out the following according to the reaction.

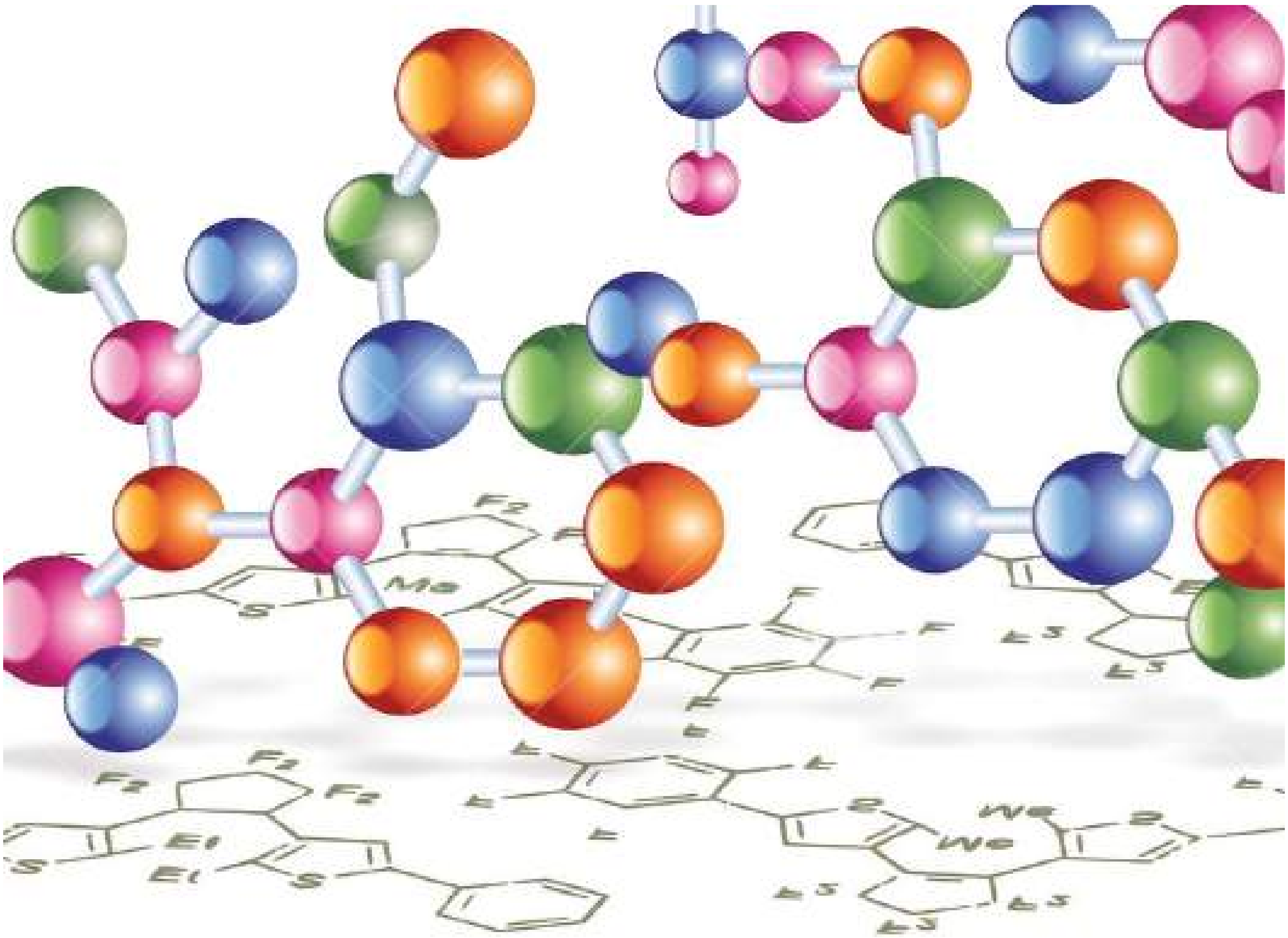
$$CaO_{(s)} + CO_{2(g)} \longrightarrow CaCO_3$$

A) What is the specific reactant for product?

B) What is the theoretical mass?

C) What is the percent yield of the product?

ORGANIC CHEMISTRY CHAPTER-4



ACHIEVEMENTS

After studying this chapter, students will be able to.

* Differentiate organic and inorganic compounds.

* Understand electronic configuration of carbon atom and its ability of forming covalent bonds.

* Named organic compounds according to general nomenclature system "IUPAC".

* Know physical and chemical properties of alkanes, alkenes and alkynes.

* Define organic components group with common formula, similar chemical properties and physical properties in organic chemistry.

* Differentiate saturated and unsaturated compounds.

* Understand the importance of organic compounds used in daily life.

4-1-PREFACE

Modern history of organic chemistry started in early 19th century. Wohler, a scientist who had contributed a lot to this branch of science resembled it to an endless forest. By the time, boundaries of organic chemistry have become clearer. Thus, it has been possible for people to understand events occurring in this limited area and also what they mean.

Organic Chemistry studies organic compounds which are commonly used in daily life and their preparation methods. These compounds are widely used in food, medicine, clothing and fuel industries. Organic compounds are different from inorganic compounds by their following properties:

1- Carbon is the main element in organic compounds. Besides, they contain elements as hydrogen, oxygen, nitrogen, sulfur and phosphorus.

2- Organic compounds usually have covalent bonds.

3- Most organic compounds are combustible and they can decompose with heat. Therefore, these compounds are accepted as the most important source of energy.

4- Reactions of organic compounds are usually slow.

5- Most organic compounds are soluble in organic solvents as alcohol, ether, gasoline, acetone and chloroform.

6- Organic compounds have a very important property called as “isomerism”.

4-2-ELECTRONIC CONFIGURATION OF CARBON ATOM

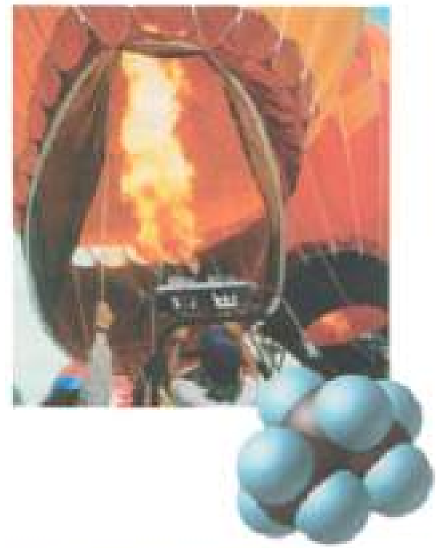
Atomic number of carbon atom is 6 and its outer electron shell is half filled with 4 electrons. Carbon atom has no possibility of gaining an electron or donating one to full out its outer shell. Therefore, carbon shares its four electrons with 4 bonds to full out outer shell. Owing to this remarkable property, carbon can form numerous organic compounds. These compounds can be in the following shapes.

Fig (4-1) shows different types of carbon chains.

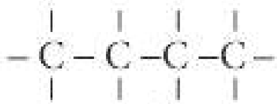
4-3-INTERMEDIATE

During chemical reactions, some bonds are broken and new bonds are formed. As a result of bond breaking, less stable intermediate products are formed. These intermediates instantly reacts and forms new bonds. Therefore, two types of bond breaking occur.

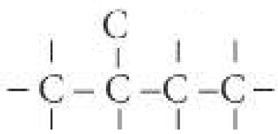
1-Homogeneous Breaking: As a result of breaking of a covalent bond between 2 atoms or groups, each group keeps one electron of the covalent bond and forms a neutral particle called as free radical.



As organic compounds are used fuel.



A



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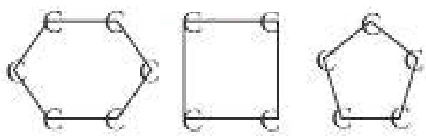

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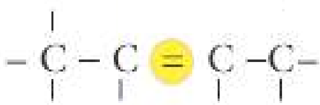


Double bond

D

Triple bond

E

Figure 4-1

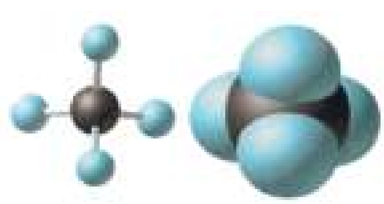
A- Unbranched single carbon chains

B- Branched carbon chains

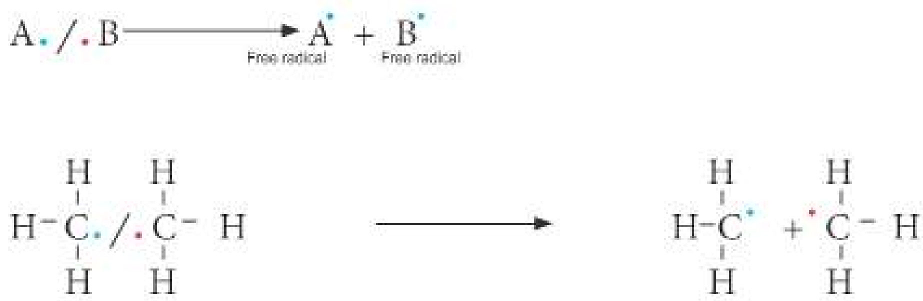
C- Cyclic, closed carbon chains

D- Carbon chains with double bond

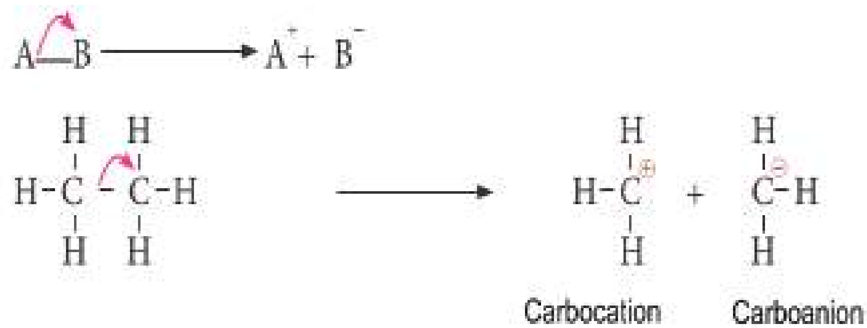
E- Carbon chains with triple bond



Methane Molecule



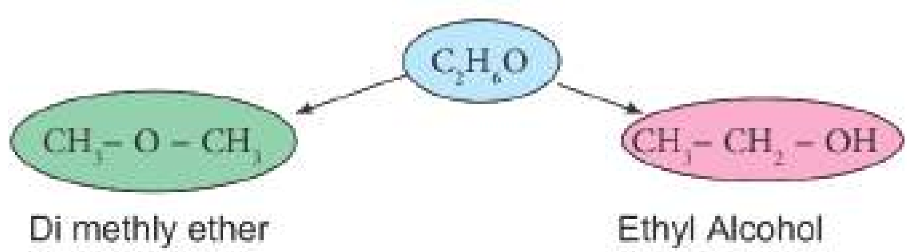
2- Non-homogeneous breaking: As a result of breaking a covalent bond between 2 atoms or groups, one of each keeps an electron pair and becomes negatively charged (carbanion). The other one becomes positively charged (carbocation).



4-4-STRUCTURAL FORMULA

A formula which shows types of elements, number of atoms, bonding orders and valences of a molecule is called as a structural formula.

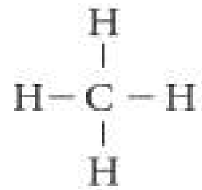
A molecular formula mostly falls short of expressing type of a compound. For example, C_2H_6O molecular formula shows either ethyl alcohol or dimethyl ether. The reason of this difference is due to the difference in bonding of atoms which form the molecule. This property is called as “isomerism”. In order to differentiate isomers, writing structural formula is necessary.



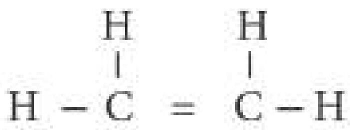
4-5-HYDROCARBONS

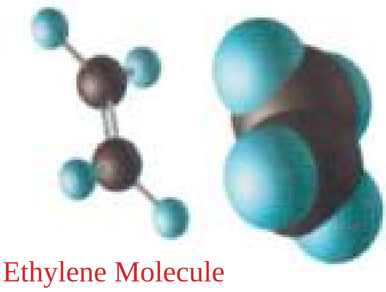
They are compounds which only consist of carbon and hydrogen atoms. They are classified according to having a closed or an open chain or being saturated or unsaturated as follows.

1-**Alkanes or Paraffins** are saturated, single bonded compounds with the general formula of $\mathbf{C_nH_{2n+2}}$. The simplest of those is methane (CH_4) gas.



2-Alkenes or olefins are unsaturated hydrocarbons with a double bond. Their general formula is $\mathbf{C_nH_{2n}}$. The simplest of those is ethylene (C_2H_4).





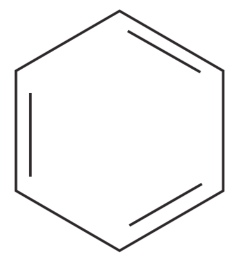
Ethylene Molecule

3-Alkynes or acetylenes are unsaturated hydrocarbons with a triple bond. Their general formula is $\mathbf{C_nH_{2n-2}}$. The simplest of those is acetylene or ethyne (C_2H_2).

$$H - C \equiv C - H$$

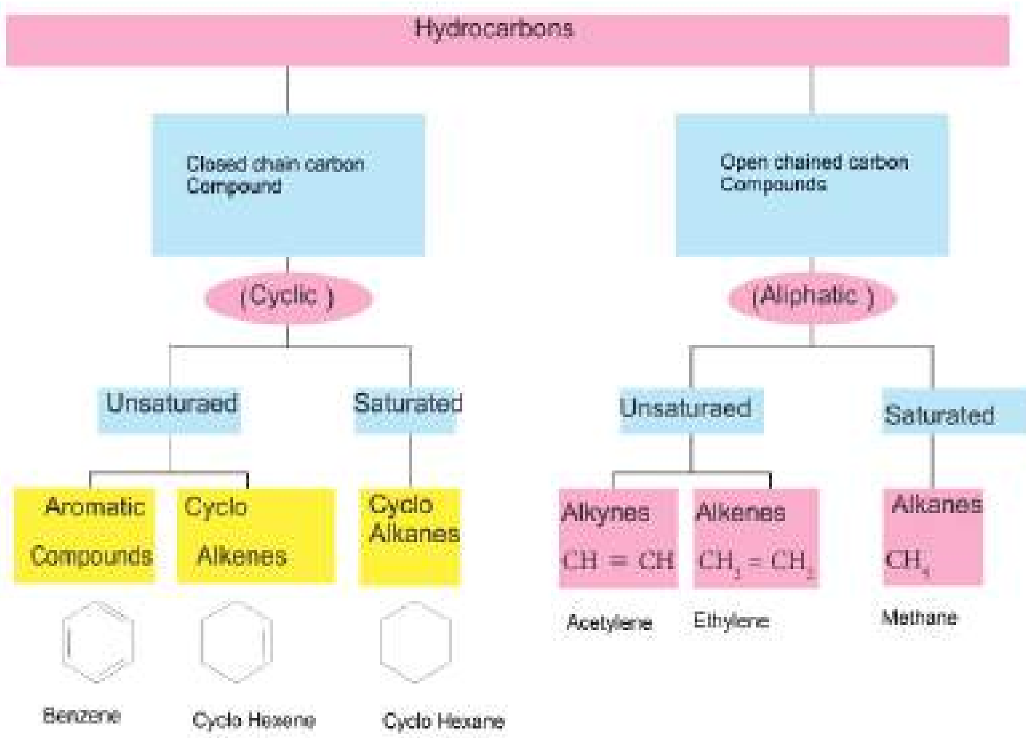
Acetylene Molecule

4-Cyclic hydrocarbons are compounds which have a carbon chain in a cyclic structure. Saturated types of those are called as cycloalkanes. e.g, cyclohexane. Unsaturated types of cyclic hydrocarbons are called as cycloalkenes. Besides, aromatic compounds or benzene or derivatives are members of this group.



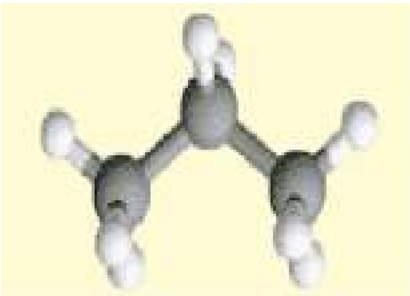
Benzene Molecule

General classification of hydrocarbons is as follows:

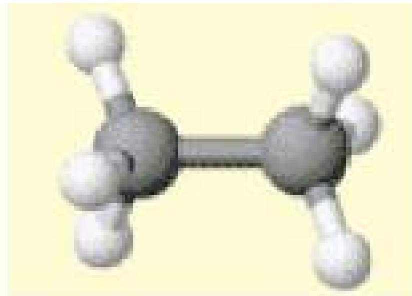


4-6-ALKANES

They are saturated hydrocarbons. They have a structure of carbon and hydrogen atoms bonded with strong and single covalent bonds. Their general formula is C_nH_{2n+2}. n is an integer and it shows number of carbon atoms. The following are examples for alkanes.



Propane



Ethane

Methane

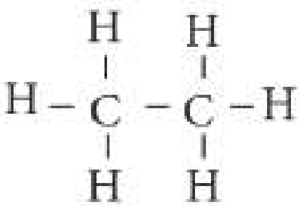
Chart 4-1 shows names and formulas for the first 10 members of alkanes.

Structural Formula	Name of Compound	Latin Prefix	Number of Carbon Atoms
CH_4	Methane	Meth-	C_1
$CH_3–CH_3$	Ethane	Eth-	C_2
$CH_3–CH_2–CH_3$	Propane	Prop-	C_3
$CH_3–CH_2–CH_2–CH_3$	Butane	But-	C_4
$CH_3–CH_2–CH_2–CH_2–CH_3$	Pentane	Pent-	C_5
$CH_3–CH_2–CH_2–CH_2–CH_2–CH_3$	Hexane	Hex-	C_6
$CH_3–CH_2–CH_2–CH_2–CH_2–CH_2–CH_3$	Heptane	Hep-	C_7
$CH_3–CH_2–CH_2–CH_2–CH_2–CH_2–CH_2–CH_3$	Octane	Oct-	C_8
$CH_3–CH_2–CH_2–CH_2–CH_2–CH_2–CH_2–CH_2–CH_3$	Nonane	Non-	C_9
$CH_3–CH_2–CH_2–CH_2–CH_2–CH_2–CH_2–CH_2–CH_2–CH_3$	Decane	Dec-	C_{10}

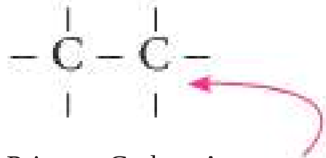
4-6-1-Types of Carbon Atoms

Carbon atoms are classified as follows according to type of bonding to each other in compounds:

1- Primary Carbon Atom: It is a carbon atom to which another carbon atom is attached.

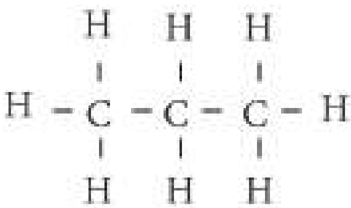


Ethane



Primary Carbon Atom

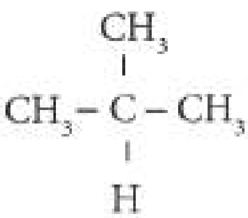
2- Secondary carbon atom: It is a carbon atom to which two carbon atoms are attached.



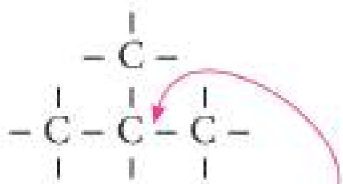
Propane

Secondary Carbon Atom

3- Tertiary carbon atom: It is a carbon atom to which 3 carbon atoms are attached.

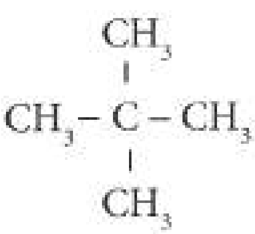


Isobutane

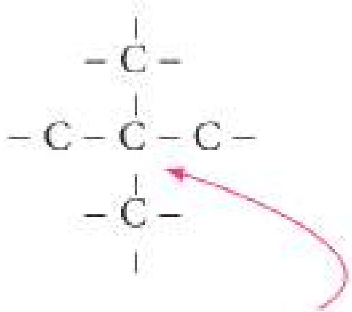


Tertiary carbon atom

4- Quaternary carbon atom: It is a carbon to which 4 carbon atoms are attached.



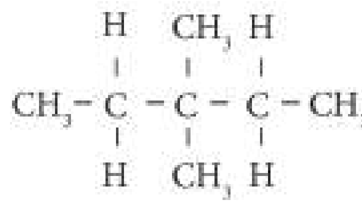
Neopantane



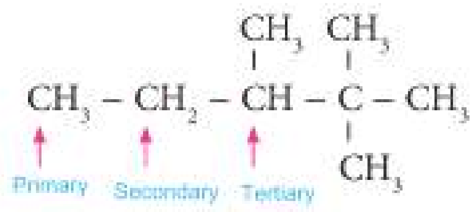
Quaternary carbon atom

Exercise 4-1

Determine the type of carbon and hydrogen atoms in the following compound.



Besides, type of hydrogen atom depends on the carbon atom it is attached to. According to this, a hydrogen atom which is attached to a primary carbon atom is called as a primary hydrogen. A hydrogen atom which is attached to a secondary carbon atom is called as a secondary hydrogen. Similarly, there is also a tertiary hydrogen atom. But quaternary hydrogen atoms do not exist. Why? Discuss with your friends.



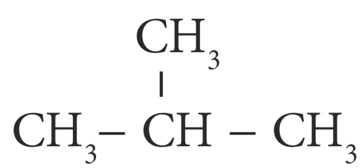
4-6-2-Nomenclature of Alkanes

There is a common and old nomenclature system used for linear and unbranched carbon atoms. While naming it starts with letter "n" (normal) and then name of alkane comes. For example:

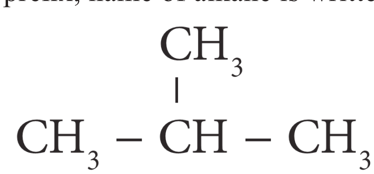
$$CH_3 - CH_2 - CH_3$$

(n- propane)

In alkanes with branched alkyl group;

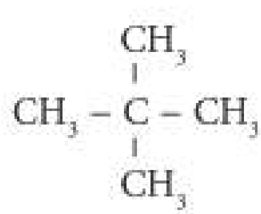


There is a common and old nomenclature with **"iso-"** prefix. After **"-iso"** prefix, name of alkane is written. For example:



(iso-butane)

For an alkane with a quaternary carbon atom, there is a common and old nomenclature which starts with the prefix **"neo-".** After neo-, name of alkane is written.



(Neo pentane)

4-6-3-Alkyl Groups

The remaining group after removal of 1 hydrogen atom from an alkane is called as an alkyl group. Its general formula is $\mathbf{C_nH_{2n+1}}$ and it is shown as (-R). **Chart 4-2** shows alkyl groups and names derived from alkane groups. In alkyl groups, in place of lost hydrogen atom, a covalent bond is put after formula as follows:

C_nH_{2n+1} —

Exercise 4-2

What is the molecular formula of the alkane with 10 carbon atoms?

Table 4-2 Nomenclature of alkyl groups

Alkyl		Alkane	
Methyl	CH_3 -	Methane	CH_4
Ethyl	C_2H_5 - CH_3 - CH_2 -	Ethane	C_2H_6 CH_3 - CH_3
Propyl n propyl Iso propyl	C_3H_7 - CH_3 - CH_2 - CH_2 - CH_3 - CH - CH_3	propane n- Propane	C_3H_8 CH_3 - CH_2 - CH_3

Example 4-1

What is the molecular formula of the alkane with four carbon atoms?

Solution:

As the general formula of alkanes is (C_nH_{2n+2})
n = number of carbon atoms, n = 4
Number of hydrogen atoms H_{2n+2}= (2 x 4) + 2 = 10
So the molecular formula is C_4H_{10}.

4-6-4-Nomenclature of Alkanes-IUPAC Nomenclature Rules:

1- Choose the longest carbon chain to be named. Then, we number carbon atoms by starting from the side closest to branching. Name is given according to the number of carbon atoms corresponding to alkane.

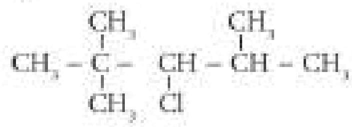
Do you know that?

The difference between an alkane and its alkyl derivative is a hydrogen atom. Name of the alkyl was derived from its alkane. –yl suffix replaces –ane suffix in the alkane. For example, propane becomes propyl.

Exercise 4-3

Write down the names for the following structural formulas.

1)

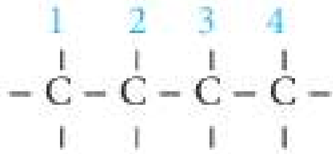


2)

Exercise 4-4

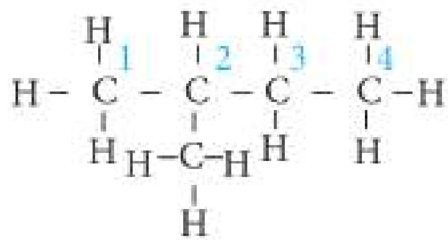
Write down structural formulas for the following compounds.

1) 2-chloro-2-methylbutane
2) 2,3-dimethylpentane



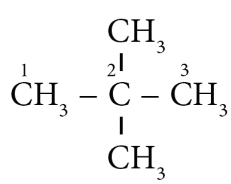
(Butane)

2- In order to determine places of alkyl groups or functional groups, we look at the number of carbon atom they are attached to.

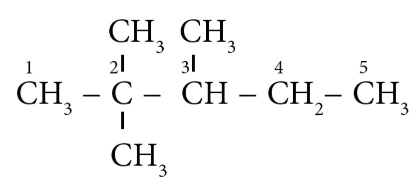


2–Methyl butane

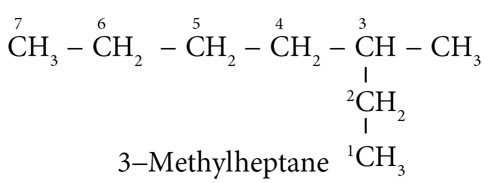
3- If there are more than one alkyl groups, they are named with respect to order of letter in Latin.
4- While naming, a comma is placed between numbers and a dash is placed between numbers and names.
5- If there are more than one of the same groups, their number is expressed by Greek numbers (-di, -tri, -tetra) before there group name. The numbers of carbon atoms to which they are attached are written separately.
1-mono, 2-di, 3-tri, 4-tetra, 5-penta
We can examine the following examples for nomenclature rules:



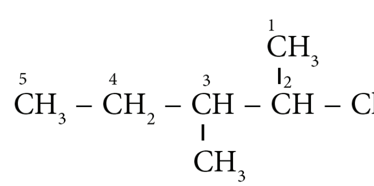
2,2–Dimethylpropane



2,2,3–Trimethylpentane



3–Methylheptane



2–Chloro–3–methylpentane

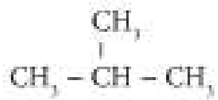
4-6-5-Isomerism

Compounds with the same molecular formula but with different structural formulas have different physical and chemical properties and they are called as **isomers.** This phenomenon starts with butane (C_4H_{10}) molecule because carbon atoms have two possibilities of bonding.
1st Possibility: Bonding of carbon atoms in an open chain (unbranched)

$CH_3-CH_2-CH_2-CH_3$

(n- butane)

2nd Possibility: Bonding of carbon atoms in a branched way.



2- Methylpropane

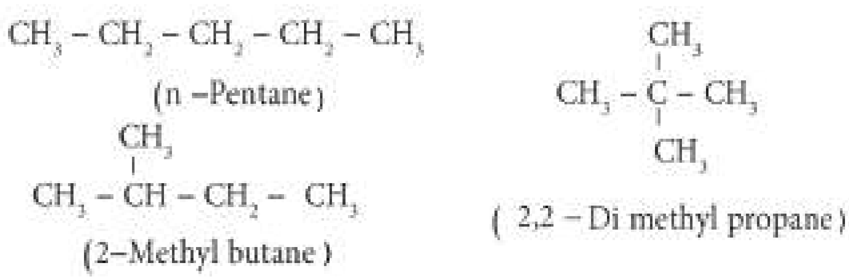
According to the old nomenclature, this compound is named as “isobutane.” As told above, butane has a single molecular formula and two different structural formulas. These two compounds which are found in nature are different by their physical and chemical properties as their structural formulas are different.

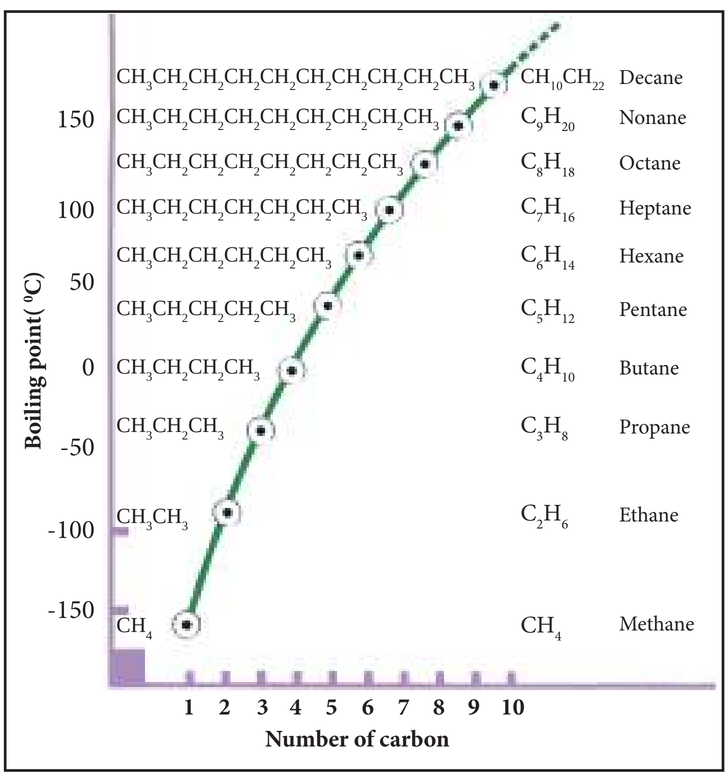
Example 4-2

Write down and name isomers of C_5H_{12}.

Solution:

C_5H_{12} has 3 isomers.





Exercise 4-5

Write down the possible structural formulas (isomers) of the alkane with the molecular formula of C_6H_{14} according to general or IUPAC system.

Exercise 4-6

Which of the following compounds has the highest boiling point?

1) C_5H_{12}
2) CH_4
3) C_8H_{18}
4) C_2H_6

Figure 4-2

The relationship between boiling points of alkanes with their number of carbon atoms.

Do you know that?

Alkanes dissolve in nonpolar solvents as oil. Therefore, it is dangerous to be exposed to vapors of alkanes. The reason is that they dissolve oil-like structures such as cell membranes, thus they cause destruction of lung tissues.

Figure 4-3

As a result of combustion of alkanes in air, CO_2 and water vapor are formed

4-6-6-Physical Properties of Alkanes

1-Solubility: Molecules of alkanes aren't polar (they are nonpolar), therefore, they aren't soluble in polar solvents as water. But they dissolve in nonpolar organic solvents as gasoline and carbon tetra chloride.

2-Boiling Point: Boiling points of alkanes increase as their molar masses increase. Methane, ethane, propane and butane are in gas form at room temperature whereas others are in liquid form. As molar mass increases, also boiling points increase. When no. of carbon atoms in molecule reaches 18, alkanes become solid. This effect on boiling points is due to the presence of weak Van der Waals attraction forces. As the distance between molecules decreases, effect of these forces increases. As molar masses of cyclic alkanes increase, surface areas of molecules which cause increase in attraction force go up. Due to the same reason, boiling points of alkanes with open carbon chains (e.g. n-pentane) are higher than boiling points of alkanes with branched carbon chains (e.g. 2-methylbuthane).

4-6-7-Chemical Properties of Alkanes

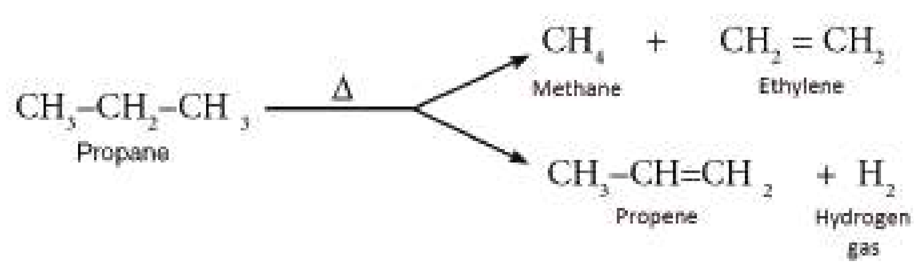
A- Chemical Reaction: As alkanes are saturated hydrocarbons, they have strong single bonds. They are less active when compared to other organic compounds as big amounts of energy are necessary to break their bonds. Alkanes don't react with strong acids as sulfuric acid and nitric acid under normal conditions. They also don't react with strong bases as sodium hydroxide or oxidizing agents as potassium permanganate.

B- Combustion Reaction: When alkanes are combusted fully in air, they yield carbon dioxide and water vapor along with smokeless blue flame. As seen in Fig. 4-3, a great amount of energy is released. Therefore, alkanes are used as fuels in industry and engines of motor vehicles. General formula of combustion of alkanes is as follows:

$$C_nH_{2n+2} + \left(\frac{3n+1}{2}\right) O_2 \longrightarrow nCO_2 + (n+1)\, H_2O + \text{Heat Energy}$$

$$CH_4 \quad + \quad 2O_2 \longrightarrow CO_2 + 2H_2O \quad + \text{Heat Energy}$$

C- Thermal Cracking: Alkanes are transformed into saturated or unsaturated compounds with heat effect in air-free medium by breaking of carbon chain or loss of hydrogen molecule. For example, propane can be cracked into two different compounds as follows:

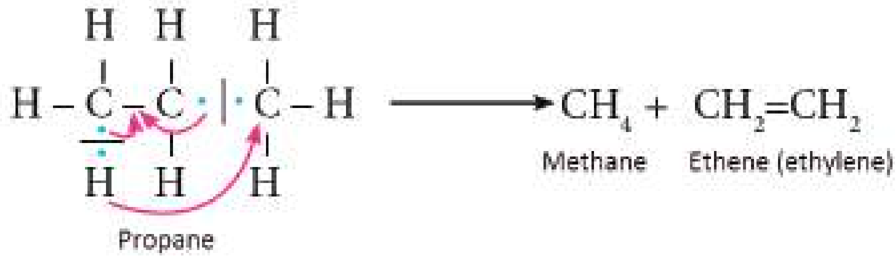


These reactions are accepted as very important as they are used in oil-refining to obtain fuels for planes, automobiles and other engines.

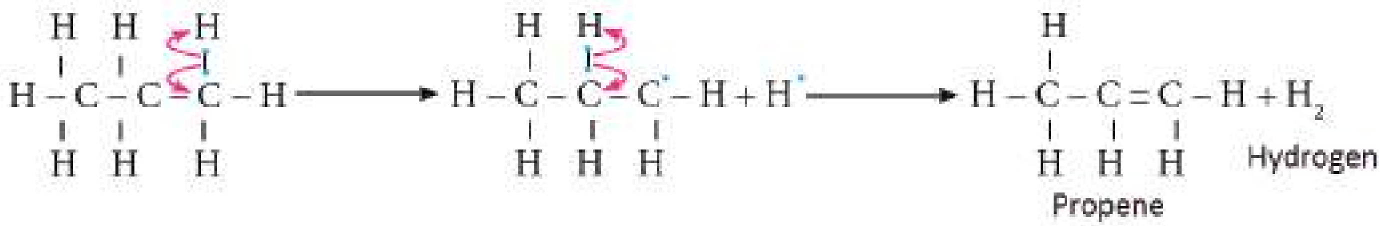
These are theoretical and actual steps which yield intermediates. In the end, actual product is obtained. Using of "reaction mechanisms" to learn actual intermediate steps of the process as follows:

1st Possibility: Free radicals form as a result of homogeneous cracking.

Generally, the largest carbon containing structure undergoes homogenous cracking. Through addition of hydrogen atom to the other radical, double bond is formed in the larger radical as follows.



2nd Possibility: To form a hydrogen atom free radical, (H-C) bond is broken. H radical is bonded with another H radical to form H_2 molecule. The remaining is propene molecule.



D-Substitution reactions

It is the substitution of one hydrogen atom in alkanes with a halide atom (Cl_2, Br_2). For example, in the reaction of methane with chlorine, the reaction continues until all hydrogen atoms in methane are replaced with chloride atoms.

$$CH_4 + Cl_2 \xrightarrow{\text{Sunlight}} \underset{\text{Methyl chloride}}{CH_3{-}Cl} + HCl$$

$$CH_3 - Cl + Cl_2 \longrightarrow \underset{\text{Dichloromethane}}{CH_2{-}Cl_2} + HCl$$

$$CH_2 - Cl_2 + Cl_2 \longrightarrow \underset{\text{Chloroform}}{CHCl_3} + HCl$$

$$CHCl_3 + Cl_2 \longrightarrow \underset{\text{Carbon tetrachloride}}{CCl_4} + HCl$$

Substitution reactions are primary and best known reactions of alkanes. These reactions can be stopped by addition of some substances.

4-6-8-Preparation of Alkanes in Laboratory

There are a few methods of preparation of alkanes in laboratory. These are mentioned below:

1- When sodium salts of carboxylic acids (R-COONa) are heated with solid sodium hydroxide or barium hydroxide, alkanes are obtained. Produced alkane has one carbon atom less than the carboxylic acid.

Exercise 4-7

Write down thermal cracking reaction of n-butane.

Do you know that?

Use of fuels as main source of energy has caused increase in carbon dioxide gas amount in upper layers of atmosphere. As a result, greenhouse effect has been produced. This effect has caused climate changes, changes in types of animals and plants and also negative results for human life.

Exercise 4-8

Obtain butane gas from sodium salt of carboxylic acid.

For example, when sodium acetate is heated with sodium hydroxide, methane gas is obtained.

$$CH_3{-}COONa + NaOH \xrightarrow{\Delta} \underset{\text{Methane}}{CH_4} + Na_2CO_3$$

When sodium propanoate (CH$_3$CHCOONa) is heated with barium hydroxide (Ba(OH)$_2$), ethane is obtained according to the following reaction:

$$2CH_3CH_2COONa + Ba(OH)_2 \xrightarrow{\Delta} \underset{\text{Ethane}}{2CH_3CH_3} + BaCO_3 + Na_2CO_3$$

2- In Grignard method, firstly, Grignard reagent is produced. In order to obtain Grignard reagent, magnesium (Mg) reacts with alkyl halide in a medium with dry ether as in the following reaction:

$$\underset{\text{Alkyl halide}}{R - X} + Mg \xrightarrow{\text{dry ether}} \underset{\text{Grignard's reagent}}{RMgX}$$

For example, Grignard reagent is obtained from methyl iodide (CH$_3$I) according to the following equation:

$$\underset{\text{Methyl iodide}}{CH_3 - I} + Mg \xrightarrow{\text{dry ether}} \underset{\text{Magnesium Methyl iodide}}{CH_3Mg\,I}$$

Alkanes are prepared from Grignard reagent by two methods.

A- If target alkane contains as many carbon atoms as Grignard reagent, the reagent is analyzed aqueously. Therefore, this method is called as aqueous analysis of Grignard reagent.

$$RMg\,X + H{-}OH \longrightarrow \underset{\text{Alkane}}{R - H} + Mg(OH)X$$

Preparation of methane is an example:

$$CH_3 - I + Mg \xrightarrow{\text{Dry Ether}} CH_3MgI$$

$$CH_3MgI + H{-}OH \longrightarrow \underset{\text{Methane}}{CH_4} + Mg(OH)\,I$$

B- If Grignard reagent reacts with an alkyl halide, no. of carbon atoms in prepared alkane is more than no. of carbon atoms in the reagent. This difference is as many as no. of carbon atoms in the alkyl halide.

$$RMgX + R - X \longrightarrow \underset{\text{Alkane}}{R - R} + MgX_2$$

$$CH_3MgI + CH_3 - I \longrightarrow \underset{\text{Ethane}}{CH_3 - CH_3} + MgI_2$$

Example 4-3

Prepare the following from ethyl chloride, 2-chloropropane and any other substances you wish.

A) n-butane B) 2-methylbutane

Solution:

A-First, we prepare Grignard reagent from ethyl chloride.

$$CH_3CH_2Cl + Mg \xrightarrow{\text{dry ether}} CH_3CH_2MgCl$$

We need an alkyl halide with two carbon atoms in order to increase number of carbon atoms and to prepare n-butane.

$$CH_3CH_2MgCl + CH_3CH_2Cl \longrightarrow CH_3CH_2CH_2CH_3 + MgCl_2$$

B-In order to prepare 2-methylbutane, we need to increase number of carbon atoms in Grignard reagent. For this, we need 2-chloropropane.

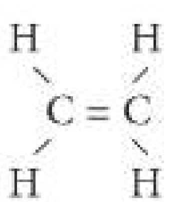
$$CH_3CH_2MgCl + \underset{\text{2- Chloro Propane}}{CH_3-\underset{\substack{|\\ CH_3}}{CH}-Cl} \longrightarrow \underset{\text{2- Methyl butane}}{CH_3-\underset{\substack{|\\ CH_2CH_3}}{CH}-CH_3} + MgCl_2$$

4-7-ALKENES (OLEFINS)

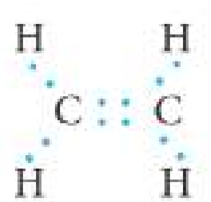
They are unsaturated hydrocarbons. They are accepted as linked chains. Members of alkenes contain less hydrogen atoms when compared with alkanes. Carbon chain has a double bond. Alkenes are shown with the general formula of (C_nH_{2n}).

$$\boxed{R-CH=CH-R' \quad \text{or} \quad R-CH=CH_2}$$

R=R' means a symmetrical alkene whereas R ≠ R' means asymmetrical alkene. Alkenes undergo addition, oxidation and substitution and combustion reactions. The simplest member of alkenes is ethylene (C_2H_4). Ethylene can be written as ;



or



or as below:

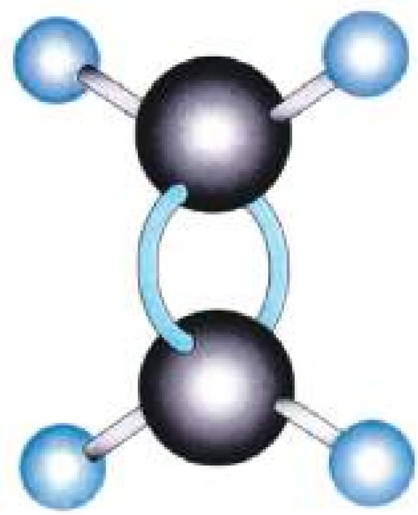
$$CH_2=CH_2$$

Exercise 4-9

Prepare the following substances using ethyl chloride and any substance you may need.

a) Ethane

b) Propane



Ethylene molecule

Exercise 4-10

Which of the following is the formula that doesn't belong to alkenes?

A) C_4H_8
B) C_5H_{10}
C) C_7H_{16}
D) C_6H_{12}

Exercise 4-11

Name the following alkenes according to general nomenclature system.

1) $CH_3-\underset{\substack{|\\ H}}{C}=CH-CH_3$

2) $CH_3-CH_2-CH=CH_2$

4-7-1-Nomenclature of Alkenes

1- Choose the longest carbon chain with double bond.

2- Number carbon atoms starting from the closest carbon to the double bond.

3- Name the longest chain as an alkane. We replace (-ane) suffix with (-ene).

4- Determine the position of double bond by picking the smaller of 2 numbers on the double bond.

5- Determine other groups with respect to the numbered carbon atom.

Table 4-3
Names of some alkenes according to IUPAC system and old (common) system

Common name	Systematic name (IUPAC)	Structural Formula
ethylene	Ethene	$CH_2 = CH_2$
propylene	Propene	$CH_3 - CH = CH_2$
butylene	1-butene	$CH_3 - CH_2 - CH = CH_2$
	2-butene	$CH_3 - CH = CH - CH_3$
pentylene	1-pentene	$CH_3 - CH_2 - CH_2 - CH = CH_2$
	2-pentene	$CH_3 - CH_2 - CH = CH - CH_3$
	3-methyl-1-butene	$CH_3 - \underset{\displaystyle CH_3}{\underset{\vert}{CH}} - CH = CH_2$
	2-methyl-1-butene	$CH_3 - CH_2 - \underset{\displaystyle CH_3}{\underset{\vert}{C}} = CH_2$

4-7-2-Geometrical Isomerism (cis-trans isomerism)

We have mentioned before that isomerism means similarity in molecular formulas but difference in structural formulas, physical and chemical properties. In alkenes, due to double bond, attached groups can't turn around the double bond freely because double bonds are shorter and stronger when compared to single bonds. But they can be positioned in 2 different orders with respect to each other. This different order seen in alkenes is called as cis-trans isomerism (geometrical isomerism). For example, including geometrical isomers, expected structural formulas for C_4H_8 molecular formula are as follows:

(A)

$CH_3 - CH_2 - CH = CH_2$

1- Butene

(B)

$CH_3 - CH = CH - CH_3$

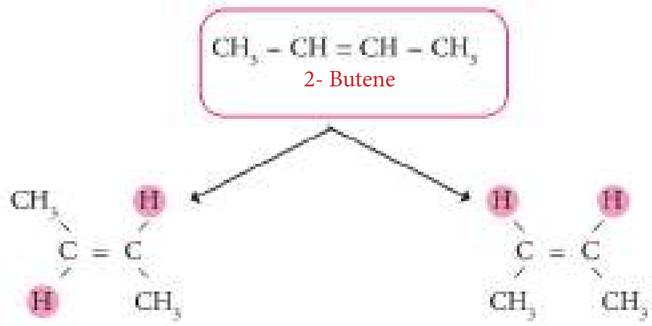
2-Butene

(C)

$CH_3 - \overset{\displaystyle CH_3}{\overset{\vert}{C}} = CH_2$

2- Methylpropene

2) Geometrical isomers of 2-butene are as follows:



trans-2- butene

Cis -2 - butene

Exercise 4-12

Write down isomers of C_5H_{10} alkene. Determine which of those is a geometrical isomer

There are two main conditions for a compound to form geometrical isomers. The first is the position of double bond (being in center not in sides), the other is no branching in carbon atoms of double bond. 2-bu tene supplies these conditions.

4-7-3-Physical Properties of Alkenes

1- First three members are in gas form, others are in liquid form

2- Boiling points increase as molar masses increase.

3- They don't dissolve in water but dissolve in organic solvents.

Fig. 4-4 shows the relationship between boiling points of alkenes and number of carbon atoms in molecular formulas.

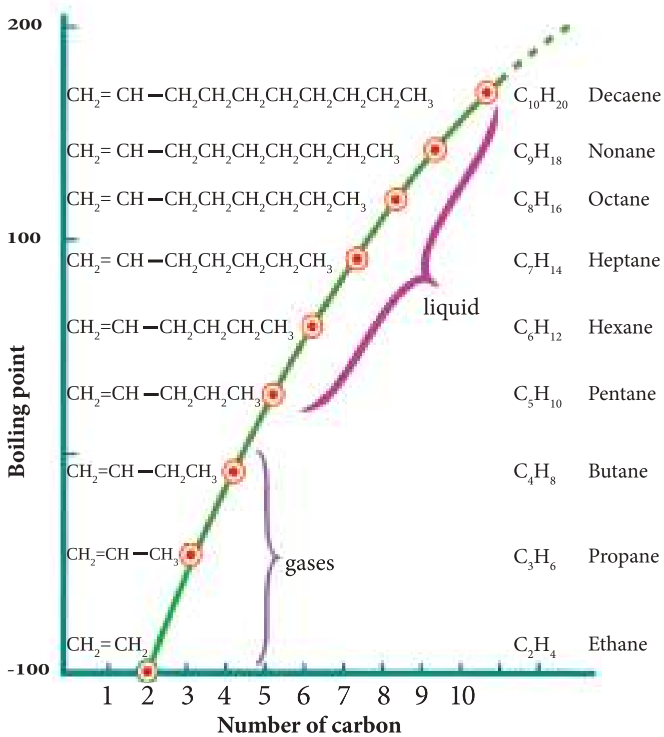


Figure 4-4

4-7-4-Reagents

1-Electrophiles

2-Nucleophiles

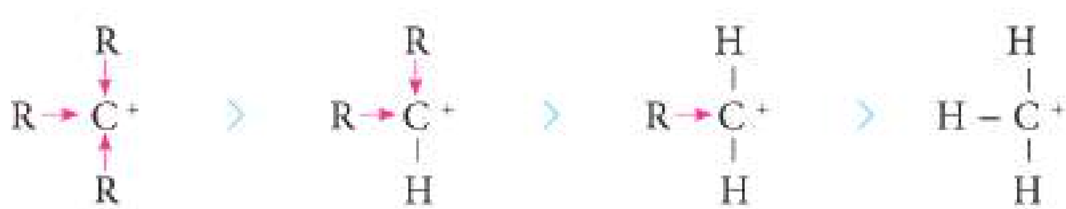
Particles (atoms, molecules or ions) which have empty orbitals and can accept a pair of electrons or electron loving reagents are called as electrophiles or Lewis acids. Particles which can share a pair of electrons by donating or nucleus loving reactants are called as nucleophiles or Lewis bases. Table 4-4 shows examples for both reagents.

Table (4-4) Examples of Electrophiles and Nucleophiles

Nucleus loving reactant Nucleophile (Lewis Bases)	Electron loving reagent Electrophile (Lewis acids)
Negative hydride ion (H^-)	Positive hydronium ion H^+
Halide ion X^-(F^-,I^-,Cl^-, Br^-)	Carbocation $-\overset{\vert}{\underset{\vert}{C}}{}^{+}$
Hydroxide Ion (OH^-)	
Carbonion $-\overset{\vert}{\underset{\vert}{C}}{}^{-}$	Boron trifluoride BF_3
Double Bond $-\overset{\vert}{C}=\overset{\vert}{C}-$	
Triple Bond $-C\equiv C-$	Polarized Carbonyl group $-\overset{\Vert}{C}-$
Ammonia NH_3	Aluminum Chloride $AlCl_3$

Stability of carbocation:

A carbocation becomes more stable as number of groups which repel bond electrons increase. For example, a carbocation which bonds 3 alkyl groups is known as the most stable type.



Tertiary most stable > Secondary > Primary > Least stable

4-7-5-Chemical Properties of Alkenes

Alkenes owe their properties to double bond (Double bond performs chemical activities of alkenes). Alkenes tend to satisfy double bond in order to become more stable. Saturated compounds with single bonds are stable (e.g. alkanes). Saturation occurs by addition of two atoms or groups to both carbon atoms of double bond. Therefore, addition includes oxidation and combustion reactions. Alkenes undergo the following reactions:

A-Addition Reactions:

Before mentioning addition reactions, it is necessary to know Markovnikov's rule. According to this rule, positive ion adds to the carbon atom of double bond with the highest no. of hydrogen atoms. Therefore, more stable carbonium ion is formed. Later, negative ion adds to the other carbon atom of the double bond.

$$R-CH=CH_2 + AB \longrightarrow R-\underset{\textstyle B}{\underset{|}{CH}}-\underset{\textstyle A}{\underset{|}{CH_2}}$$

Carbon Atom with least number of hydrogen atom

Carbon Atom with highest number of hydrogen atom

reaction mechanism

$$R-CH=CH_2 \xrightarrow{A^+} \underset{\text{carbonium ion}}{R-\overset{+}{C}H-\underset{\textstyle A}{\underset{|}{CH_2}}} \xrightarrow{B^-} R-\underset{\textstyle B}{\underset{|}{CH}}-\underset{\textstyle A}{\underset{|}{CH_2}}$$

Types of addition reactions are given below.

1-Hydrogenation (Hydrogen addition)

Alkenes react with hydrogen and saturate in presence of platinum, palladium and nickel as catalysts along with heat and pressure.

$$\underset{\text{Ethylene}}{CH_2=CH_2} + H_2 \xrightarrow[\text{suitable solvent}]{Pt} \underset{\text{Ethane}}{CH_3-CH_3}$$

This method is used to prepare alkanes and hydrogenation of vegetable oils.

Hydrogenation of alkenes is carried out in presence of platinum catalyst. According to this, hydrogen molecule is divided into positive and negative hydrogen ions. Firstly, through addition of positive hydrogen ion, carbocation is formed. Then, negative hydrogen ion is added as shown below:

$$H_2 \xrightarrow{Pt} H^+ + H^-$$

$$\underset{\text{Propene}}{CH_3-CH=CH_2} \xrightarrow{H^+} \underset{\text{Carbonium ion}}{CH_3-\overset{+}{C}H-CH_3} \xrightarrow{H^-} \underset{\text{Propane}}{CH_3-CH_2-CH_3}$$

2- Halogenation

When an alkene compound is mixed with bromine solution (red-colored) dissolved in carbon tetrachloride, we see that red color disappears as a sign of interaction of bromine with the double bond. As a result, a double halide compound is formed. This process is used to determine presence of a double bond and to differentiate an alkane from an alkene.

$$\underset{\text{Ethylene (colorless)}}{CH_2 = CH_2} + \underset{\text{Bromine (red)}}{Br_2} \xrightarrow{CCl_4} \underset{\text{1,2- Di bromo ethane colorless}}{Br - CH_2 - CH_2 - Br}$$

3-Addition of hydrogen halide-HX (HCl or HBr)

Addition occurs according to the following equation.

$$\underset{\text{2-methyl-1-propene}}{CH_3 - \underset{\displaystyle CH_3}{\overset{|}{C}} = CH_2} + HBr \xrightarrow{Br^-} \underset{\text{2-bromo-2-methylpropane}}{CH_3 - \overset{\displaystyle Br}{\underset{\displaystyle CH_3}{\overset{|}{\underset{|}{C}}}} - CH_3}$$

Exercise 4-13

While 2-bromopropane is formed from the reaction of hydrogen bromide with propene, 1-bromo-propane isn't formed. Why not?

4-Addition of Strong Sulfuric Acid to Alkenes

When ethylene gas (which is an alkene) is passed through concentrated sulfuric acid and hydrolyzed according to the simplified equation below, alcohol (ethyl alcohol) is formed.

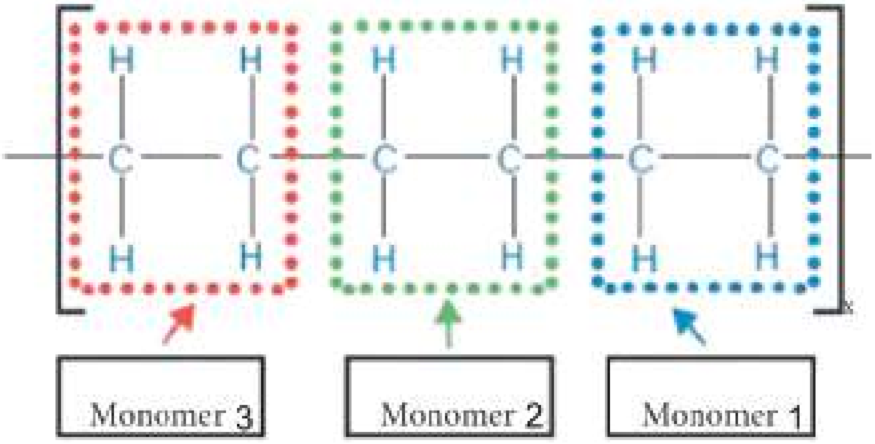
$$CH_2 = CH_2 + H_2O \xrightarrow{H_2SO_4} \underset{\text{Ethyl alcohol}}{CH_3 - CH_2 - OH}$$

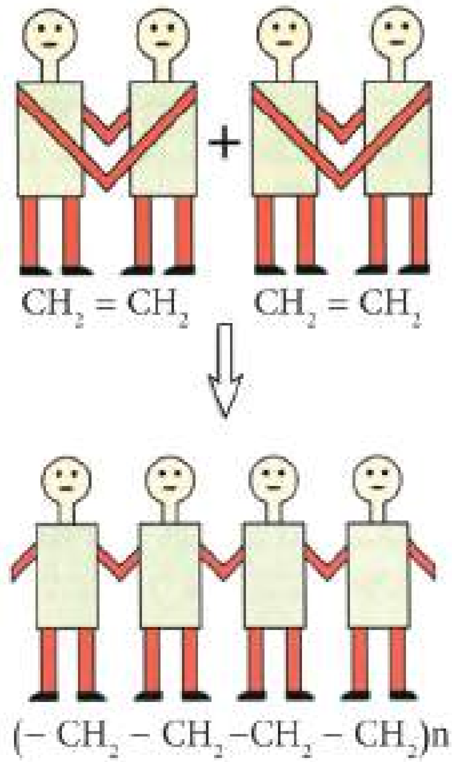
This reaction is used to separate alkenes from alkanes after thermal cracking process. The method is also accepted as a commercial process to produce alcohol. Intermediate product (ethyl hydrogen sulfate) decomposes through hydrolysis and forms alcohol and concentrated sulfuric acid.

B-Polymerization

It is a type of addition reaction of alkenes. Small alkene molecules which are called as monomers build up in presence of a suitable catalyst (e.g. sulfuric acid). As a result, saturated molecules which are called as polymers with large molar masses are produced. This reaction is called as polymerization. Polymers are used in plastic production. For example, from the reaction of ethylene molecules, a substance called as polyethylene is produced.

$$n\,(CH_2 = CH_2) \xrightarrow{H_2SO_4} \underset{\text{Polyethylene}}{[-CH_2-CH_2(CH_2-CH_2)_nCH_2-CH_2-]_x}$$





Polymerization of ethylene

C- Combustion Reactions

Alkenes burn with a smoky flame in air (because ratio of carbon is higher in alkenes than in alkanes). As a result of combustion, CO_2 and water vapor (H_2O) are formed. Besides, energy is released.

$$CH_2 = CH_2 + 3O_2 \longrightarrow 2CO_2 + 2H_2O + \text{Energy}$$

D- Oxidation Reactions

When alkenes are mixed with cold dilute potassium permanganate ($KMnO_4$ –purple-colored) solution, the strong oxidizing agent potassium permanganate breaks double bond and purple color disappears rapidly. At the end of oxidation, a substance called as glycol and manganese dioxide (MnO_2)-a brown precipitate- are formed.

$$3CH_2 = CH_2 + \underset{\text{purple solution}}{2KMnO_4} + 4H_2O \longrightarrow \underset{\text{Ethylene glycol}}{3\underset{OH}{CH_2} - \underset{OH}{CH_2}} + 2KOH + \underset{\text{dark brown precipitate}}{2MnO_2}$$

When hot concentrated potassium permanganate solution is used, ethylene is oxidized completely as in the following equation.

$$CH_2 = CH_2 + 4KMnO_4 \longrightarrow 2CO_2 + 4KOH + 4MnO_2$$

This method is used to separate alkenes and alkanes as in bromine method told in the 2nd method.

4-7-6-Preparation of Alkenes in Laboratory

1-Dehydration of Alcohols

In order to accomplish this reaction, several catalysts as concentrated sulfuric acid are used. Ethylene is produced when concentrated sulfuric acid is mixed with ethyl alcohol and heated until 165 °C.

Do you know that?

Polymerization reactions have a great importance especially in plastic industry. They are also used in production of artificial rubber and many other polymers.

Exercise 4-14

Show how 2-methylpropene can be differentiated from butane using bromine solution dissolved in CCl_4 via chemical equations.

Exercise 4-15

Prepare 1-butene using a proper alcohol and necessary substances.

$$\underset{\text{Ethyl alcohol}}{CH_3-CH_2-O-H} \xrightarrow[165\,°C]{H_2SO_4} CH_2=CH_2 + H_2O$$

We should emphasize here that hydrogen atom is lost from the other carbon atom which is bonded to the carbon with OH group. You will learn the mechanism of this reaction in the following chapters.

Example 4-4

Prepare propene using a proper alcohol and necessary substances.

Solution:

We need to choose an alcohol with 3 carbon atoms in order to prepare propene $CH_3CH{=}CH_2$.

In this case, we choose propyl alcohol $CH_3CH_2CH_2OH$. The reaction occurs according to the following equation:

$$\underset{\text{Propyl alcohol}}{CH_3CH_2CH_2OH} \xrightarrow[165\,°C]{H_2SO_4} CH_3CH{=}CH_2 + H_2O$$

Exercise 4-16

Prepare 1-butene using proper alkyl halide and necessary substances.

2-Removal of Hydrogen Halide (Dehydrohalogenation) from Alkyl Halides

Alkenes can be prepared from the reaction of a strong base as KOH dissolved in alcohol (alcohol assumes the role of a catalyst here) and an alkyl halide (R-X). Alkenes is easily released.

$$\underset{\text{2-chloropropane}}{CH_3-\underset{\displaystyle Cl}{\underset{|}{CH}}-CH_3} + KOH \xrightarrow[\Delta]{\text{Alcohol}} \underset{\text{Propene}}{CH_3-CH{=}CH_2} + KCl + H_2O$$

Reaction Mechanism:

Notice that hydrogen is removed from the other carbon atom which is bonded to the carbon with halide (chloride).

4-8-ALKYNES-ACETYLENES

They are the third type of hydrocarbons. They have the general formula of C_nH_{2n-2} and the structural formula of $R-C\equiv C-H$ or $R-C\equiv C-R$.
Their characteristic is to have a triple bond. Their first member is acetylene ($-C\equiv C-$), therefore they are also called as acetylenes.
3 pairs of electrons between 2 carbon atoms form a triple bond in acetylenes.

Acetylenes undergo addition reactions just as olefins do. The reason for this is that triple bonds can react with electrophiles. Therefore, catalysts are used in most addition reaction of acetylenes. Besides, a hydrogen atom on the carbon of a triple bond is more active than a hydrogen atom on the carbon of a double bond, thus, it can undergo a substitution reaction with a metal. Hydrogen atom on a triple bond is accepted as acidic.

Acetylene gas

4-8-1-Nomenclature of Alkynes

1- The longest chain including both carbon atoms of a triple bond is chosen. Then, the chain is numbered as the carbon of triple bond will get the smallest number. The longest chain is given the name of corresponding alkane. "-ane" suffix of alkane is replaced with "-yne" suffix. The position of the triple bond is mentioned by the smaller number on both carbon atoms.

2- Names of branches are given. Their positions are mentioned through the number of the chain. The following examples show nomenclature of some alkynes.

$$H-C\equiv C-H \quad \text{Ethyne}$$

$$CH_3-C\equiv C-H \quad \text{Propyne}$$

$$CH_3-CH_2-C\equiv C-H \quad \text{1-butyne}$$

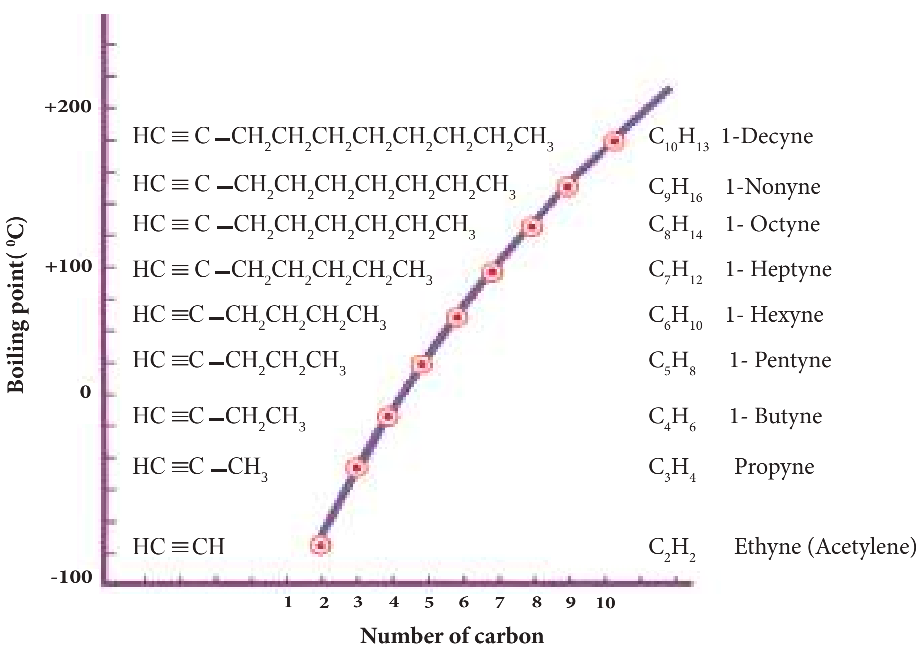
$$CH_3-C\equiv C-CH_3 \quad \text{2-butyne}$$

$$\begin{array}{c} \quad\quad CH_3 \\ \quad\quad | \\ CH_3-C-C\equiv C-H \\ \quad\quad | \\ \quad\quad CH_3 \end{array} \quad \text{3,3-dimethyl-1-butyne}$$

4-8-2-Physical Properties of Alkynes

1-Their boiling points increase with increasing molar masses. The first 4 members are in gas form whereas the remaining are in liquid form.

2-They dissolve in water and in other polar solutions slightly, but they are highly soluble in organic solvents.



4-8-3-Chemical Properties of Alkynes

Alkynes which have a structural formula of (R - C ≡ C - H) possess two active groups. These are:

1) First active group: triple bond

2) Second active group: weak acidic hydrogen which can react with metal acetylides as sodium acetylide.

A- Addition Reactions

Presence of unsaturated triple bond in alkynes enables them to behave like alkenes (olefins) and shows similar chemical activities. The reason is alkynes tend to saturate this bond completely or partially. Addition reactions are as follows:

1-Hydrogenation:

It is possible to saturate alkynes by reacting them with hydrogen gas in presence of nickel or platinum catalysts. This process functions in two steps. In the first step, an alkene (olefin) is produced; in the second, an alkane is obtained.

$$\underset{\text{propyne}}{CH_3-C\equiv C-H} \xrightarrow[H_2]{Ni} \underset{\text{Propene}}{CH_3-CH=CH_2} \xrightarrow[H_2]{Ni} \underset{\text{propane}}{CH_3-CH_2-CH_3}$$

2-Addition of Hydrogen Halide

Addition of hydrogen halide (HX) acids is similar to the reaction of hydrogen halide addition to alkenes which we have mentioned before. But here it occurs in two steps and compounds with two halides (e.g. 2,2-dibromopropane) are produced.

$$CH_3-C\equiv C-H + HBr \longrightarrow \underset{\text{2- bromo propene}}{CH_3-\overset{\overset{Br}{|}}{C}=CH_2} \xrightarrow{HBr} \underset{\text{2,2- di bromo propane}}{CH_3-\overset{\overset{Br}{|}}{\underset{\underset{Br}{|}}{C}}-CH_3}$$

3-Addition of Halides (Br_2, Cl_2)

A halide molecule adds to a triple bond readily and haloalkenes are produced. Then, by addition of one more halide, double bond of alkene is saturated and an alkane is obtained.

$$\underset{\text{propyne}}{CH_3-C\equiv C-H} + Br_2 \longrightarrow \underset{\text{1,2- dibromopropene}}{CH_3-\overset{\overset{Br}{|}}{C}=\overset{\overset{Br}{|}}{C}-H} \xrightarrow{Br_2} \underset{\text{1,1,2,2-tetrabromopropane}}{CH_3-\overset{\overset{Br}{|}}{\underset{\underset{Br}{|}}{C}}-\overset{\overset{Br}{|}}{\underset{\underset{Br}{|}}{C}}-H}$$

B) Acetylide Reactions

They are the reactions of acidic hydrogen atom of triple bond carbon. An acetylide is a compound derived from active metals as sodium and calcium. As result of decomposition of those in aqueous medium, an alkyne is produced.

$$H-C\equiv C-H + Na \longrightarrow \underset{\text{sodium acetylide}}{H-C\equiv \bar{C}\,Na^+} + \frac{1}{2}H_2$$

$$H-C\equiv C-H + 2Na \longrightarrow \underset{\text{disodium acetylide}}{Na^+\bar{C}\equiv \bar{C}\,Na^+} + H_2$$

4-8-4-Difference between an acidic alkyne and non-acidic alkyne

Through use of Tollens' Reagent, acidic and non-acidic alkynes as 1-butyne and 2 butyne can be differentiated completely. Tollens' Reagent is a solution of ammonia-silver ion obtained from solution of silver oxide or silver nitrate in ammonia. Tollens' Reagent ($Ag(NH_3)_2OH$) reacts with 1-butyne and forms a white precipitate, silver acetylide. But Tollens' Reagent doesn't react with 2-butyne as it doesn't have an active acidic hydrogen atom.

$$\underset{\text{1-butyne}}{CH_3-CH_2-C\equiv C-H} + \underset{\text{Tollens' Reagent}}{Ag(NH_3)_2OH} \longrightarrow \underset{\text{white precipitate}}{CH_3-CH_2C\equiv C^-Ag^+\downarrow} + 2NH_3 + H_2O$$

$$\underset{\text{2- butyne}}{CH_3-C\equiv C-CH_3} + Ag(NH_3)_2OH \longrightarrow \underset{\text{No reaction}}{N.R}$$

4-8-5-Preparation of Alkynes

A-Preparation of Acetylene gas in industry and laboratory

1- Acetylene is produced from the aqueous reaction of calcium carbide as shown in the following equation.

$$CaC_2 + 2H_2O \rightarrow H-C\equiv C-H + Ca(OH)_2$$

Calcium Carbide acetylene gas (ethyne)

2- Acetylene gas is obtained from heating methane gas in air-free medium at high temperatures as shown in the following equation.

$$2CH_4 \xrightarrow{1500°C} H-C\equiv C-H + 3H_2$$

B-Preparation of Alkynes with high molar masses

Acetylenes are prepared from acetylene gas. It is first transformed into sodium acetylide. Then, an alkyne is produced from the reaction of sodium acetylide with a proper alkyl halide. For example, 1-butyne and 2-butyne can be obtained from acetylene gas.

$$\underset{\text{acetylene}}{CH \equiv CH} + Na \longrightarrow \underset{\text{sodium acetylide}}{CH \equiv C^- Na^+} + \frac{1}{2} H_2$$

$$CH \equiv C^- Na^+ + I - CH_2 - CH_3 \longrightarrow \underset{\text{1-butyne}}{CH \equiv C - CH_2 - CH_3} + NaI$$

$$\underset{\text{acetylene}}{H - C \equiv C - H} + 2Na \longrightarrow Na^+ C^- \equiv C^- Na^+ + H_2$$

$$Na^+ C^- \equiv C^- Na^+ + 2CH_3 - I \longrightarrow \underset{\text{2-butyne}}{CH_3 - C \equiv C - CH_3} + 2NaI$$

Exercise 4-17

$CH_3C \equiv CH + 2HBr \longrightarrow$

Which of the following compounds is the product of the reaction above?

A) $CH_3CBr_2CH_3$

B) $CH_3CHBrCH_2Br$

C) $CH_3CH_2CHBr_2$

D) $BrCH_2CH_2CH_2Br$

Example 4-5

Prepare 2-pentyne from propyne and ethyl iodide.

Solution:

In order to obtain sodium propynide, propyne and sodium must react.

$$CH_3C \equiv CH + Na \longrightarrow CH_3C \equiv C^- Na^+ + \frac{1}{2} H_2$$

By the reaction of sodium propynide (Sodium 1-propyn-1-ide) with ethyl iodide, we get 2-pentyne according to the following equation.

$$CH_3C \equiv C^- Na^+ + I - CH_2 - CH_3 \rightarrow CH_3C \equiv C - CH_2 - CH_3 + NaI$$

BASIC CONCEPTS

Alkanes: They are saturated hydrocarbons. Carbon atoms are bonded to each other with (C- C) single bonds.

Carbon chains in organic compounds :

1-Unbranched carbon chains

2-Branched carbon chains

3-Closed carbon chains

4-Chains with double bond (alkenes) or chains with triple bond as acetylenes (alkynes).

Alkenes: They are unsaturated hydrocarbons. $(CH_2 = CH_2)$ A double bond exists between carbon atoms. Their general formula is C_nH_{2n}.

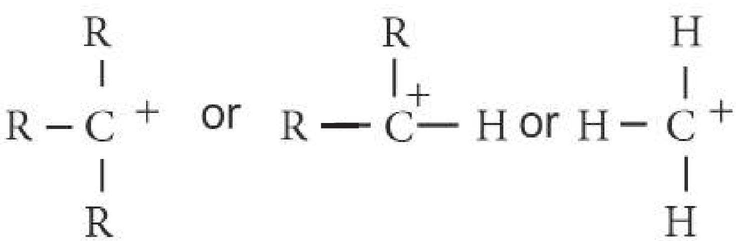
Alkynes: They are unsaturated hydrocarbons. $CH \equiv CH$ A triple bond exists between carbon atoms.

Alkyl Group: The group which remains after removal of one hydrogen atom from an alkane. Its general formula is C_nH_{2n+1}. For example, methyl group $-CH_3$ or ethyl group $-C_2H_5$.

Electrophile: Reactants which can accept a pair of electrons, electron loving reactants or particles (atoms, molecules or ions) with empty orbitals are called as electrophiles.

Nucleophile: Particles (atoms, molecules or ions) which can donate a pair of electrons and nucleus loving reactants are called as nucleophiles.

Carbocation : A carbon atom ion which has 3 hydrogen atoms or another group on it. It carries a positive charge.



Polymerization: It is one of the addition reactions of alkenes (olefins). Here, molecules of monoalkenes combine with each other with a proper catalyst and form saturated compounds called as polymers with high molar masses.

Grignard reagent: In order to prepare Grignard reagent, Mg interacts with an alkyl halide in dry ether medium. Its general formula is (X-Mg-R).

Sodium acetylide: It is a compound obtained from the reaction of sodium metal with hydrogen atom in acetylene.

$$R-C\equiv C-H + Na \longrightarrow \underset{\text{Sodium acetylide}}{RC\equiv C^- - Na^+} + \frac{1}{2}H_2$$

Substitution Reactions: They occur by replacing a hydrogen atom in an organic compound with another atom or group (e.g. –Br or –Cl).

QUESTIONS OF CHAPTER- 4

Reminder: In some questions, you may need to know atomic masses. You may refer to tables at the end of the book.

4.1) What are the most significant properties of organic compounds? What are the properties which differentiate organic compounds from inorganic compounds?

4.2) What is a hydrocarbon? How is it classified?

4.3) What is the distinctive property of carbon atom from other atoms?

4.4) What is a homologous series? What uses does it have?

4.5) Define the following.

A) Isomer

B) Structural formula

4.6) Give the Name the following structural formulas according to IUPAC system.

A) $$\begin{array}{ccccc} & & CH_3 & & \\ & & | & & \\ CH_3 & - & CH & - & CH_3 \end{array}$$

B) $CH_3 - CH = CH - CH_3$

C) $CH_3 - C \equiv C - CH_3$

D) $CH_3 - C \equiv C - H$

E) $$\begin{array}{ccccc} & & CH_3 & & \\ & & | & & \\ CH_3 & - & C & - & CH_3 \\ & & | & & \\ & & CH_3 & & \end{array}$$

F) $CH_3 - CH = CH_2$

4.7) What are the common and systematic names of the following?

A) $$\begin{array}{ccc} & & CH_3 & & \\ & & | & & \\ CH_3 & - & C & = & CH_2 \end{array}$$

B) $CH_3 - CH_2 - CH_3$

C) $CH_2 = CH_2$

D) $H - C \equiv C - H$

4-8) Write down structural formulas for the following compounds.

A) 2-methyl-2-butene

B) 2-methyl-1-pentene

C) 2,2-dimethylpropane

4.9) What is the structural formula of alkane (paraffin) with the molecular formula of C_5H_{12}?

4.10) As atomic mass of H is 1 and atomic mass of C is 12, what are the structural formulas for alkanes with a molar mass of 72 g/mol?

4.11) Fill in the blanks.

1. Write down general formulas of the following:

-Alkanes ……..

-Alkenes ………

-Alkynes ………

2. Write down structural formulas of the following:

-Alkanes ……..

-Alkenes ……

-Alkynes ……

3. What are the functional groups of the following?

-Alkenes ……

-Alkynes ……

4.12) Explain the following

A) Why do we have to write down structural formulas sometimes?

B) A quaternary hydrogen atom or higher don't exist. Why not?

C) As molar masses of substances increase, boiling points also increase. Why?

D) Alkanes are insoluble in water. Why not?

E) Alkanes aren't active compounds. Why not?

F) When we add HBr to propene, 2-bromopropane is produced but not 1-bromopropane. Why not?

G) Addition of concentrated sulfuric acid to alkenes and analysis of its product in aqueous medium is both commercially and industrially important. Why?

H) Tollens' reagent doesn't react with 2-butyne, but reacts with 1-butyne. Why?

4.13) Write down preparation methods of each of the following.

A) Propane B) propene C) propyne

4.14) How do you prepare the following starting from ethyl chloride and using necessary substances?

A) Ethane

B) Ethylene

Also tell how you differentiate those above in laboratory.

4.15) How do you prepare the following by starting from calcium carbide and using necessary substances?

A) Propyne

B) 2-butyne

Also tell how you differentiate those above.

4.16) Express the following reactions with their structural formulas.

1- Addition of hydrogen bromide to propene

2- Withdrawing KCl from ethyl chloride with KOH in alcoholic medium.

3- Withdrawing water from ethyl alcohol by concentrated sulfuric acid at 165°C.

4- Oxidation of ethylene by hot concentrated potassium permanganate.

5- Reaction of sodium acetylide with 2-chloropropane.

4-17) Circle the correct answer:

1-Alkanes:

A) are always in gas form.

B) are soluble in water.

C) contain single covalent bonds.

2- In which of the following molecules is the general formula of alkynes is valid?

A) C_3H_8 B) C_3H_6 C) C_3H_4

3-Which of the following reagents is used to differentiate ethylene and ethane?

A) Red bromine water

B) Lime water

C) Silver nitrate solution

4.18) Which of the following molecules is an alkyne?

A) $C_{15}H_{32}$

B) $C_{20}H_{38}$

C) C_9H_{20}

4.19) What is the molecular formula of the alkene which consists of 4 carbon atoms?

4.20) Which of the following is the number of covalent bonds in the molecular formula of acetylene ($H - C \equiv C - H$) ?

A) 3

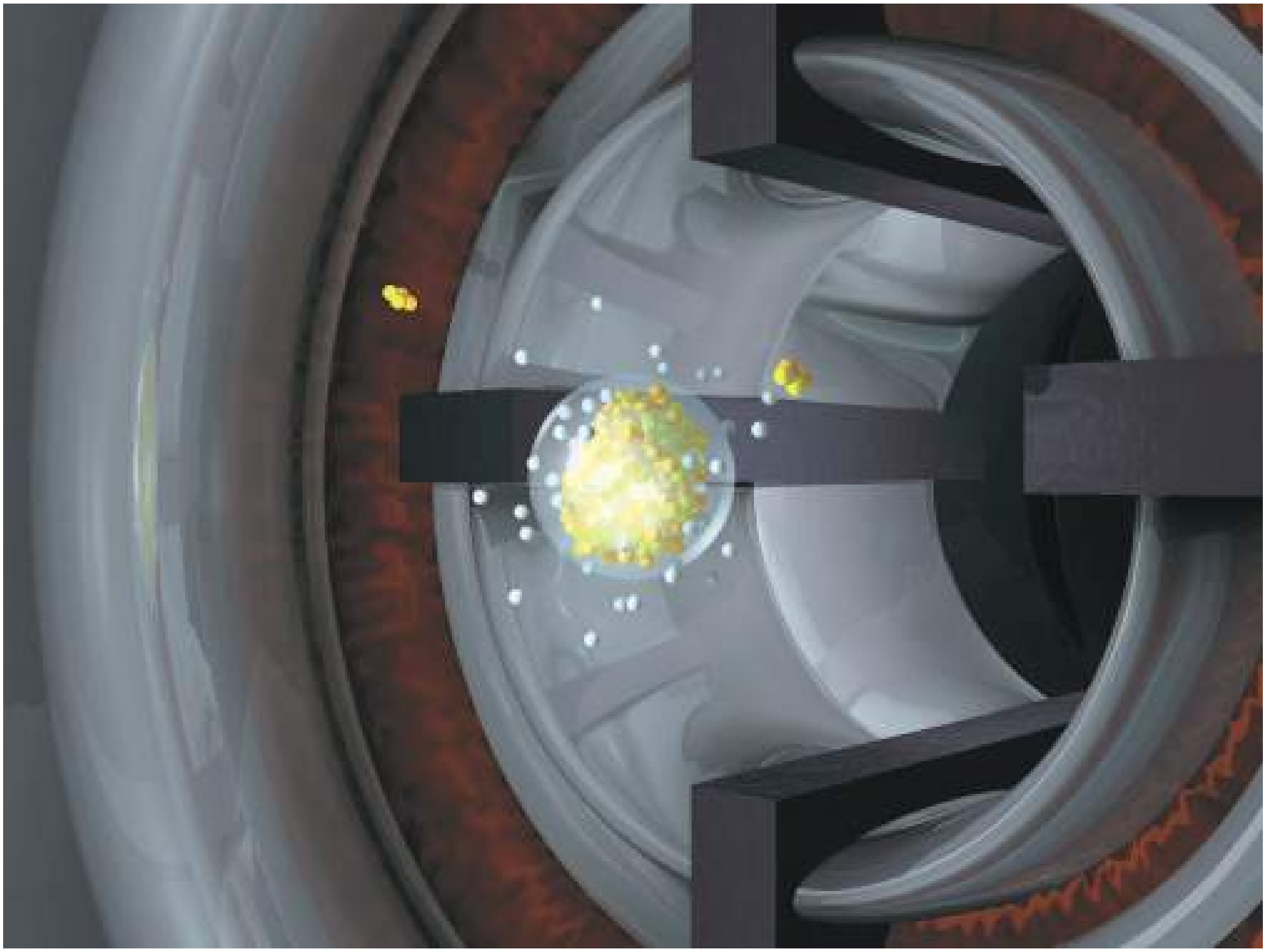
B) 2

C) 5

4.21) Write down the reaction showing addition of hydrogen gas to 2-butyne by using proper catalyst.

NUCLEAR CHEMISTRY

CHAPTER-5



ACHIEVEMENTS

After studying this chapter, students will be able to.

*Know nucleus and its importance in nuclear reactions.

*Know isotopes and their types. They learn their uses in medicine, science and archeology fields.

*Learn how to determine half lives of isotopes and understand the importance of ^{14}C isotope.

*Know the definition of radioactivity and can tell types of it.

*Write and balance nuclear reactions.

*Differentiate nuclear fission and nuclear fusion and learn their properties.

*Understand dangers of ionized radiation, its effects on living beings and ways of protecting from it.

Big Bang

5-1-PREFACE

Scientists have agreed that our universe were created as a result of "Big Bang". If we consider the temperature of Big Bang was at a level of a few billion centigrades, we can understand that inconsiderable amount of energy and particles (protons, neutrons and electrons) were released by this event. From these particles, different elements were formed. Matter was in plasma state (fourth state of matter). Plasma is a sea of positive nuclei and negative electrons.

5-2-NUCLEUS

When scientists studied different elements, they noticed that they were made up of tiny particles called as atoms. Atoms consist of main particles represented by nucleus. Nucleus is the axis around which nuclear chemistry evolves. Though nucleus is too small, it carries most of atomic mass inside and its mass is much greater than mass of electrons. Electrons are negatively charged (e^-) and they turn around nucleus at a great speed. Amount of positive charge in an atom is equal to amount of negative charge. Therefore, it is electrically neutral. The reason of positive charge of atom is due to particles called as protons. These are positively-charged and shown with (p+). In an atom, there are also neutral particles. These are called as neutrons and shown with(n°). Number of nucleons in an atom is equal to sum of protons and neutrons. This is named as mass number and shown with "A". Number of protons in an atom is called as atomic number and shown with "Z". This is also equal to number of electrons in a neutral atom. This number determines the position of an element in the Periodic Table. These numbers can be shown with the following order around symbol of an element.

mass number $\longrightarrow$ A

atomic number $\longrightarrow$ Z

$$^{A}_{Z}X$$

$$N = A - Z$$

If we know mass number and atomic number of an element, we can find number of neutrons (n) from this information.

5-3-ISOTOPES

An element can have different mass numbers but its atomic number doesn't change. Isotopes are different forms of a single element. Isotope means "the same place". Therefore, isotopes of an atom fill the same place in the Periodic Table because there are same numbers of protons in the nuclei of isotopes of an element. But they have different number of neutrons. Elements in nature are divided into two: Some have isotopes whereas others not. For example, hydrogen element has 3 isotopes. These are:

-Normal hydrogen $^{1}_{1}H$

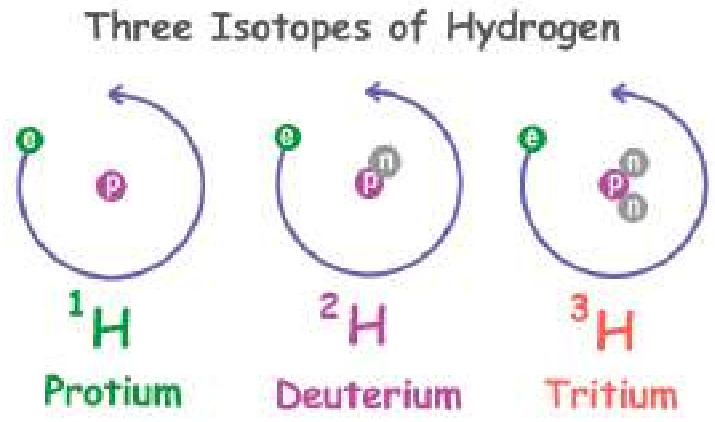
-Heavy hydrogen or deuterium $^{2}_{1}D$ or $^{2}_{1}H$

-Heaviest hydrogen tritium $^{3}_{1}T$ or $^{3}_{1}H$

Do you know that?

In Arabic, word "zerre" is used to define a type of tiny red ant. This ant is the smallest type of ant family. This name "zerre" is also used for tiny objects seen when sunlight passes through a narrow hole.

Figure 5-1



These isotopes show a difference of abundance in nature. Table 5-1 shows ratios of abundance of those isotopes in nature.

Most of hydrogen atoms are found in water (H_2O). In every 6000 normal water molecules, there is only 1 molecule of (D_2O) (heavy water). In order to obtain heavy water, normal water is electrolyzed.

In this process, normal hydrogen is isolated more easily than heavy hydrogen. As the time goes by in hydrolysis, concentration of D_2O builds up. D_2O is used in electric energy production, in reactors as an inhibitor and in many fields.

You need to know that isotopes have same number of electrons and protons. Therefore, they have similar chemical properties. The reason is that number of electrons determines chemical properties of an atom. The difference between mass numbers of an atom depends on sum of numbers of neutrons and protons. As this number determines nuclear properties of an atom, nuclear properties of isotopes are different.

Table 5-1 Isotopes of hydrogen element and its abundance in nature

Name of isotopes	Symbol	Number of ^{0}n	Percentage %
Hydrogen	$^{1}_{1}H$	0	99.984%
Deuterium	$^{2}_{1}H$, $^{2}_{1}D$	1	0.015%
Tritium	$^{3}_{1}H$, $^{3}_{1}T$	2	quite rare and radioactive

As a different example, Uranium has 3 isotopes.

$$^{238}_{92}U ,\ ^{235}_{92}U ,\ ^{234}_{92}U$$

Most elements are found as a mixture of two or more isotopes in nature. Ratio of abundances of isotopes of an element in nature is called as "isotopic abundance". This ratio can be different for every element. We find atomic mass from average of masses of all isotopes of an element. Atomic mass is obtained by multiplication of mass number of each isotope with isotopic abundance and sum of these multiplications. It is measured with atomic mass unit (amu).

According to this, **1amu = 1.66×10^{-24}**

Exercise 5-1

Study the symbols of elements below. Then answer the questions.

$^{31}_{15}P$ $^{12}_{6}C$ $^{23}_{11}Na$

1-What does the number on lower left corner of an element symbol show?

2- What does the number on upper left corner of an element symbol show?

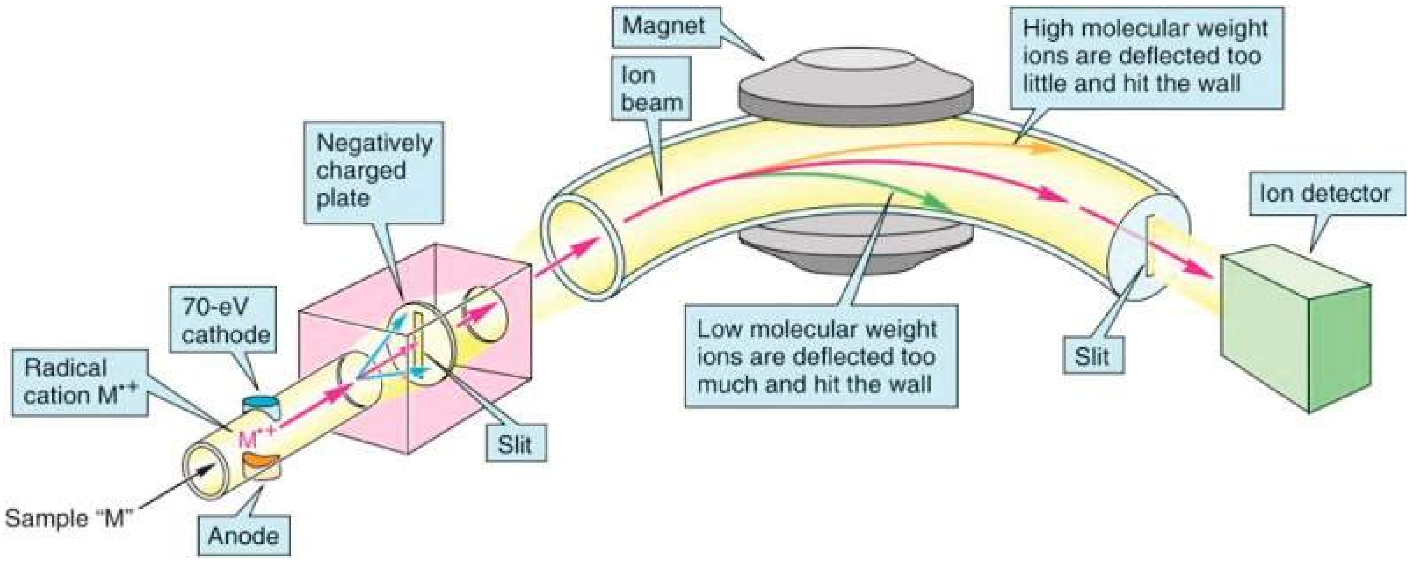
3-Find out number of neutrons (n°) of each element above.

Table 5-2 IMPORTANT ISOTOPES OF SOME ELEMENTS

Elem	Atomic mass of isotopes	Isotopes	percent (%) abundance in nature
Hydrogen	1.0078	${}^{1}_{1}H$	%99.984
	2.0141	${}^{2}_{1}H$	%0.015
	2.0160	${}^{3}_{1}H$	
Carbon	12.000	${}^{12}_{6}C$	%98.893
	13.0033	${}^{13}_{6}C$	%1.1070
	14.0032	${}^{14}_{6}C$	
Nitrogen	14.0031	${}^{14}_{7}N$	%99.634
	15.0001	${}^{15}_{7}N$	%0.366
Oxygen	15.9949	${}^{16}_{8}O$	%99.759
	16.9991	${}^{17}_{8}O$	%0.0674
	17.9992	${}^{18}_{8}O$	%0.0239
Fluorine	18.9984	${}^{19}_{9}F$	%100
Neon	19.9924	${}^{20}_{10}Ne$	%90.48
	20.9937	${}^{21}_{10}Ne$	%0.27
	21.99014	${}^{22}_{10}Ne$	%9.25
Sodium	22.9868	${}^{23}_{11}Na$	%100
Phosphorus	30.9738	${}^{31}_{15}P$	%100
Chlorine	34.9689	${}^{35}_{17}Cl$	%75.53
	36.9659	${}^{37}_{17}Cl$	%24.47
Bromine	78.9184	${}^{79}_{35}Br$	%50.54
	80.9163	${}^{81}_{35}Br$	%49.46
Iodine	126.9045	${}^{127}_{53}I$	%100
Lead	203.9731	${}^{204}_{82}Pb$	%1.48
	205.9745	${}^{206}_{82}Pb$	%23.6
	206.9759	${}^{207}_{82}Pb$	%22.6
	207.9766	${}^{208}_{82}Pb$	%52.3
Thorium	232.0382	${}^{232}_{90}Th$	%100

Carbon isotope $^{12}_{6}C$ was chosen as a reference for most applications. The reason is that its mass of 12 units was measured very precisely with mass spectrometers (Fig 5-2) and by many other instruments. Measurement of other elements were made by taking 1/12 of mass of $^{12}_{6}C$ element as a basis (refer to Chapter 1 for definition of atomic mass unit).

Figure 5-2



Atomic mass of an element can be calculated from isotopic abundances as in the following equation.

$$\text{Atomic mass of an element} = \frac{[(\text{mass of }1^{st}\text{ isotope})(\%\text{ abundance})] + [(\text{mass of }2^{nd}\text{ isotope})(\%\text{abundance})] + [...]}{100}$$

Example 5-1

^{35}Cl makes up 75.53% of all Chlorine in nature. ^{37}Cl makes up 24.47% of it. According to this, calculate atomic mass of chlorine.

Solution:

As chlorine has only two isotopes, the equation is written as follows:

$$\text{Atomic mass of chlorine} = \frac{[(\text{mass of }1^{st}\text{ isotope})(\%\text{ abundance})] + [(\text{mass of }2^{nd}\text{ isotope})(\%\text{abundance})]}{100}$$

$$\text{Atomic mass of chlorine} = \frac{(34.9689 \times 75.53) + (36.9659 \times 24.47)}{100}$$

$$= 35.4576 \text{ amu}$$

Exercise 5-2

Calculate mass of Boron (B) atom which has 81.2% of ^{11}B and 18.8% of ^{10}B isotopes in nature.

Isotopes have several applications in medicine to diagnose and treat many illnesses. For example, Cobalt isotope (^{60}Co) is used in tumor treatments, Iodine isotope (^{131}I) in treatment of swelling of thyroid gland (goiter). (Figure 5-3) Scientists were able to estimate ages of rocks, meteors and fossils by using uranium isotopes $^{238}_{92}U$, $^{235}_{92}U$, $^{234}_{92}U$ and Thorium isotopes $^{232}_{90}Th$. Radioactive isotopes are also used in industry. They are used in measurements of thicknesses of plates, in industry of measurement instruments of liquid and gas flows.

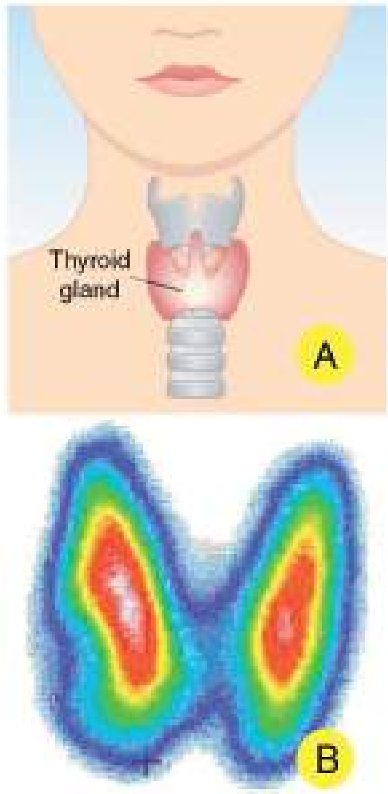
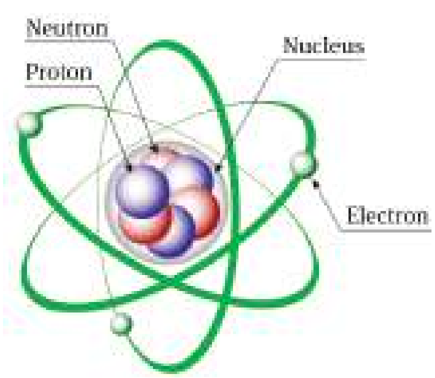


Figure 5-3
A) Thyroid gland
B) X-ray film of thyroid gland



Structure of the atom

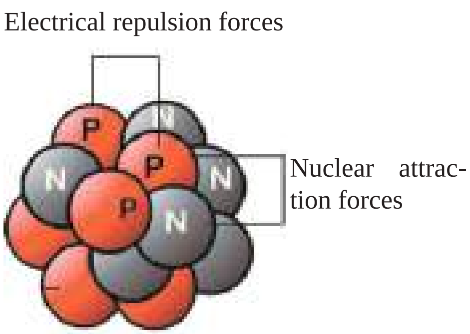


Figure 5-4

Besides, by using radioactive isotopes, leak points in tanks and pipes of dangerous liquids and gases, and also water leakages underground can be detected without a digging process. In agriculture, radioactive isotopes are used as fertilizers and to increase soil productivity.

5-4-VOLUME AND MASS OF NUCLEUS

Although atom is extremely small, it is possible to measure its volume very precisely. Diameter of an atom is equal to 100 millionth of 1 centimeter (1/100000000). This can be imagined as 100 million atoms were placed on the head of a pin which has a diameter of 0.001 centimeter.

Nucleus is accepted as center of gravity of an atom and energy source. Nucleus is even much smaller than atom itself. Volume of nucleus is ten thousandth of the volume of an atom. Cloud around nucleus is made up of orbitals in which electrons turn around . Structure of atom resembles solar system. According to this, while the sun represents nucleus, planets evolving in orbits are similar to electrons. These planets are at a certain distance from the sun. They are bound to solar system with law of attraction. Similarly, electrons are too far away from nucleus relatively but due to law of attraction, nucleus and electrons make up atom as a single body. Table 5-3 shows some properties of main particles of atom we have studied in secondary school.

Table 5-3 Mass of atom, charge, symbol and its components

Particle	Symbol	Type of charge	Mass (g)
Electron	e^-	-1	9.11×10^{-28}
Proton	p^+	+1	1.672×10^{-24}
Neutron	n^0	neutral	1.674×10^{-24}

5-5-NUCLEAR STABILITY

While non-radioactive isotopes are stable, radioactive isotopes aren't. There are 280 isotopes in nature. 50 of those are radioactive. Scientists have produced many isotopes up to 500 artificially. In order to accomplish this, atoms of some elements were bombarded with neutrons.

Reasons of instability of some radioactive isotopes depend on the ratio between numbers of neutrons and protons ($n°/p^+$). This ratio is 1/1 for stable nuclei. In other words, for stable nuclei this ratio is equal to 1 exactly. But if this ratio is greater than 1, nuclei become unstable. Therefore, this kind of nuclei emit some rays to reach stability. This phenomenon is called as **radioactivity.**
The most stable nuclei have a ratio of ($n°/p^+$) close to 1 and also small atomic numbers. As atomic numbers increase, ratio of neutrons to protons becomes greater than 1. The reason is the relationship between nuclear attraction force and static electrical repulsion force of protons. Nuclear attraction force affects protons and neutrons equally and it is accepted as the greatest attraction force in nature. As atomic number increases, also number of protons in nucleus increases. Related to this, static repulsion force between protons also builds up. To overcome nuclear attraction force, more neutrons are needed. (Figure 5-4)

5-6-NUCLEAR BINDING ENERGY

Nucleus of helium atom consists of 2 protons and 2 neutrons:

Mass of proton p^+: (1.00728 amu)

Mass of neutron n^0: (1.00866 amu)

Therefore, mass of nucleus of helium atom can be calculated as follows:

Mass of 2 protons= (1.00728 × 2 = 2.01456 amu)

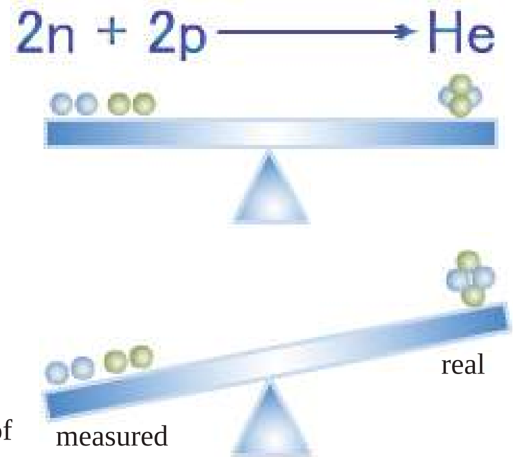
Mass of 2 neutrons= (1.00866 × 2 = 2.01732 amu)

Total mass of protons and neutrons = 2.01732 + 2.01456 = 4.03188 amu

Mass of nucleus of helium: 4.03188 amu

When we compare real mass (4.00151 amu) of helium with total of masses of particles in its nucleus, a mass difference of (0.03037 amu) can be observed.

If a mass spectrometer is used, it can be observed that this difference (lost mass) is transformed into energy according to Einstein's equation ($E=mc^2$). This energy is called as nuclear binding energy. This energy is also the energy which enables to overcome repulsion of positive protons and also keeps protons and neutrons inside tiny nucleus. In order to calculate transformation of mass into energy, we will follow the steps shown in the following example:



The difference between real and measured masses of nucleus of helium atom

Einstein's Equation: $E = mC^2$
E = Energy
m = mass
C = speed of light (3×10^8 m/s)

Example 5-2

If mass difference of nucleus of Helium is accepted as 0.03037 amu, calculate nuclear binding energy of Helium nucleus.

Solution:

We convert mass from amu to kg.

$$m = 0.03037 \text{ amu} \times \frac{1.66 \times 10^{-27} \text{ kg}}{1 \text{ amu}} = 0.050414 \times 10^{-27} \text{ kg}$$

To calculate binding energy, we use Einstein's equation.

$$E = mC^2 = 0.050414 \times 10^{-27} \text{ kg} \times (3 \times 10^8 \text{ m/s})^2$$

$$= 0.454 \times 10^{-11} \text{ kg . m}^2/\text{s}^2$$

As 1 joule (J), kg.m^2/s^2, it is found that $E = 0.454 \times 10^{-11}$ J.

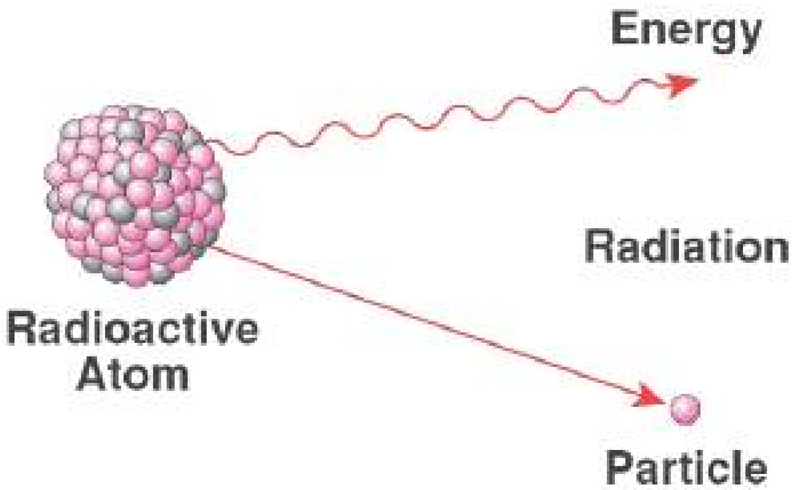
Exercise 5-3

Calculate nuclear binding energy of nucleus of lead element with 82 protons and 125 neutrons. As mass of proton is 1.00728 amu and mass of neutron is 1.00866 amu, mass of lead nucleus is 207.2 amu.

5-7-RADIOACTIVITY

In 1896, scientist Henry Becquerel observed that photograph films were blurred in the drawer of his desk on which he kept some
Uranium samples. He thought some kind of invisible rays emitted by Uranium caused this event by affecting the film. Madame Curie and her husband studied carefully Becquerel's studies in his lab and proposed some proofs on two new elements (Radium and Polonium) they discovered. In both Becquerel and Curie's hands, a weigh-able amount of elements was accumulated.
Radioactivity unit was accepted as Curie to honor studies of Madame Curie.
Some unstable radioactive nuclei decay spontaneously and form new and stable isotopes. This phenomenon is called as radioactivity. During this process, high energy nuclear rays are emitted and as a result, nuclei of some elements become more stable. See Figure 5-6.

Figure 5-6
Radioactivity
The nucleus of a radioactive atom



Marie Curie
(1867-1934)

Pierre Curie
(1859-1906)

Nucleus of Uranium isotope can be given of as an unstable (radioactive) atom. This nucleus undergoes decay. Rate of decay of nucleus depends on components of nucleus and its energy level. According to studies, there are 3 kinds of radiation. Ability of penetration of these radiation types are different (Figure 5-7). These 3 radiations were named after the first 3 letters of Greek alphabet (Alpha α, Beta β, Gamma γ).

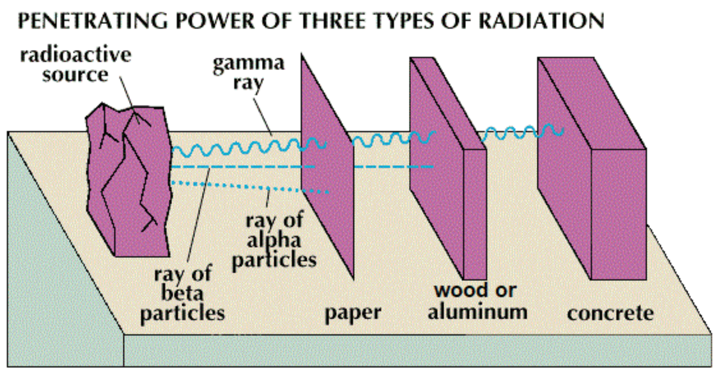


Figure 5-7
Types of radiation and their characteristics

1-Alpha Particles (α)

These are positive-charged particles. As an alpha particle is made up of 2 protons and 2 neutrons, it represents a helium nucleus. It is shown with α and ${}^{4}_{2}He$, it has the largest mass among radiation types. Its properties;

1) Alpha particles have a great effect on matter. Because these particles cause electrons to be freed when they hit substances, they cause them to be ionized.

2) Effect of these particles on matter is very short as alpha particles combine with 2 electrons which have been freed. This occurs through ionization of matter. Alpha particles are transformed into Helium atom according to the following equation.

$${}^{4}_{2}He\ (\alpha) + 2e^{-} \rightarrow {}^{4}_{2}He$$

Alpha particles   Helium gas atom

Disintegration of uranium isotope ${}^{238}_{92}U$ due to radioactivity:
Uranium changes into Thorium isotope ${}^{234}_{90}Th$ and releases alpha particles. Figure 5–8

$$ {}^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He\ (\alpha) $$

Uranium Thorium Alpha particles

Path of alpha particles can be cut by a thin sheet of paper as shown in Figure 5-7.

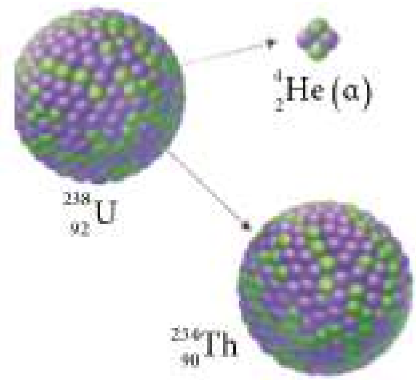


Figure 5-8
Emission of alpha particles from radioactive nucleus of uranium isotopes

2-Beta particles (β)

They are equivalent to electrons. With respect to alpha particles, they penetrate more into matter. Because volumes of alpha particles are bigger than volume of electrons. Therefore, electrons can reach more electron orbitals of atoms. β particles are shown with ${}^{0}_{-1}e$ symbol. Beta radiation can be stopped with a piece of wood Fig. (5-7). The following equation shows emission of beta particles from neutron to produce a proton.

$$ {}^{1}_{0}n \rightarrow {}^{1}_{1}p + {}^{0}_{-1}e $$

For example, carbon isotope ${}^{14}_{6}C$ decays and changes into nitrogen isotope ${}^{14}_{7}N$. Notice that atomic number goes up by 1 whereas mass number remains unchanged. Fig 5-9.

*****While along with emission of beta particle, 1 neutron changes into 1 proton, also 1 electron and 1 anti-electron (neutrino) are released. We will mention this in future.

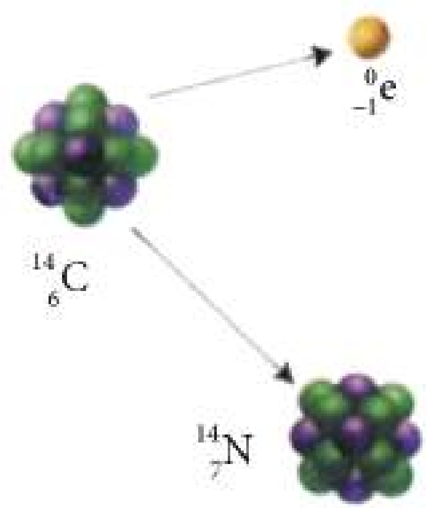


Figure 5-9
Emission of beta particles through disintegration of carbon isotopes ${}^{14}_{6}C$

3-Gamma Rays (γ)

Gamma rays are made up of neutral electromagnetic waves and they have a speed close to speed of light. These are the strongest rays. They are more effective in penetration of matter, stronger and they can travel further. They are also accepted as the most dangerous type of radiation. Path of these rays can only be slowed down by concrete barrier. Figure 5-7. The following equation shows disintegration of uranium isotope ${}^{238}_{92}U$ as a result of alpha and gamma radiation and formation of Thorium isotope.

$$ {}^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He\ (\alpha) + {}^{0}_{0}\gamma $$

Figure 5-10 shows effects of magnetic and electric fields on all types of radiation.

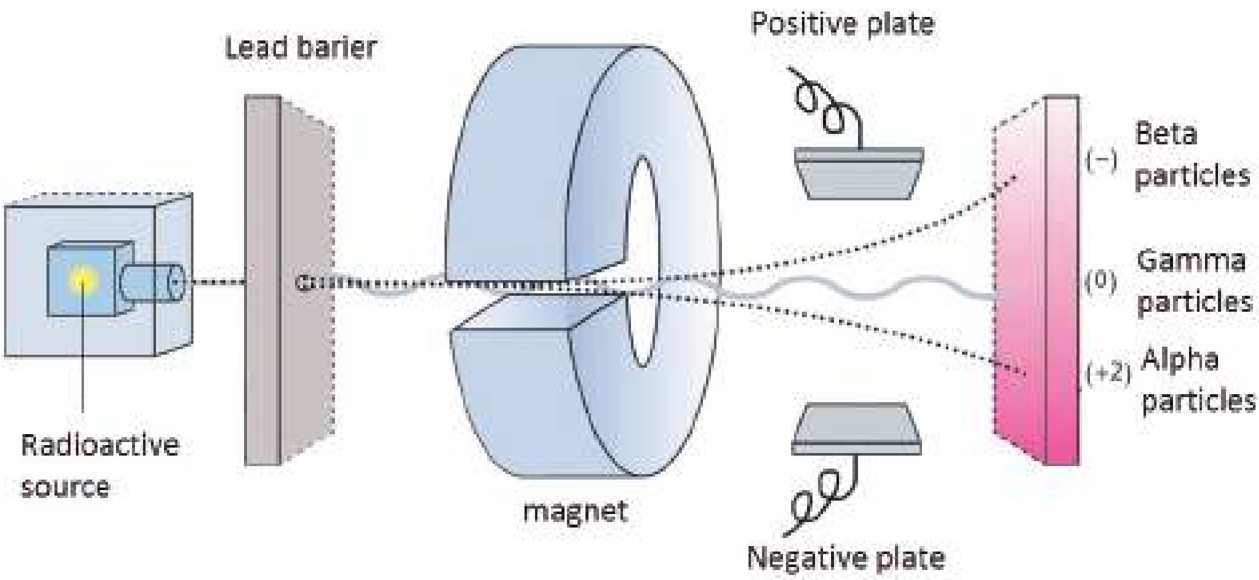


Figure 5-10
Effect of magnetic and electric fields on radioactive particles released from a radioactive source

Table 5-4 Properties of radioactive types

Type of radion	Property	Speed	Charge	Behaviour electric field	Stopping radiation
Alpha (α)	Helyum nucleus ($^{4}_{2}He$)	10% of speed of light	Positve (+2)	It deviates by approaching Negative plate	Papre, clothing
Beta (β)	High speed ($^{0}_{-1}e$)	90% of speed of light	Negative (-1)	It deviates by approaching Positive plate	Wood or Aluminium barrier
Gamma (γ)	High-energy electromag-netic wave	Speed of light	(0)	Not effected	Concrete or 10 cm lead barriers don't stop them but only reduces their effects.

Properties of Radioactive Elements

1- All compounds of radioactive elements are also radioactive.
2- A radioactive element is radioactive in all states (solid, liquid, gas).
3- Nucleus of a radioactive element doesn't emit alpha and beta particles at the same time. It can emit either alpha or beta emission at one time. But it can emit gamma radiation in both cases.
4- Radioactivity of a sample isn't affected from outside factors as pressure and temperature.
5- Emission of alpha or beta particles from nucleus of a radioactive element changes it into another element's nucleus.

5-8-INTENSITY OF RADIATION

It shows number of decays in 1 second. According to this, when it is said that intensity of cobalt source is 50 thousand Becquerel, it means at every second, 50 thousand nuclei decay at this source with Becquerel unit. 1Bq is number of decays at 1 second. Curie unit (Ci) is equal to 37 million Becquerel.

5-9-HALF-LIFE TIME

Its symbol is ($t_{1/2}$). It is the time necessary for decay of half of radioactive material. In other words, it is consumption of half of nuclei of radioactive material. Isotopes of different elements have a constant, known half- life for each. Most radioactive isotopes decay in a few steps (successive radiolysis chain). In the end, radioactive material becomes stable. These are called as **radioisotope elements** while real isotopes before decay are called as **main elements**. Fig. 5-11 shows decay curves of radioactive elements and half-life of each.

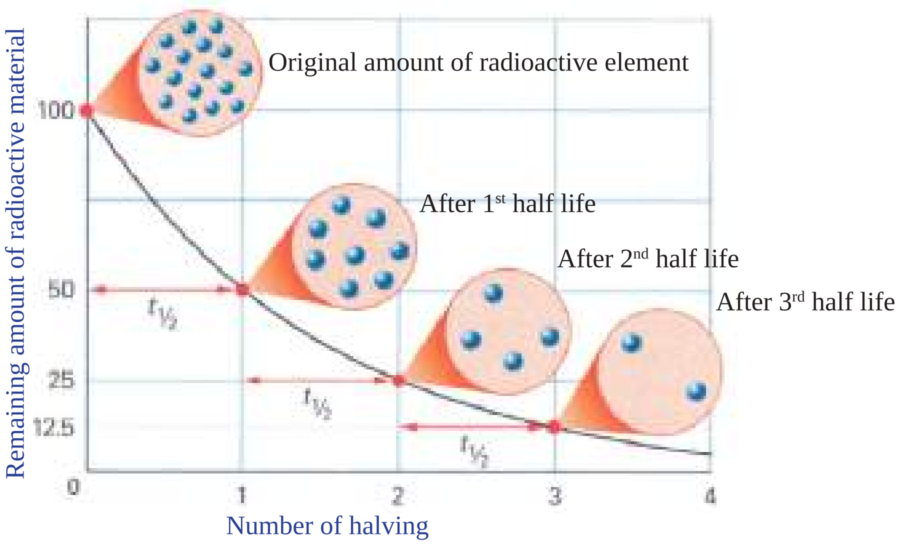


Figure 5-11
Decay curve and half lives of radioactive elements

Most of the time, half lives of isotopes depend on atoms which they are made up of. Outside factors as temperature, pressure, chemical environment, magnetic and electric fields have no effect. Therefore, half-lives of isotopes don't change by time, this is a natural constant value. Each radioactive nucleus has a characteristic half-life. Stable nuclei decay slowly. Therefore, they have a longer half-life. But less stable nuclei decay rapidly. Their half-lives are too short values not longer than a few parts of a second. Half-life of $^{14}_{6}C$ is almost 5730 years. Half -life of Uranium $^{238}_{92}U$ is 4.6 $\times 10^9$ years and half-life of Potassium $^{40}_{19}K$ is 1.3 $\times 10^9$ years. Half-life of Polonium $^{218}_{84}P$ is 3 minutes; of Astatine $^{218}_{85}At$ is too short with 1.6 seconds.

Do you know that?

If we had 1 kg of $^{238}_{92}U$ element, half kg of it would turn into lead $^{206}_{82}Pb$ element after 4.5 billion years passed.

Radioactive elements with short half-lives are used in nuclear medicine in treatment of some illnesses especially against cancer tumors (these tumors cause change in cell structure and fast cell division). These elements don't pose a dangerous radioactivity source for patients. Hal-life times are also used to estimate ages of trees and fossils. Most of carbon dioxide in atmosphere is made up of $^{12}_{6}C$ and a small amount of it is of $^{13}_{6}C$. Both these isotopes aren't radioactive. Additionally, radioactive $^{14}_{6}C$ is found in very small amounts in nature.

Decay time of this isotope ($^{14}_{6}C$)is constant and their amounts remain unchanged. The reason is the effect of cosmic rays on nitrogen $^{14}_{7}N$. $^{14}_{7}N$ is found in atmosphere and by decaying it turns into $^{14}_{6}C$. As everybody knows, plants use carbon dioxide gas in atmosphere during photosynthesis. Ratio of $^{14}_{6}C$ atoms to $^{12}_{6}C$ atoms in plants and carbohydrates is equal to that of atmosphere.

But this ratio is being reduced by disruption of life cycle (by destruction of forests, death of living organisms). As $^{14}_{6}C$ atoms are radioactive and they constantly decay, at the end of 5730 years which is half-life of ratio of $^{14}_{6}C$ to $^{12}_{6}C$ in organisms is half of atmosphere. In order to determine ages of fossils, some of their samples are burnt and carbon dioxide gas is formed, then ratio of $^{14}_{6}C$ to $^{12}_{6}C$ is calculated. From this ratio and by some special calculations, ages of fossils can be estimated. This method is also used in study of archaeological remains.

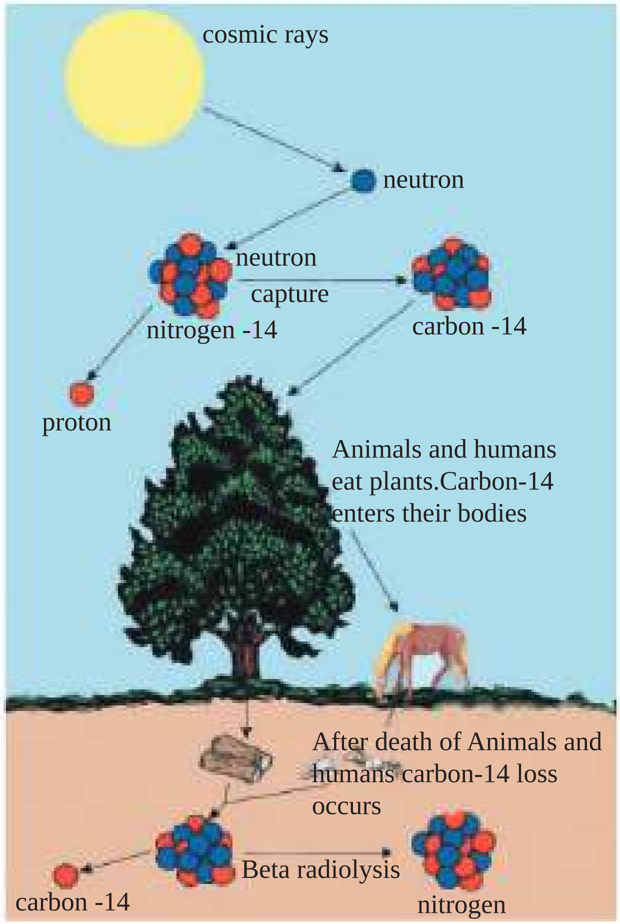


Figure 5-12
Cycle of Carbon-14 isotopes in living organism

Example 5-3

Carbon isotope $^{14}_{6}C$ with a half-life of ($t_{1/2}$) 5730 years decays by emitting beta particles. For an isotope sample with a mass of 2×10^{-2} g, make the following calculations:

1- Find out the time passed in three half-lives.
2- How many grams remain from the isotope at the end of third half-life?

Solution:

This kind of questions are solved by the following formula:

$$N_t = \frac{N_0}{2^{(t/t_{1/2})}}$$

N_o is original amount of matter. N_t is amount of remaining radioactive material after t time and $t_{1/2}$ has passed.

1- In order to calculate three half-lives, half-life time is multiplied with number of half-times:

Half-life time = 3 × 5730 years = 17190 years

2- In order to calculate amount of remaining isotope after third half-life has passed, the following formula is used:

$$N_t = \frac{N_0}{2^{(t/t_{1/2})}} = \frac{2 \times 10^{-2}\ g}{2^{(17190/5730)}} = 0.25 \times 10^{-2}\ g$$

5-10-NUCLEAR REACTIONS

Changes which occur in nucleus and cause it change into another nucleus are called as **nuclear reactions.** For example, release of alpha rays from Uranium isotope causes formation of Thorium isotope . As in chemical reactions, nuclear reactions can also be expressed with mathematical equations.

Nuclear reactions can be expressed as in the following equation:

$$^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He$$

In nuclear reactions, arithmetical sum of atomic numbers and mass numbers must be equal on both sides. Chart 5-5 shows particles released in nuclear equations.

Neutron	$^{1}_{0}n$
Proton	$^{1}_{1}H(p^{+})$
Electron	$^{0}_{-1}e$
Alpha	$^{4}_{2}He$
Beta	$\beta^{-}(^{0}_{-1}e)$
Gamma	$^{0}_{0}\gamma$

Example 5-4

In the following nuclear equation, find out atomic number and mass number of X element.

$$^{212}_{84}Po \rightarrow {}^{4}_{2}He + X$$

Solution:

Mass number of Polonium isotope is 212 and its atomic number is 84. When alpha particles are released, X element is formed as in the equation.

Mass number of element X = 212 - 4 = 208

Atomic number of element X = 84 - 2 = 82

The result

$$^{212}_{84}Po \rightarrow {}^{4}_{2}He + {}^{208}_{82}X$$

Exercise 5-4

1) Manganese isotope ^{56}Mn decay in 2.6 hours by emitting beta rays. How much of 1 g of Manganese-56 sample remains after 10.4 hours passed?

2) Half-life of Phosphorus ^{32}P isotope is 14.3 days. How many grams of 4 g of ^{32}P isotope remain after 57.2 days has passed?

Exercise 5-5

A) In the following nuclear equations, find the name of particle added to $^{22}_{11}Na$ isotope:

$$^{22}_{11}Na + x \rightarrow {}^{22}_{10}Ne$$

B) Find out atomic number and mass number of element X in the following nuclear equation.

$$^{253}_{99}Es + {}^{4}_{2}He \rightarrow {}^{1}_{0}n + x$$

5-11-TYPES OF NUCLEAR REACTIONS

We can divide nuclear reactions into 4 parts:

1- Radioactive spontaneous disintegration

2- Non-spontaneous nuclear reaction

3- Nuclear Fission

4- Nuclear Fusion

5-11-1- Radioactive Spontaneous Disintegration

Nuclei of heavy unstable elements are transformed into lighter and more stable nuclei spontaneously. For this, alpha, beta and gamma particles are released through radiation. For example, transformation of Uranium isotope into Thorium isotope spontaneously and alpha radiation.

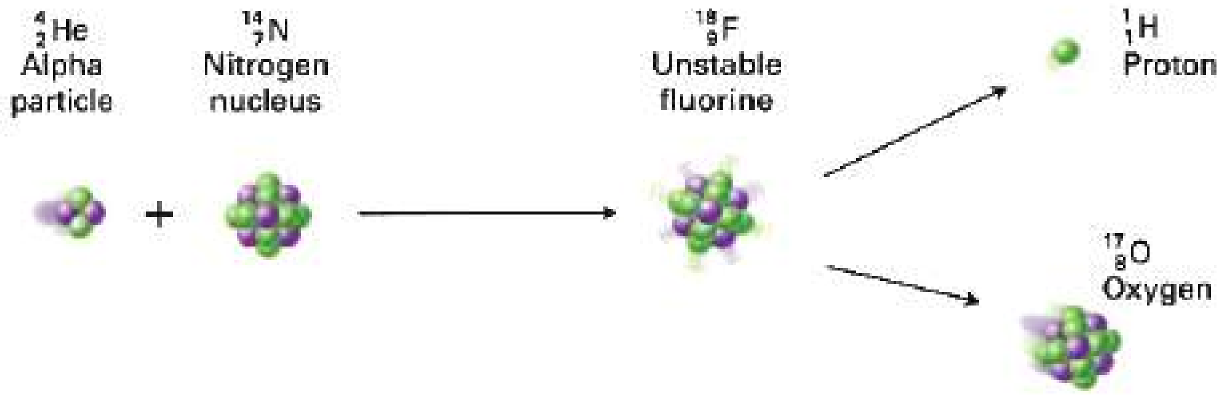
5-11-2- Non-spontaneous Nuclear Reaction

It occurs through bombardment of nucleus by particles or light nuclei.
1) As in the following equation, nucleus is bombarded with neutron (proton emission).

$$^{35}_{17}Cl + ^{1}_{0}n \longrightarrow ^{35}_{16}S + ^{1}_{1}H$$

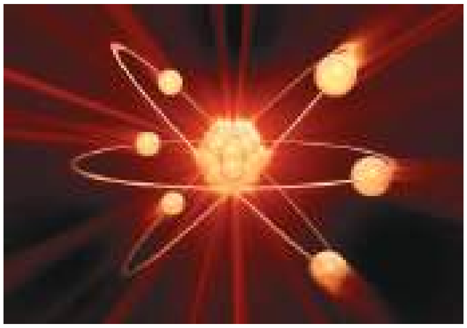
2) As in the following equation, nucleus is bombarded with alpha particles:

$$^{4}_{2}He + ^{14}_{7}N \longrightarrow ^{18}_{9}F$$



$$^{18}_{9}F \longrightarrow ^{17}_{8}O + ^{1}_{1}H$$

5-11-3-Nuclear Fission



By looking at the photo above, can you imagine the extent of the explosion and how it occurred? In 1934, a German scientist discovered that disintegration of Uranium atom occurred very quickly and produced a great amount of energy. This would cause a very big explosion. Nuclear fission is division of a heavy nucleus into two average massed nuclei, formation of new elements and producing a great amount of heat and radiation. As shown in Figure 5-13, Uranium $^{238}_{92}U$ and Plutonium $^{239}_{94}Pu$ isotopes are used as nuclear fission materials. If left uncontrolled, nuclear fission can lead up to great explosions just as in an atomic bomb explosion. The photograph above pictures the great energy after explosion of an atomic bomb.

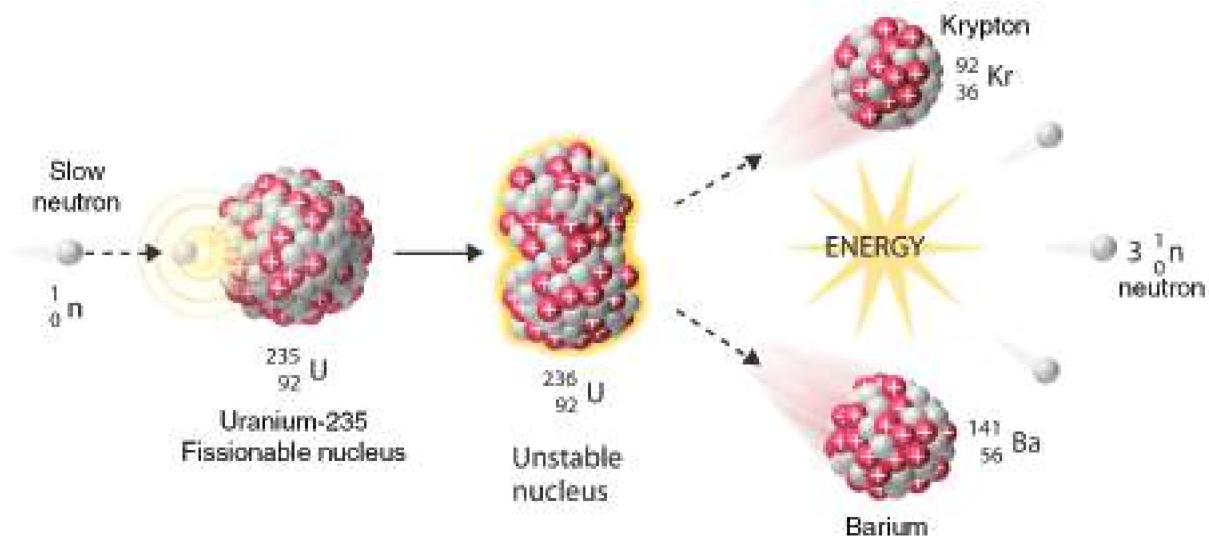


Figure 5-13
Fission of uranium $^{235}_{92}$U nucleus

One of nuclear fission applications is made in nuclear reactors by controlling amount of released energy during fission to produce electric energy. Fig. 5-14 shows a nuclear power plant which produces electric energy.

5-11-4-Nuclear Fusion

We all know the sun supplies necessary energy for life. But how is this energy produced?

Many reactions occur in the sun. One of those reactions is nuclear fusion. As there are fissionable nuclei, some lighter nuclei can combine with each other. Nuclear fusion is formation of heavier nuclei from combination of lighter nuclei. From fusion of hydrogen isotopes Deuterium $^{2}_{1}$D and Tritium, $^{3}_{1}$T Helium atom $^{4}_{2}$He is produced. Meanwhile, a great amount of energy is released as shown in Figure 5-15.

$^{3}_{1}T + ^{2}_{1}D \longrightarrow ^{4}_{2}He + ^{1}_{0}n + \text{Energy}$

Figure 5-14
A nuclear power plant which produces electric energy

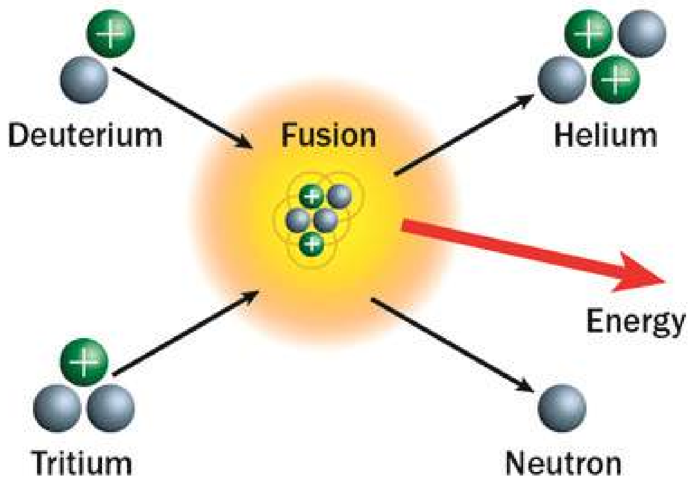
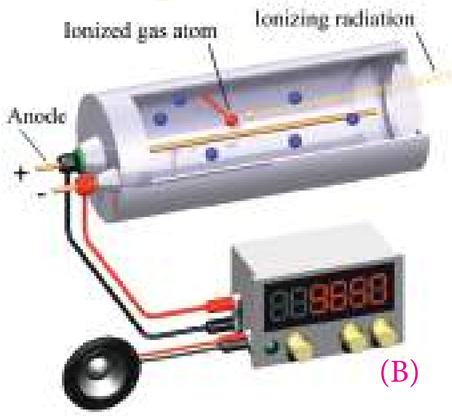


Figure 5-15
Nuclear fusion and formation of nucleus of Helium $^{4}_{2}$He atom


(A)



(B)

Figure 5-16
A) Geiger counter
B) part of Geiger counter

Figure 5-17
Personnel who works at a place with radiation activity is attracting a film badge on his clothes.

In order to accomplish nuclear fusion, a great amount of energy is needed. When this fusion occurs, an immense amount of energy in form of heat and radiation is released as in the sun. The sun supplies us heat, light and life. Otherwise, if this reaction weren't possible, the sun, stars and life wouldn't exist. Energy released from fusion is much greater than energy released from nuclear fission. Hydrogen bomb is an example of nuclear fusion.

5-12-DETECTION OF RADIATION

There are a few methods used to detect activity of radioactive materials.

5-12-1-Geiger Counter

This counter is used to detect radiation activity of radioactive substances. Operation principle of this instrument, shown in Fig. 5-16, depends on ionization of Argon gas in the instrument by high energy rays. Ionization of Argon gas is transformed into electric pulses. These cause a digital counter work or produce audible clicks. This sound points radiation activity from a radioactive substance.

5-12-2-Film Badge

This film is a plastic plate covered with silver bromide (AgBr). This plate is sensitive to amount of rays which pass through it. Amount of rays is determined from the extent of exposure of the film. These special film badges are specially protected. As shown in Fig. 5-17, they are installed on clothes of staff who work at places with radiation activity.

5-13-RADIATION DOSE

It is the amount of radiation absorbed by unit mass of an object in energy units. This amount is measured with Gray (Gy) unit.

1 Gy = 1 J/kg

This amount is measured with (Rad) unit according to another system.

1 Gy = 100 Rad

5-14-IONIZING RADIATION

All types of nuclear radiation have very high energy and they have an ability of ionizing substances they pass through. Therefore, in addition to X-rays, these rays are accepted as ionizing. Ionizing radiation is one type of energy. Their invisibility makes them dangerous. Moreover, people can't feel those as they feel sunlight or heat of fire so they don't try to keep away from them. Ionizing rays penetrate body and transfer their energy to it. With respect to dose which body has absorbed, their harm can continue from a few hours to tens of years. Therefore, protection principles must be considered for X, gamma, cosmic rays and alpha, beta radiation.

5-15-EFFECTS OF IONIZING RADIATION ON MOLECULES IN LIVING BEINGS

1-Physical Risks: They cause leukemia, bone marrow cancer, thyroid, bone cancers and malign tumors. This kind of radiations are accepted as factors that cause shortening life time. Besides, they weaken immune system of people. In case of exposure of embryos to 1-5 Rad of radiation, risk of leukemia after birth is definitely determined.

2-Genetic Risks: Ionizing radiations weakens ability of fertilization and cause infertility along with some genetic deformations. Besides, they can cause genetic mutations. Exposure to ionizing radiation also affects ratio of male/female among newborns.

5-16-RADIOACTIVE DECOMPOSITION OF WATER

Water is accepted as main source of life and as the most abundantly found solvent in nature. From radioactive decomposition of water, negative and positive ions are formed. Later, these ions decompose to high energy ions and free radicals. As a result, these ions and free radicals cause great changes in organic components of cells and in sensitive parts (chromosomes) by combining with them. Harm of radiation in cells and tissues doesn't mean cease of all duties of cells because live tissues and cells have abilities to heal effects of radiation.

5-17-WAYS OF PROTECTING FROM RADIATION

There are three main methods of protecting from ionizing radiation people can face:

1-Time: Harmful effects of radiation increases by time which a person is exposed to radiation.

2-Distance: Amount of radiation a person is exposed decreases as distance between a person and radiation source increases. Safe distance is determined depending on amount of radiation energy and amount of radiation activity of the source.

3-Protective Shield: Effect of radiation decreases as thickness of protective shield around radiation source increases. Thickness of shield depends on type and energy of radiation.

BASIC CONCEPTS

Nuclear Fission: Process of division of heavy nuclei to lighter and more stable nuclei. Formed nuclei have average masses. In this phenomenon, a great amount of energy is released.

Nuclear Reaction: A reaction which affects and changes nucleus of an atom.

Alpha Particle: Positive-charged particles. They are released during decay of radioactive elements. They consist of 2 protons and 2 neutrons and represent nucleus of a Helium atom.

Beta Particle: Negative-charged particles. They are released during some types of radiation.

Nuclear Reactor: Reactors in which nuclear fission reactions are used under control to produce energy from newly formed radioactive nuclei.

Radioactive Nucleus: Unstable nuclei, they release radioactive rays.

Geiger Counter: This counter is used to detect radiation activity of radioactive materials in different fields.

Half-Life Time: It is the time for half of a radioactive substance to decay. In other words, consumption of half of nuclei of radioactive material.

QUESTIONS OF CHAPTER-5

5.1) Half-life of Polonium-210 isotope is 138.4 days. What is the mass (in mg) of remaining Polonium-210 after 415.2 days passed? Initial mass of isotope is 2 mg.

5.2) Half-life of Cobalt-60 isotope is 5.27 years. What is the amount (in mg) of remaining Cobalt-60 with an initial amount of 10mg after 52.7 years passed?

5.3) What is the reason of alpha particles being less effective than beta and gamma rays? (Alpha particles are bigger in charge and mass.)

5.4) Complete the following nuclear equations and balance them. Find out mass number and atomic number of element X in each equation.

1) ${}^{27}_{13}Al + {}^{4}_{2}He \rightarrow {}^{30}_{14}Si + x$

2) ${}^{214}_{83}Bi \rightarrow {}^{4}_{2}He + x$

3) ${}^{27}_{14}Si \rightarrow {}^{0}_{-1}e + x$

4) ${}^{59}_{27}Co + x \rightarrow {}^{60}_{27}Co$

5.5) The following are decayed radioactive isotopes by release of alpha particles. Write down balanced equations for decay of each isotope.

1) ${}^{238}_{94}Pu$

2) ${}^{210}_{83}Bi$

3) ${}^{218}_{84}Po$

4) ${}^{230}_{90}Th$

5) ${}^{222}_{86}Rn$

5.6) As mass of a proton is 1.00728 amu, mass of one neutron is 1.00866 amu and mass of Polonium is 219.213amu, calculate the nuclear binding energy of Polonium ${}^{218}_{84}Po$ nucleus.

5.7) Radioactive isotope of lead element (Pb) decays and yields isotope of Bismuth (Bi) element and releases beta particles. Complete the decaying equation. Calculate atomic number and lost mass.

$${}^{210}_{?}Pb \rightarrow {}^{?}_{83}Bi + {}^{0}_{-1}e$$

5.8) Write down symbol and charge of each of the following:

1- Alpha particles
2- Beta particles
3- Gamma rays

5.9) The following radioactive elements decay by releasing beta particles. Write down decaying process of the following as a balanced nuclear equation:

1- Carbon ${}^{14}_{6}C$
2- Strontium ${}^{90}_{38}Sr$
3- Potassium ${}^{40}_{19}K$
4- Nitrogen ${}^{13}_{7}N$

5.10) How do atomic number and mass number of a nucleus change in case of the following radiations are released?

1- Alpha particles
2- Beta particles
3- Gamma ray

5.11) Tell the difference between radioactive and non-radioactive isotopes.

5.12) Determine which of the following is the most stable isotope.

1) ${}^{14}_{6}C$, ${}^{12}_{6}C$
2) ${}^{1}_{1}H$, ${}^{3}_{1}H$
3) ${}^{18}_{8}O$, ${}^{16}_{8}O$
4) ${}^{15}_{7}N$, ${}^{14}_{7}N$

5.13) What is the reason of using radioactive isotopes with short half-lives in medicine for diagnosis and treatment?

5-14) A patient was given a 20 mg-dose of Iodine-13 (${}^{131}I$). As half-life of this substance is 8 days, how much of it remains in body after 40 days?

5.15) Explain how nuclear fission occurs.

5.16) What is the difference between a nuclear reaction occurring in the sun and a nuclear reaction occurring in a reactor?

5.17) What is the use or purpose of using a film badge while working with ionizing radiation sources?

5.18) Circle the correct answer in the following.

1- If a radioactive element decays with beta radiation: Which of the following statements is true?

A- Atomic number changes.
B- Number of neutrons remains constant.
C- Isotope loses a proton.
D- Mass number changes.

2- Half-life of radioactive element Radon ${}^{222}Rn$ is 3.8 days. How much of 20 g of this element remains after 15.2 days passes?

A- 5.0 g
B- 12.5 g
C- 1.25 g
D- 2.50 g

5.19) What is the particle shown with “?” in the following reaction?

$${}^{27}_{13}Al + {}^{4}_{2}He \rightarrow \ ? + {}^{30}_{15}P$$

5.20) Tell the name of the particle released in each of the following equations.

1) $^{59}_{26}Fe \rightarrow ^{59}_{27}Co + ^{0}_{-1}e$

2) $^{59}_{27}Co + ^{1}_{0}n \rightarrow ^{60}_{27}Co$

5.21) How is heavy water produced?

5.22) What are the differences between chemical and nuclear properties?

5.23) Choose the correct answer in parenthesis in the following questions:

1) In the symbol $^{A}_{Z}X$, what does A show? (Atomic number, mass number, number of neutrons, number of electrons)

2) What does number 238 show in Uranium ^{238}U? (Number of neutrons, mass number, number of protons, atomic number)

3) Which isotope of hydrogen does $^{2}_{1}D$ represent? (normal, the heaviest, heavy, none)

4) How can beta particles be stopped? (by paper, by air, by a piece of wood)

5) In the following nucleus reaction, $^{238}_{92}U \longrightarrow ^{234}_{90}Th + ? + ^{0}_{0}\gamma$ what is the particle shown with "?" ?

6) Half-life of Polonium $^{218}_{84}Po$ is 3 minutes. How much (in grams) of 60 g of ^{218}Po remains after 9 minutes? (60, 7.5, 15, 30)

7) Which of the following occurs when nucleus of an element emit negative beta particle?
A- Mass number remains constant; atomic number decreases,
B- Mass number remains constant; atomic number increases,
C- Mass number remains constant; atomic number remains constant,
D- Mass number decreases; atomic number decreases.

5.24) Explain the reasons of the following.

1) Alpha particles ionize atoms in air when they pass through air. Why?

2) Alpha particles deviate in electric and magnetic fields. Why?

3) Gamma rays aren't influenced by electric and magnetic fields and don't ionize gases. Why not?

4) Effect of gamma rays is much greater than of alpha and beta particles. Why?

5) Positive-charged protons don't go far away from each other although they are inside nucleus. Why not?

6) Free radicals formed from water are dangerous. Why?

5.25) Tell what the following mean.
1- Half-life
2- Radiation activity of a radioactive element
3- Radioisotope element

5.26) Fill in the following blanks.

1- Instruments which detect radiation are called as…….

2- Nucleus of Lead $^{210}_{82}Pb$ turns into Bismuth $^{210}_{83}Bi$ isotope when it emits …..

3- Combination of two lighter nuclei to form a heavy nucleus is called as ….

4- Radiation activity of a radioactive sample isn't influenced by ……. but it only depends on …….

5- Decays occurring in 1 second are called as ……. and they are measured with ……. unit.

6- More stable nuclei ….. slowly and they have …….. half-lives. Less stable nuclei …… quickly and they have ……… half-lives.

5.27) Write down two harms caused by exposure of our bodies to radiation.

5.28) $^{14}_{7}N$ isotope makes up of 99.63% of total nitrogen in nature. For $^{15}_{7}N$, this ratio is 0.37%. According to those, calculate atomic mass of nitrogen.

5.29) What are numbers of protons, neutrons and electrons in each of the following?

1) $^{38}_{19}K$

2) $^{235}_{92}U$

3) $^{68}_{31}Ga$

4) $^{13}_{7}N$

5) $^{59}_{26}Fe$

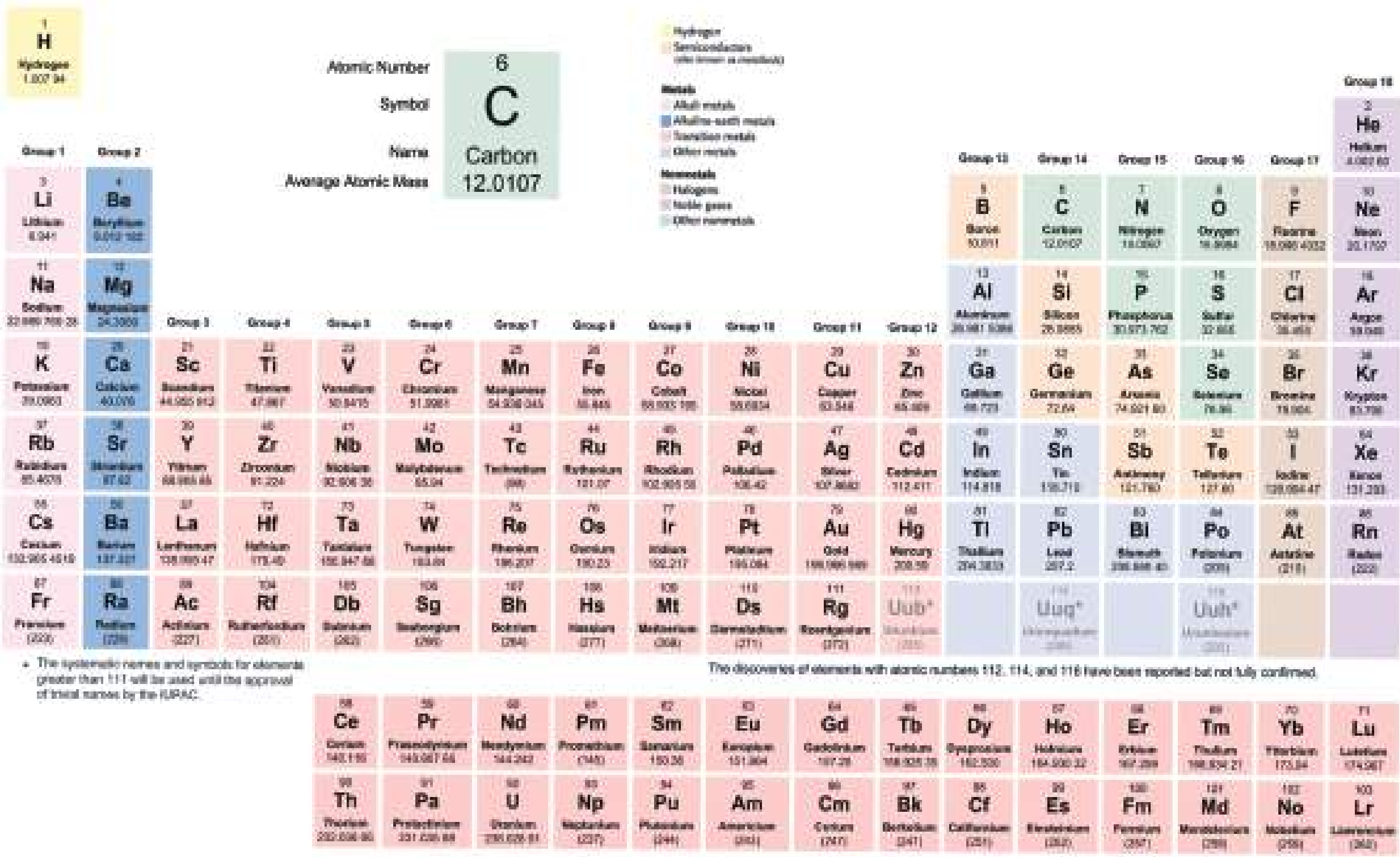
List of Elements

ATOMIC NUMBER	NAME	SYMBOL	RELATIVE ATOMIC MASS	GROUP	PERIOD
1	Hydrogen	H	1.00794	1 / IA	1
2	Helium	He	4.002602	18 / VIIIA	1
3	Lithium	Li	6.941	1 / IA	2
4	Beryllium	Be	9.012182	2 / IIA	2
5	Boron	B	10.811	13 / IIIA	2
6	Carbon	C	12.0107	14 / IVA	2
7	Nitrogen	N	14.0067	15 / VA	2
8	Oxygen	O	15.9994	16 / VIA	2
9	Fluorine	F	18.9984032	17 / VIIA	2
10	Neon	Ne	20.1797	18 / VIIIA	2
11	Sodium (Natrium)	Na	22.98976928	1 / IA	3
12	Magnesium	Mg	24.3050	2 / IIA	3
13	Aluminium (Alumi-num)	Al	26.9815386	13 / IIIA	3
14	Silicon	Si	28.0855	14 / IVA	3
15	Phosphorus	P	30.973762	15 / VA	3
16	Sulfur	S	32.065	16 / VIA	3
17	Chlorine	Cl	35.453	17 / VIIA	3
18	Argon	Ar	39.948	18 / VIIIA	3
19	Potassium (Kalium)	K	39.0983	1 / IA	4
20	Calcium	Ca	40.078	2 / IIA	4
21	Scandium	Sc	44.955912	3 / IIIB	4
22	Titanium	Ti	47.867	4 / IVB	4
23	Vanadium	V	50.9415	5 / VB	4
24	Chromium	Cr	51.9961	6 / VIB	4
25	Manganese	Mn	54.938045	7 / VIIB	4
26	Iron (Ferrum)	Fe	55.845	8 / VIII	4
27	Cobalt	Co	58.933195	9 / VIII	4
28	Nickel	Ni	58.6934	10 / VIII	4
29	Copper (Cuprum)	Cu	63.546	11 / IB	4
30	Zinc	Zn	65.39	12 / IIB	4
31	Gallium	Ga	69.723	13 / IIIA	4
32	Germanium	Ge	72.64	14 / IVA	4
33	Arsenic	As	74.92160	15 / VA	4
34	Selenium	Se	78.96	16 / VIA	4
35	Bromine	Br	79.904	17 / VIIA	4
36	Krypton	Kr	83.798	18 / VIIIA	4
37	Rubidium	Rb	85.4678	1 / IA	5
38	Strontium	Sr	87.62	2 / IIA	5
39	Yttrium	Y	88.90585	3 / IIIB	5

40	Zirconium	Zr	91.224	4 / IVB	5
41	Niobium	Nb	92.906 38	5 / VB	5
42	Molybdenum	Mo	95.94	6 / VIB	5
43	Technetium	Tc	97.9072*	7 / VIIB	5
44	Ruthenium	Ru	101.07	8 / VIII	5
45	Rhodium	Rh	102.905 50	9 / VIII	5
46	Palladium	Pd	106.42	10 / VIII	5
47	Silver (Argentum)	Ag	107.8682	11 / IB	5
48	Cadmium	Cd	112.411	12 / IIB	5
49	Indium	In	114.818	13 / IIIA	5
50	Tin (Stannum)	Sn	118.710	14 / IVA	5
51	Antimony (Stibium)	Sb	121.760	15 / VA	5
52	Tellurium	Te	127.60	16 / VIA	5
53	Iodine	I	126.904 47	17 / VIIA	5
54	Xenon	Xe	131.293	18 / VIIIA	5
55	Caesium (Cesium)	Cs	132.9054519	1 / IA	6
56	Barium	Ba	137.327	2 / IIA	6
57	Lanthanum	La	138.90547	-	6
58	Cerium	Ce	140.116	-	6
59	Praseodymium	Pr	140.90765	-	6
60	Neodymium	Nd	144.242	-	6
61	Promethium	Pm	144.9127	-	6
62	Samarium	Sm	150.36	-	6
63	Europium	Eu	151.964	-	6
64	Gadolinium	Gd	157.25	-	6
65	Terbium	Tb	158.92535	-	6
66	Dysprosium	Dy	162.500	-	6
67	Holmium	Ho	164.930 32	-	6
68	Erbium	Er	167.259	-	6
69	Thulium	Tm	168.93421	-	6
70	Ytterbium	Yb	173.04	-	6
71	Lutetium	Lu	174.967	3 / IIIB	6
72	Hafnium	Hf	178.49	4 / IVB	6
73	Tantalum	Ta	180.94788	5 / VB	6
74	Tungsten (Wolfram)	W	183.84	6 / VIB	6
75	Rhenium	Re	186.207	7 / VIIB	6
76	Osmium	Os	190.23	8 / VIII	6
77	Iridium	Ir	192.217	9 / VIII	6
78	Platinum	Pt	195.084	10 / VIII	6
79	Gold (Aurum)	Au	196.966569	11 / IB	6
80	Mercury (Hydrargy-rum)	Hg	200.59	12 / IIB	6
81	Thallium	Tl	204.3833	13 / IIIA	6

82	Lead (Plumbum)	Pb	207.2	14 / IVA	6
83	Bismuth	Bi	208.98040	15 / VA	6
84	Polonium	Po	208.9824*	16 / VIA	6
85	Astatine	At	209.9871*	17 / VIIA	6
86	Radon	Rn	222.0176*	18 / VIIIA	6
87	Francium	Fr	223.0197*	1 / IA	7
88	Radium	Ra	226.0254*	2 / IIA	7
89	Actinium	Ac	227.0277*	-	7
90	Thorium	Th	232.03806*	-	7
91	Protactinium	Pa	231.03588*	-	7
92	Uranium	U	238.02891	-	7
93	Neptunium	Np	237.0482*	-	7
94	Plutonium	Pu	244.0642*	-	7
95	Americium	Am	243.0614*	-	7
96	Curium	Cm	247.0704*	-	7
97	Berkelium	Bk	247.0703*	-	7
98	Californium	Cf	251.0796*	-	7
99	Einsteinium	Es	252.0830*	-	7
100	Fermium	Fm	257.0951*	-	7
101	Mendelevium	Md	258.0984*	-	7
102	Nobelium	No	259.1010*	-	7
103	Lawrencium	Lr	262.1097*	3 / IIIB	7
104	Rutherfordium	Rf	261.1088*	4 / IVB	7
105	Dubnium	Db	262	5 / VB	7
106	Seaborgium	Sg	266	6 / VIB	7
107	Bohrium	Bh	264	7 / VIIB	7
108	Hassium	Hs	277	8 / VIII	7
109	Meitnerium	Mt	268	9 / VIII	7
110	Darmstadtium	Ds	271	10 / VIII	7
111	Roentgenium	Rg	272	11 / IB	7
112	Ununbium	Uub	285	12 / IIB	7
113	Ununtrium	Uut	284	13 / IIIA	7
114	Ununquadium	Uuq	289	14 / IVA	7
115	Ununpentium	Uup	288	15 / VA	7
116	Ununhexium	Uuh	292	16 / VIA	7
117	Ununseptium	Uus	-	17 / VIIA	7
118	Ununoctium	Uuo	294	18 / VIIIA	7

Periodic Table of the Elements



* The systematic names and symbols for elements greater than 111 will be used until the approval of trivial names by the IUPAC.

The discoveries of elements with atomic numbers 112, 114, and 116 have been reported but not fully confirmed.

The atomic masses listed in this table reflect the precision of current measurements. (Each value listed in parentheses is the mass number of that radioactive element's most stable or most common isotope.)