

بِسْمِ اللَّهِ الرَّحْمَنِ الرَّحِيمِ

(In the Name of Allah, the Most Compassionate, the Most Merciful.)

CHEMISTRY (TECH)

9

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States of Matter and Phase Changes

Student Learning Outcomes

After studying this chapter, students will be able to:

- Define chemistry as the study of matter, its properties, composition, and interactions with other matter and energy. Or Study of earth (solids), Air (gases), Sea (liquids) and Sky (plasma) and their interaction with each other.
- Explain with examples that chemistry has many sub-fields and interdisciplinary fields. Some examples include:
 - Biochemistry
 - Polymer Chemistry
 - Environmental Chemistry
 - Physical Chemistry
 - Inorganic Chemistry
 - Astrochemistry
 - Medicinal Chemistry
 - Geochemistry
 - Analytical Chemistry
 - Organic Chemistry
 - Nuclear Chemistry
- Explain the allotropic forms of solids (some examples may include diamond, graphite, and fullerenes)
- Explain the difference between elements, compounds and mixtures
- Identify solutions, colloids, and suspensions as mixtures and give an example of each
- Explain the effect of temperature on solubility and formation of unsaturated and saturated solutions

1.1 What is Chemistry?

Chemistry is the branch of science which deals with the properties, composition and the structure of substances. It also deals with the physical and chemical changes in matter and the laws or principles which govern these changes.

Determination of composition represents finding out percentages of elements and compounds in a sample of matter. Structure of matter means the arrangement of atoms in matter. Both physical and chemical changes may be brought about by absorption or evolution of energy.

Branches of Chemistry

To understand the widely spread complex subject of chemistry and to concentrate on its specific aspects, chemistry is divided into many distinct branches. These branches have distinct areas of study for the scientists to focus on and to achieve breakthroughs and advancements.

1. Physical Chemistry

This branch investigates how substances behave at atomic and molecular levels. It provides clear explanation as to how fundamental physical laws governing our world cause atoms and molecules to show specific characteristics and in turn react to give huge structures related to life. Physical chemistry is also used to predict and change the rates of reactions and thus optimize the conditions to carry out the reactions on industrial scale.

2. Inorganic Chemistry

It is the study of the synthesis, composition, properties and structure of elements and compounds that contain little or no carbon. An inorganic substance can be composed of metals, nonmetals or a mixture of these, salts, acids and bases. Inorganic compounds are used as fertilizers, medicines, catalysts, pigments, coatings and much more.

3. Organic Chemistry

It is the branch of chemistry that deals with the carbon compounds (hydrocarbons and their derivatives) other than its simple salts like carbonates, bicarbonates, oxides and carbides. In this branch, we study the structure, formation, properties, composition and reactions of carbon containing compounds. Organic compounds are found in all forms of life and are also essential for life.

4. Environmental Chemistry

It is the scientific study of the chemical and biochemical phenomena that occur in this planet. In this subject, we study the sources, reactions, effects and fates of chemical species in the air, soil and water environments. Without this, it would be impossible to study the effects that humans have on the environment through the release of chemicals. It helps in understanding the causes, effects and solutions of different types of pollution.

5. Analytical Chemistry

This branch of chemistry deals with the analysis of different substances. It involves separation, identification and determination of the concentration of the components present in material things. Nowadays the field of analytical chemistry generally involves the use of modern and sophisticated instruments to analyze the matter.

6. Biochemistry

It is the branch of chemistry in which we understand life through chemical processes. It is the study of chemical substances and vital processes occurring in living organisms. Biochemistry provides insights into the structure and function of molecules such as proteins, carbohydrates, lipids and nucleic acids.

7. Nuclear Chemistry

Nuclear chemistry deals with the reactions taking place in the nucleus of an atom. It deals with radioactivity, nuclear processes and transformation in the nuclei of atoms. Nuclear chemistry has many applications in agriculture, medicine, industry and research.

8. Polymer Chemistry

Polymers are large molecules made by linking together a series of building blocks. Polymer chemistry focuses on the properties, structure and synthesis of polymers and macromolecules. Many materials present in the living organisms including proteins, cellulose and nucleic acids are naturally occurring polymers.

9. Geochemistry

The study of chemical composition of Earth and its sources and minerals is called geochemistry. Apart from its use in minerals exploration, geochemical mapping today has applications in environmental monitoring, forestry and medical research.



Interesting Information!

Geothermal heat pump uses a pump to transfer underground water into the buildings during the winter to heat them and in the summer to cool them.

10. Medicinal Chemistry

In this branch of chemistry, the chemist tries to design and synthesize

medicines or drugs which are beneficial for mankind. It includes the discovery, delivery, absorption and metabolism of drugs in human body.

11. Astrochemistry

It is the study of molecules and ions recurring in space and interstellar space. In this discipline we study the abundance and reactions of atoms, molecules and ions in the universe and interaction of these species with radiation.

Exercise

A lunar mission has recently brought samples from the Moon. The following experiments were then carried out on it. Point out the branch of chemistry these experiments are related to.

Experiment	Branch of Chemistry
1. Determining its composition	
2. Studying the physical properties of materials it contains	
3. Carrying out chemical reactions with usual inorganic reagents	

1.2 Differences between Elements, Compounds and Mixtures

Elements	Compounds	Mixtures
An element is the simplest form of matter. It is a pure substance containing the same kind of atoms.	A compound is a pure substance. It is formed by the chemical combination of two or more atoms of different elements.	Mixture is an impure compound. A sample of matter having more than one type of elements or compounds mixed together in any ratio, is called a mixture.
It is not possible to break down an element into simpler particles by ordinary chemical reactions.	In a compound, the atoms of elements must combine together by a fixed ratio by weight. For example, in water (H_2O) hydrogen and oxygen are present in a fixed ratio of 1:8 by weight.	Each component of a mixture retains its identity and specific properties.

When an element exists in the form of atoms, it is represented by a symbol. For example, sodium and calcium are represented by their symbols Na and Ca.	It is possible to break a compound into its constituent elements by a chemical reaction. For example, ammonia can be converted back to nitrogen and hydrogen by a suitable chemical reaction.	A mixture may be homogeneous or heterogeneous. For example, the solution of common salt in water is a homogeneous mixture while a sample of rock is a heterogeneous mixture.
Gaseous elements exist in the form of independent molecules, for example, nitrogen (N_2), oxygen (O_2) and chlorine (Cl_2). Noble gases, however, exist as mono atomic molecules.	The properties of a compound are always different from the elements from which it is formed. For example, the properties of water are different from those of hydrogen and oxygen.	The components of a mixture are not chemically bound together and they can be separated by physical methods.
For example, helium (He) and argon (Ar).	Compounds exist in the form of molecules, for example hydrogen chloride (HCl), ammonia (NH_3) and water (H_2O). Compounds may also exist as network arrangement of their atoms. For example, ionic compounds like NaCl and covalent compounds like sand (SiO_2).	The properties of a mixture are the sum of those of its components.

Exercise

1. Which elements are found in pure state on Earth?
2. Which elements are present in very small amounts on Earth?

1.3 Allotropic Forms of Substances

Elements may exist in more than one structural forms which can exhibit quite different physical and chemical properties. These forms are called allotropic forms and phenomenon is called allotropy. Element oxygen exists in two allotropic forms namely oxygen (O_2) and ozone (O_3).

Similarly, carbon exists in three main allotropic forms, diamond, graphite and Buckminster fullerene. Diamond has a giant macromolecular structure whereas graphite has a layered structure of hexagonal rings of carbon. Buckminster fullerene (C_{60}) consists of spheres made of atoms arranged in pentagons and hexagons. Fullerenes are stable at high temperatures and high pressures. Being covalent in nature, they are soluble in organic solvents. The fullerene structure is unique in that the molecule is not charged, has no boundaries and has no unpaired electrons. They have a cage like structure. Fullerene C_{60} has a low melting point. It is soft and cannot conduct electricity. Element sulphur also exists in two crystalline allotropic forms i.e. rhombic and monocline; the former is more stable than the latter, Fig (1.1).

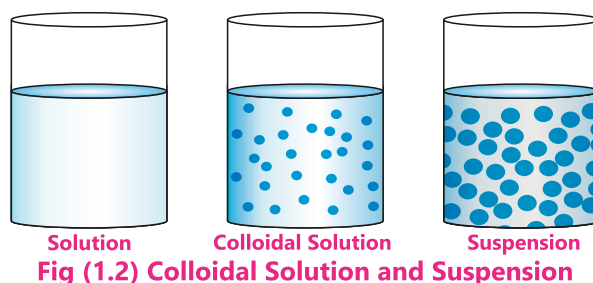


Fig (1.1) Allotropic Forms of Carbon and sulphur

1.4 Solution, Colloidal Solution and Suspension

A solution is such a mixture in which solute particles are completely homogenized in the solvent e.g. dissolution of sodium chloride or copper sulphate in water. The solute particles in such a solution cannot be seen by the naked eye. If solution is filtered, these particles pass through the pores of filter paper leaving no residue. Such a solution is called a true solution. A suspension, on the other hand, is a mixture in which solute particles do not dissolve in the solvent. We can actually see these particles. If a suspension is kept for some time,

these particles settle down. Again, if this suspension is filtered, the particles in it cannot pass through the pores of filter paper and can be collected as a residue. A mixture of chalk in water is an example of suspension. Fig (1.2).



Besides a true solution and a suspension, there is a third form of solution called colloidal solution. In this type, the solute particles do not homogenize with solvent. These particles are a little bit bigger than the solute particles present in the true solution but not big enough to be seen with a naked eye like the particles present in a suspension. If kept for some time, the particles of a colloidal solution do not settle down. On filtration, these particles pass through the filter paper like the particles of a true solution. Starch solution and white of an egg are the common examples of colloidal solutions.



Interesting Information!

Mixtures are closely related to our everyday lives. The air we breathe, the foods we generally consume, the fluids in our body, the solids like steel we use, are all either homogeneous or heterogeneous mixtures.

1.5 Effect of Temperature on the Solubility of Solutes

The solubility of a solute is the amount of solute which can dissolve in 100g of a solvent at a particular temperature.

Change in temperature has different effects on the solubility of different compounds. Usually the solubility increases with the increase in temperature but it cannot be taken as a general rule. There are a large number of compounds whose solubility in water increases with the increase in temperature e.g. potassium nitrate (KNO_3), silver nitrate (AgNO_3) and potassium chloride (KCl) etc. The solubility of sodium chloride in water does not increase appreciably with the increase in temperature. The solubility of compounds like lithium carbonate (Li_2CO_3) and calcium chromate (CaCrO_4) decreases with the increase in temperature. The solubility of gases in water also decreases with the increase in temperature, Fig (1.6).

Similarly, the solubilities of copper sulphate and sodium nitrate also increase with increase in the temperature. However, the solubility of calcium hydroxide decreases with the increase in temperature.

Exercise



1. Tick (✓) the correct answer.

- (i) Hazardous effects of shopping bags are studied in:
- (a) Geochemistry (b) Inorganic Chemistry
(c) Analytical Chemistry (d) Environmental Chemistry
- (ii) The man-made polymer is:
- (a) Starch (b) Polystyrene (c) Protein (d) Cellulose
- (iii) The crystals of which substance has rhombic shape?
- (a) Brass (b) Sulphur (c) Graphite (d) Bronze
- (iv) Which liquid among the following is a colloidal solution?
- (a) Milk (b) Slaked lime used for white wash
(c) Vinegar solution (d) Mixture of AgCl in water
- (v) Which of the following is a heterogeneous mixture?
- (a) A solution of calcium hydroxide in water
(b) A solution of potassium nitrate in water
(c) Hot chocolate
(d) Concrete mixture
- (vi) When the tiny visible particles of a substance are dispersed through a medium, the mixture is named as:
- (a) True solution (b) Colloid
(c) Suspension (d) Saturated solution
- (vii) A solution of KClO_3 has a solubility of about $13.2\text{g per } 100\text{ cm}^3$ at 40°C . How its solubility will be affected, if you decrease the temperature?
- (a) The solubility will increase
(b) The solubility will decrease
(c) The solubility will remain the same
(d) The solubility will first increase with temperature and then it will decrease
- (viii) You are studying the rate of hydrolysis of starch under different conditions of temperature. In which branch of chemistry this topic will fall?
- (a) Organic Chemistry (b) Analytical Chemistry
(c) Biochemistry (d) Physical Chemistry

2. Questions for Short Answers

- i. Why is there a need to divide Chemistry into many branches? Give three reasons.
- ii. Reactions may take place due to electrons present outside the nucleus or they may take place inside the nucleus. Which branches of Chemistry cover these two types of reactions?
- iii. What types of problems are solved in analytical chemistry?
- iv. Both graphite and graphene have hexagonal layered structures. What is the difference?
- v. What is the importance of graphene?
- vi. Which form of matter do most of the material things in this world belong to?

3. Constructed Response Questions

- i. Give the reason of brilliance shown by diamond. Can you improve it?
- ii. Why do different compounds have different solubilities in water at a particular temperature?
- iii. Why NaCl can not be crystallized from water just like KNO_3 ?
- iv. Why graphite is slippery to touch? Which property of graphite enables it to be used as lubricant?

4. Descriptive Questions

- i. Mention the name of the branch of Chemistry in which you will study each of the following topics.
 - (a) Rate of a reaction
 - (b) Digestion of food in human body
 - (c) Properties of colloidal solution
 - (d) Ecosystem
 - (e) Reactions taking place during fireworks
 - (f) Measurement of the absorption of wavelength with the help of ultraviolet spectrometer
- ii. What are allotropic forms? Explain the allotropic forms of carbon and sulphur. How does coal differ from diamond?
- iii. Define solubility of a solute. How does the solubility of solutes change with the increase in temperature?
- iv. Differentiate between the areas which are studied under inorganic and organic chemistry.

Atomic Structure

Student Learning Outcomes

After studying this chapter, students will be able to:

- Explain the structure of the atom as a central nucleus containing neutrons and protons surrounded by electrons in shells
- State that, orbits (shells) are energy levels of electrons and a larger shell implies higher energy and greater average distance from nucleus
- State that electrons are quantum particles with probabilistic paths whose exact paths and locations cannot be mapped (with reference to the uncertainty principle)
- Explain that a nucleus is made up of protons and neutrons held together by strong nuclear force
- Explain that an atomic model is an aid to understand the structure of an atom.
- State the relative charge and relative masses of a subatomic particles (an electron, proton and neutron)
- Interpret the relationship between a subatomic particle, their mass and charge.
- Illustrate the path that positively and negatively charged particles would take under the influence of a Uniform Electric Field.
- Define proton number / atomic number as the number of protons in the nucleus of an atom.
- Explain that the proton number is unique to each element and used to arrange elements in periodic table
- Define nucleon number / atomic mass as sum of number of protons and neutrons in the nucleus of an atom.
- Define relative atomic mass as the average mass of isotopes of an element compared to 1/12th of mass of an atom of Carbon-12
- Interpret and use the symbols for atoms and ions

Elements are very different from one another. A large number of elements exist as solids and very few are present as liquids while the rest exist as gases. All these elements are however made up of atoms.

Have you ever thought why these elements are so different from one another? Sulphur looks very different from gold which, in turn, is very different

from bromine (Fig: 2.1). Similarly iron is a heavy metal while aluminium and zinc are light metals. Metals are mostly lustrous while non-metals like sulphur and carbon appear dull. The difference in the properties of elements is due to the difference in the properties of their constituent atoms.



Fig (2.1) Different Elements look different

Write down the names of the elements shown in the above pictures.

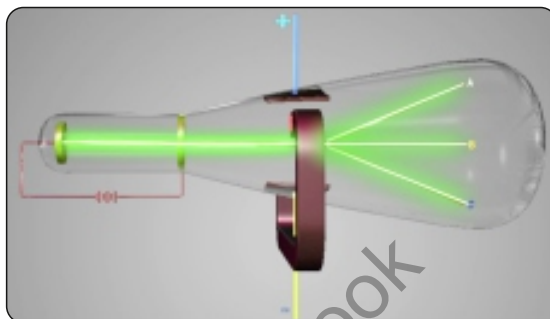
2.1 Structure of Atom

The idea of atom was first proposed in Greece when the philosopher Democritus declared that all matter is made of tiny particles. He named this particle as atom, a particle that cannot be further subdivided. Around 1800, the English chemist John Dalton did many experiments with compounds that provided the evidence for the existence of atoms. In the last decade of nineteenth century, a group of scientists were trying to pass electricity through gases at reduced pressure. During the course of these experiments known as 'Discharge Tube Experiments', they discovered that atoms are no longer the smallest particles of matter; rather there exist particles that are even smaller than atoms. In other words, atoms are composed of negatively charged particles called electrons and positively charged particles called protons. It was also discovered that a proton is 1836 times heavier than an electron. In a discharge tube, the presence of the

negatively charged particles was ascertained because of their deflection towards the positive plate in an electric field. Similarly, the presence of positively charged particles was confirmed due to their deflection towards the negative plate.

Discovery of Electrons

A discharge tube is a hard glass tube provided with two metallic electrodes and a vacuum pump to evacuate the gas present in it Fig (2.2). When a very high voltage is applied to a gas at a very low pressure present in a glass tube the glass surface behind the positive electrode started to glow, due to the rays emitted from the cathode.



**Fig (2.2) Discharge Tube Experiment
Cathode Rays**

These rays were named as cathode rays. In 1897, British physicist Joseph John Thomson studied the properties of cathode rays by passing them through the oppositely charged electric plates. It was observed that cathode rays bent towards the positively charged plate showing that they carry negative charge. Thomson also installed two magnets on either side of the discharge tube and noticed that cathode rays were also diverted by the magnetic field. Thomson used the findings of his experiments to calculate the mass to charge ratio of cathode rays which finally proved that cathode rays are in fact, negatively charged material particles. These particles were later named as electrons. It was also shown that electrons are the subatomic particles of all elements.

Discovery of Proton

The presence of positively charged particles in an atom had been first observed by E. Goldstein in 1886. It was based on the concept that atoms are electrically neutral having same number of positive and negative charges. He performed a series of experiments with a gas-discharge tube having a perforated cathode. A new type of rays were produced from the anode which moved towards the cathode. He called these new rays as canal rays or anode rays.

The properties of these rays seemed to vary depending on the gas

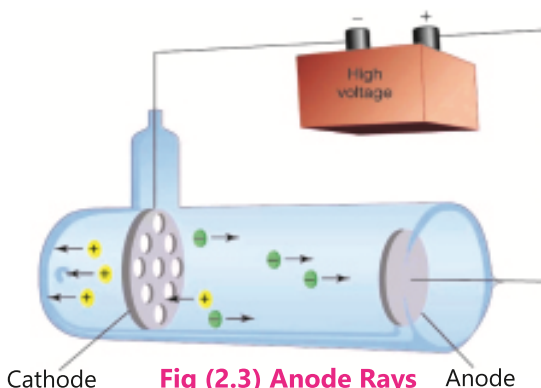


Fig (2.3) Anode Rays

used in the discharge tube. In fact what he discovered was gas ions and this also included hydrogen ions (H^+). Goldstein at that time knew nothing about its significance.

In 1917, Rutherford performed experiments which proved that the hydrogen nucleus is present in other nuclei. Rutherford thought that a hydrogen nucleus or a proton must be the fundamental building block of all nuclei and also possibly a new fundamental particle as well.

Discovery of Neutron

Later on, in 1933, another particle neutron was also discovered, which is known to carry no charge. The mass of a neutron is almost the same as that of a proton.

These three particles i.e. electron, proton and neutron were given the name fundamental particles and are shown to be present in all atoms (except ordinary hydrogen which does not have a neutron) irrespective of the fact that these atoms behave very different from one another. It was, however, also shown that the number of these particles is different in different atoms.

After the discovery of these particles, a very important question arose as to how these particles are arranged in a tiny place called atom. In other words, what is the structure of an atom?

Discovery of Nucleus

Lord Rutherford, in 1911, provided answer to this question. He carried out a remarkable experiment in which he hit a stream of special type of particles to a



Interesting Information!

Although the nucleus is less than one hundred-Thousandth ($1/100,000$) of the size of the atom, it contains more than 99.9% of the mass of the atom.

very thin gold foil. From this experiment he concluded that an atom has two portions. A tiny central portion which he called as nucleus and a relatively large area surrounding this, which he called extra nuclear portion.

It was also discovered that almost all the mass of an atom is concentrated in the nucleus because both the heavy particles i.e. protons and neutrons are found to be present here. In the nucleus these two particles are held together by a strong nuclear force.

Particle	Charge	Mass
Electron	$-1.6022 \times 10^{-19} \text{ C}$	$9.109 \times 10^{-31} \text{ Kg}$
Proton	$+1.6022 \times 10^{-19} \text{ C}$	$1.673 \times 10^{-27} \text{ Kg}$
Neutron	0.0	$1.675 \times 10^{-27} \text{ Kg}$

Bohr's Atomic Model

In 1913, Niels Bohr proposed a model for the hydrogen atom which is called Bohr's atomic model. According to this model the electron can revolve around the nucleus of the atom in specific paths called orbits or shells. When the electron is revolving in one of these orbits, its energy is fixed. Also, when the electron is present in the orbit which is closest to the nucleus, its energy is minimum and it is called the ground state of the atom. The orbits that are further away from the nucleus possess successively greater energy. The electron is not allowed to occupy a space in between the orbits.

Since electron present in each shell has a fixed energy, these shells are also named as energy levels. The shell which is nearest to the nucleus is called first shell or K shell and the electron present in it has a fixed value of energy. Similarly, the second shell will also be at a definite distance from the nucleus which will, of course, be greater than the first shell and electron present in it will also possess greater value of energy. Similarly, electron may also be present in the third or higher shells.

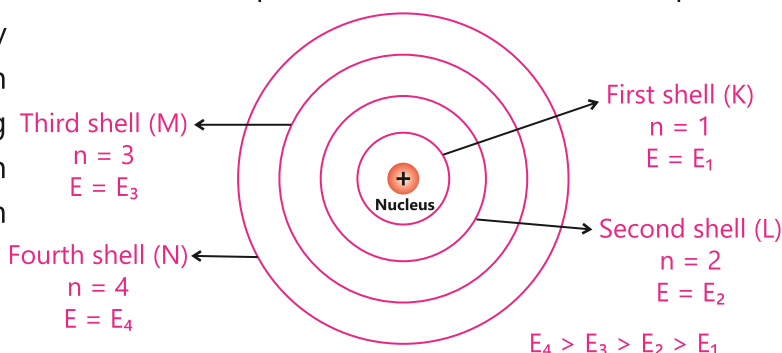
Each shell is further sub-divided into sub-shells or orbitals. The number of sub-shells present in a shell is equal to the value of n for that shell. For the first shell, ($n = 1$), it will, therefore, have only one sub-shell which is called **s** sub-shell. For ($n = 2$), there will be two sub-shells **s** and **p**. The second shell will, therefore, have two sets of sub-shells. The third shell ($n = 3$) has three sub-shells **s**, **p** and **d**. The fourth shell consists of four sub-shells **s**, **p**, **d** and **f**.



Interesting Information!

The size of an atom is so small that it is not possible to see it with naked eye. However, a transition electron microscope can be used to see atoms.

According to modern approach, electrons are like charged clouds whose location around the nucleus cannot be predicted with one hundred percent certainty. All we can say is that there is a certain probability of finding the electron at a certain probable distance from the nucleus.





Interesting Information!

The largest atom cesium is approximately nine times bigger than the smallest atom helium.

In order to find out the number of electrons which can be accommodated in these extra nuclear shells, the scientists have devised a formula called $(2n^2)$ formula where n can have values 1, 2, 3 ... and so on and they represent the number of shells. Shells have also been named as **K, L, M, N** ... and so on. When the value of n is one, it means first or K shell and the maximum number of electrons which can be present in this shell is $(2 \times 1^2 = 2)$. K shell can, however, have less than two electrons or it may not have any electron.

Similarly when $(n = 2)$, it means second or L shell and the maximum number of electrons it can accommodate is $(2 \times 2^2 = 8)$. For $(n = 3)$ or M shell, it can accommodate 18 electrons at the most. This process will go on until the electrons present in an atom are finished.

The maximum number of electrons which can be accommodate in sub-shells **s, p, d** and **f** are 2, 6, 10 and 14 respectively. In the first shell there are 2 electrons which shall go to (**s**) sub-shell. In the second shell, 8 electrons will be further sub-divided, s-subshell will have 2 electrons and 6 electrons will be accommodated in p-subshell.

2.2 Atomic Number and Mass Number

Electrons, protons and neutrons are called the fundamental particles of all types of matter. In other words, the atoms of all the elements present in this world contain same electrons, protons and neutrons. However, an atom of one element differs from an atom of another element because it contains different number of the fundamental particles.

The number of protons present in the atoms of an element is always fixed and it is called the atomic number of that element. Since an atom, as a whole, is electrically neutral, the number of electrons present in an atom will be the same as the number of protons. Atomic number of an element is represented by **Z**. Atomic number of an element is unique to that element and the element is identified by this number. In a periodic table of elements, the elements are arranged according to ascending order of their atomic numbers.

The total number of protons and neutrons present in an atom almost accounts for the total mass of that atom and hence it is called its nucleon number or mass number. It is represented by **A**. The mass of electron being very small is

not included in the mass number.

Just like atomic number, the atom of an element may also be identified from its mass number. For example, the number of protons present in an oxygen atom is 8, so its atomic number is 8 while the total number of protons and neutrons present in it is 16, so its mass number is 16. Information about the atomic number and the mass number is often included with the symbol of any element. The atomic number is written as a left subscript while the mass number as a left superscript. So, oxygen atom would be symbolized as $^{16}_8\text{O}$. Similarly, carbon atom symbolized as $^{12}_6\text{C}$ will have 6 protons and 6 neutrons.

The number of neutrons **N** present in an atom can be calculated if its atomic number **Z** and mass number (**A**) are known.

$$N = A - Z$$

Thus, the number of neutrons in sulphur atom, symbolized as $^{32}_{16}\text{S}$, can be calculated as $32 - 16 = 16$.



Interesting Information!

Changing the number of electrons of an element forms ions, while changing the number of neutrons of an element forms isotopes.

Sample Example:

Calculate the number of neutrons, protons and electrons in barium $^{137}_{56}\text{Ba}$

Solution: Each barium atom will have 56 protons and 56 electrons. The number of neutrons in barium will be calculated as follows:

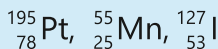
$$N = A - Z$$

$$N = 137 - 56 = 81$$

So each $^{137}_{56}\text{Ba}$ atom will have 81 neutrons, 56 protons and 56 electrons.

Exercise

Calculate the number of neutrons, protons and electrons in the following atoms.



2.3 Relative Atomic Mass

Ever since the existence of atom was recognized, the chemists were trying to find out a method which would allow them to compare the masses of different atoms. It was necessary because without knowing the relative masses of atoms,

we would not know in which ratio of masses we mix the reactants to carry out a chemical reaction. In 1961, the chemists adopted a new scale for the measurement of the relative masses of atoms. The unit mass on this scale is $1/12^{\text{th}}$ of the mass of lighter isotope of carbon taken as 12. The mass of one atom of carbon on this scale is exactly 12 and the masses of the other atoms are measured relative to this unit. The relative atomic mass of an element is thus defined as the mass of an atom of that element relative to the mass of light isotope of carbon taken as 12. The relative atomic masses of elements are expressed in atomic mass unit (amu). It is defined as one-twelfth the mass of an atom of carbon-12.

$$1\text{amu} = 1.67377 \times 10^{-27} \text{ kg}$$

For example, the mass of one atom of hydrogen-1 on atomic mass scale is 1.007 amu, similarly the mass of one atom of sulphur-32 is 31.972 amu.

Exercise

How would you compare the masses of the atoms of Mg and Cl?

Exercise



1. Tick (✓) the correct answer.

- (i) How many electrons can be accommodated at the most in the third shell of the elements?
(a) 8 (b) 18 (c) 10 (d) 32
- (ii) What information was obtained from discharge tube experiments?
(a) Structure of atom was discovered.
(b) Neutrons and protons were discovered.
(c) Electrons and protons were discovered.
(d) Presence of nucleus in an atom was discovered.
- (iii) What does keep the particles present in the nucleus intact?
(a) Particles are held together by strong nuclear force.
(b) Particles are held together by weak nuclear force.
(c) Particles are held together by electrostatic force.
(d) Particles are held together by dipolar force.
- (iv) How do electrons keep themselves away from the oppositely charged nucleus?
(a) By keeping themselves stationary
(b) By revolving around the nucleus
(c) Due to their wave-like nature
(d) A magnetic field around the nucleus keeps them away

2. Questions for Short Answers

- i. Why is it said that almost all the mass of an atom is concentrated in its nucleus?
- ii. Why are elements different from one another?
- iii. How many neutrons are present in ${}_{83}^{210}\text{Bi}$?
- iv. How can an atom absorb and involve energy?

3. Constructed Response Questions

- i. Why does the energy of electron increase as we move from first shell to second shell?
- ii. Why is it needed to lower the pressure of the gas inside the discharge tube?
- iii. During discharge tube experiments, how did the scientists conclude that the same type of electrons and protons are present in all the elements?

4. Descriptive Questions

- i. Describe how electron and proton were discovered?
- ii. Explain the structure of a hydrogen atom.
- iii. Explain the concepts of atomic number, mass number and relative atomic mass.

Chapter 3

Chemical Bonding

Student Learning Outcomes

After studying this chapter, students will be able to:

- Describe that noble gas electronic configuration, octet and duplet rules help predict chemical properties of main group elements
- Compare between the formation of cations and anions
- Account for the electropositive and electronegative nature of metals and non-metals.
- Define ionic, covalent, coordinate covalent and metallic bonds
- Differentiate between ionic compounds and covalent compounds. (The following points need to be included in the respective definitions:
 - a.** Ionic Bond as strong electrostatic attraction between oppositely charged ions
 - b.** Covalent bond as strong electrostatic attraction between shared electrons and two nuclei
 - c.** Metallic bond as strong electrostatic attraction between cloud/sea of delocalized electrons and positively charged cations)
- Explain the properties of compounds in terms of bonding and structure
- Compare uses and properties of materials such as strength and conductivity as determined by the type of chemical bond present between their atoms.
- Justify the availability of free charged particles (electrons or ions) for conduction of electricity in ionic compounds (solid and molten) covalent compounds and metallic bonds.
- Recognize that some substances can ionize when dissolved in water. (e.g. acids dissolves in water and conduct electricity)
- Justify the suitability of usage of graphite, diamond and metals for industrial purposes (Some examples may include: **a.** graphite as lubricant or an electrode **b.** diamond in cutting tools **c.** metals for wires, and sheets)
- Draw the structure of ionic and covalent compounds along with their formation. {some examples can include: **a.** ionic bonds in binary compounds such as NaBr, NaF, CaCl_2 using dot-and-cross diagrams and Lewisdot structures simple molecules including H_2 , O_2 , N_2 , H_2O , CH_4 , NH_3 , HCl , C_2H_4 , CO_2 , and similar molecules using dot and-cross diagrams and Lewis-dot structures)

3.1 Why do atoms form chemical bonds?

Atoms have a tendency to decrease their energy. They can do this by combining with other atoms. It is a natural phenomenon because it increases the stability of atoms.

How do atoms succeed in lowering their energy? The early chemists had started thinking about this a long time ago. They finally succeeded to get an answer only when the noble gases He, Ne, Ar, Kr and Xe were discovered. Helium has two electrons in its outermost shell while all other noble gases have eight electrons in their outermost shells. We also know about these gases that neither their atoms combine with themselves nor with other atoms. The probable reason for this lack of reactivity was their stability. It was suggested that these gases were stable due to the presence of two electrons in helium and eight electrons in the outermost shells of the rest of the gases. This gave rise to a principle that having two electrons (for hydrogen and helium which have only the first shell) or eight electrons in the outermost shell meant stability and hence unreactivity as well. This principle was named as Duplet or Octet Rule.

The discovery of duplet or octet rule was followed by another similar suggestion that atoms form bonds because they would like to lower their energy by completing their duplet or octet. For example, for sodium atom it is easy to lose one electron and stabilize itself than to gain seven electrons while completing its octet. Sodium atom, therefore, adopts the energetically easier path and loses its electron to form a bond. In the same way, it is energetically favourable for hydrogen atom to lose one electron to become proton (H^+) or gain one electron to become hydride ion (H^-). In the latter case, it completes its duplet.

Alkali and alkaline earth metals are therefore expected to be electropositive metals which will form bonds with electronegative elements like oxygen and chlorine. Although, in the beginning, octet rule played a significant role in understanding the nature of a chemical bond, yet further investigations found it to be less important.

3.2 Chemical Bond

A chemical bond is a force of attraction between atoms which holds them together in the form of a molecule or a compound.

When atoms of different substances approach each other, there are two possibilities. They may attract or repel each other. If the forces of attraction between them dominate the forces of repulsion, the energy of the system gets

lowered and as a result the two atoms will react to form a new molecule. Conversely, the two atoms simply move away from each other.



Important Information!

The arrangement of electrons around the nucleus of an atom in shells and sub-shells is called electronic configuration.

Types of Bonds

We shall consider here three types of bonds.

- (1) Ionic bond
- (2) Covalent bond
- (3) Coordinate covalent bond

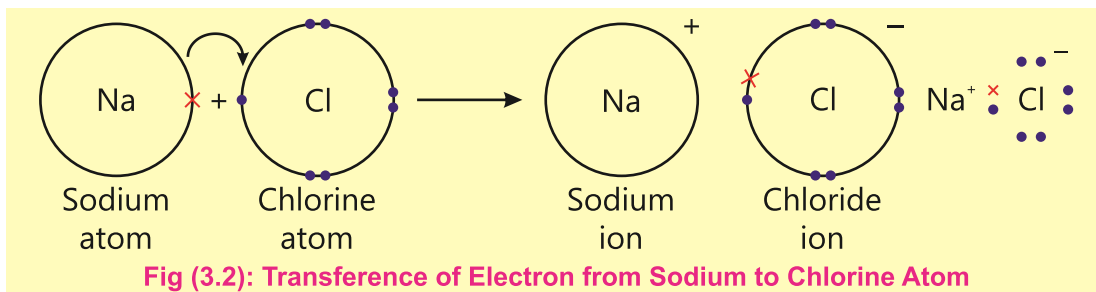
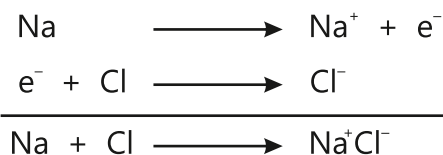
3.2.1 Ionic Bond

This chemical bond is formed as a result of the tendency of atoms to lose or gain electron or electrons to acquire the electronic configuration of the nearest noble gas because this is a more stable electronic structure. Let us take the example of the formation of a simple and important compound, sodium chloride. This compound is formed when the elements sodium and chlorine react chemically. The electronic configurations of these elements are shown in Fig (3.1).

	1st shell	2nd shell	3rd shell
$_{11}\text{Na}$	2	8	1
$_{17}\text{Cl}$	2	8	7

Fig (3.1): Electronic Configurations of Sodium and Chlorine

An electron from the outermost shell of sodium atom is transferred to the outermost shell of chlorine atom and in doing so, both these atoms acquire the electronic configurations of their nearest noble gases. (Fig 3.2)

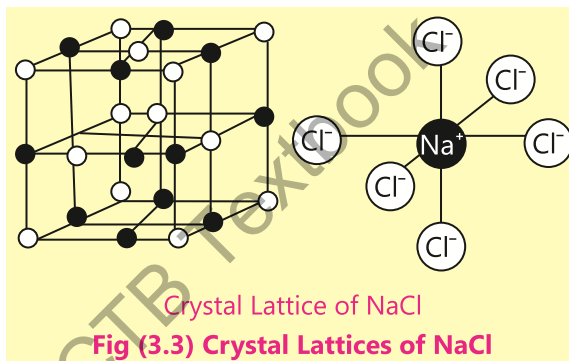


Similarly, sodium also reacts with fluorine and bromine to give sodium fluoride and sodium bromide respectively.

It should be noted here that an electron or electrons, which take part in a chemical reaction, come only from the outermost shells of the atoms. Sodium chloride, formed as a result of the chemical reaction mentioned on the previous page contains the positively charged sodium ions (Na^+) and the negatively charged chloride ions (Cl^-). These oppositely charged ions are then held together by the electrostatic force of attraction. The chemical bond, thus formed, is called an Ionic or an Electrovalent Bond and the compounds having such a bond are called ionic compounds.

These ions then surround each other three dimensionally to form a crystal lattice.

Examples of ionic compounds are KCl, Mg F, NaF, KBr, CaF_2 .
Figures of crystal lattices of NaCl. (Fig 3.3)



Exercise

1. What types of elements form ionic bonds?
2. What are the conditions for an ionic bond to be formed?

An ionic bond is therefore a bond which is formed by the complete transference of electron or electrons from one atom to another atom.

3.2.2 Covalent Bond

During the formation of an ionic bond, the atoms lower their energy by the transference of an electron (s) and thus acquire the electronic configurations of the nearest noble gas. However, it is not the only way by which atoms can lower their energy. Some atoms decrease their energy by mutually sharing their electrons. This can be explained as follows:

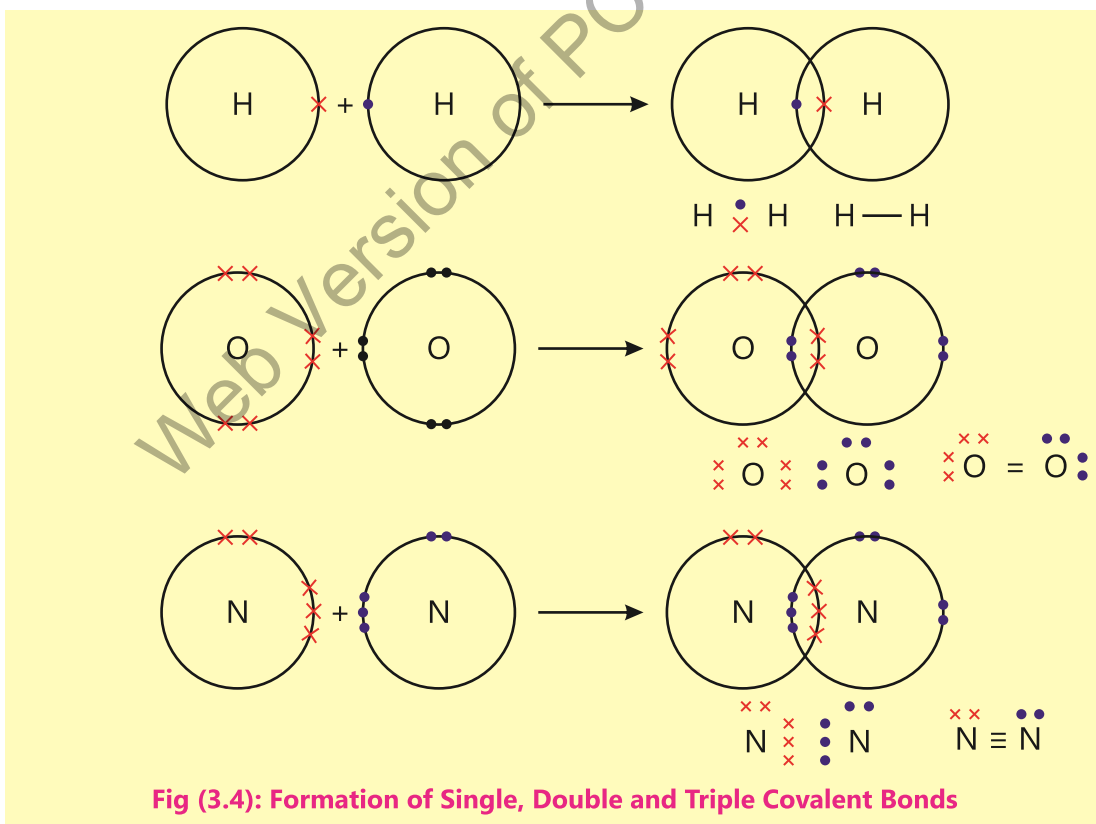
When two atoms approach each other in order to form a bond, they undergo important changes in their energy. The electrons belonging to one atom will come under the attractive influence of the nucleus of the other atom. This is the new force of attraction and will be responsible for lowering the energy. The electrons and the nucleus of one atom will also repel the electrons and the nucleus of the other atom. This is the force of repulsion and will obviously increase the energy. The two atoms will bring themselves at such a distance so

that the attractive forces dominate the repulsive forces. The total energy at this distance will be minimum and thus a stable molecule is formed. **A covalent bond is therefore a bond formed by the mutual sharing of an electron pair provided by the bonded atoms. This is called a single covalent bond.**

In some compounds, the atoms share two electrons each to form a double covalent bond. In the same way atoms can share three electrons each to form a triple covalent bond. Double and triple covalent bonds have two and three electron pairs respectively which are mutually shared between the two atoms. A single covalent bond is represented by a single line(–), a double covalent bond is represented by two lines (=) while a triple covalent bond is represented by three lines(≡). The mutually shared electrons may be shown by a dot or a cross. The formation of single, double and triple covalent bonds in different molecules is explained in the examples shown in Fig (3.5).

Exercise

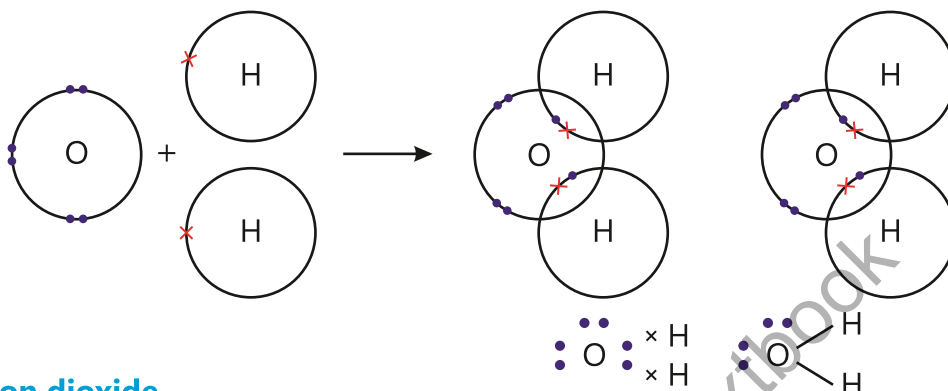
1. What type of elements form covalent bond?
2. How covalent bond is different from an ionic bond?



Formation of Covalent Compounds

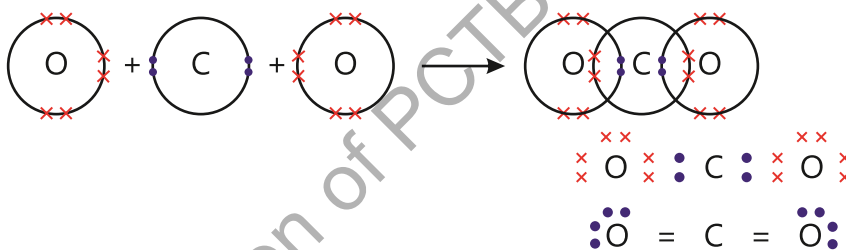
Water

A water molecule is formed when two hydrogen atoms share their electrons separately with the electrons of one oxygen atom.

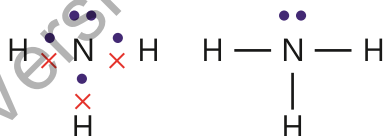


Carbon dioxide

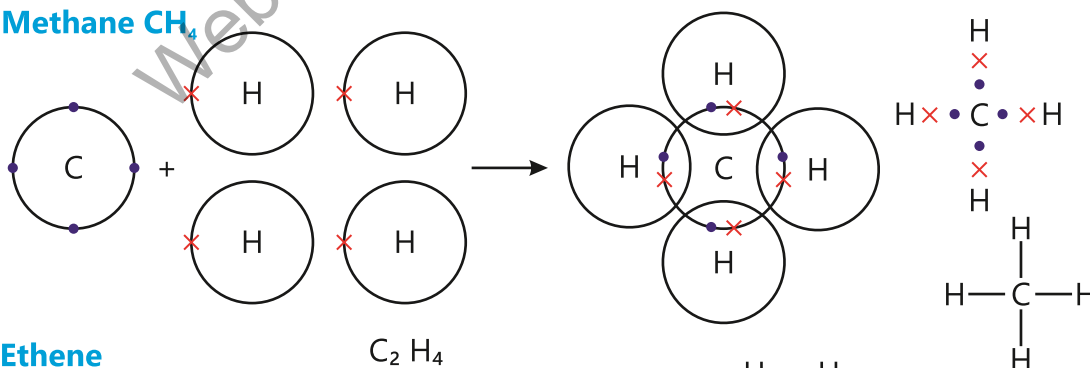
A carbon dioxide molecule is formed when an atom of carbon shares its four electrons with two oxygen atoms. Each oxygen atom also shares two electrons.



Ammonia



Methane CH₄



Ethene

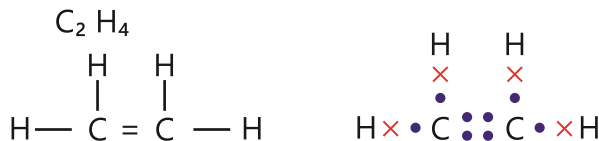
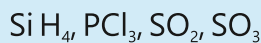


Fig (3.5): Formation of Covalent Compounds

Exercise

Draw electron dot and cross structure of the following compounds.



It is quite clear from the examples shown above that after mutually sharing their electrons, the bonded atoms acquire the electronic configurations of the nearest noble gas.

3.2.3 Coordinate Covalent Bond

Coordinate covalent bond is a type of covalent bond in which the shared electron pair is donated by one atom only. This bond is formed when a species has an electron pair to donate to another species. The species which donates the electron pair, is called a donor while that which accepts it is called an acceptor. An arrow head ($\rightarrow$) pointing towards the acceptor represents this type of bond. Following examples will help to explain this bond.

Hydroxonium Ion (H_3O^+)

Acids provide protons (H^+) when dissolved in water. This proton has an empty outer shell and can accept one of the two pairs of electrons present on the oxygen atom in water molecule. As a result of this, a hydroxonium ion (H_3O^+) is formed. (Fig 3.7)

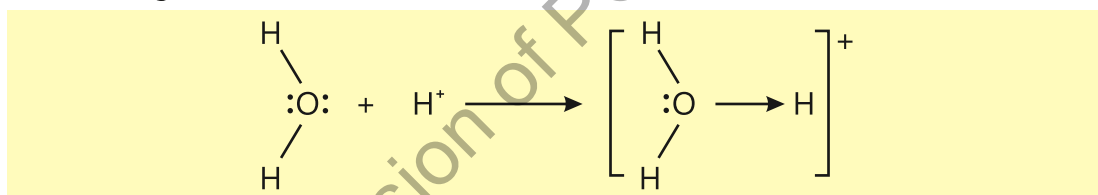
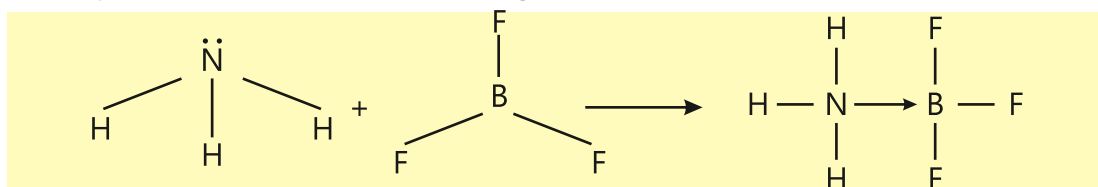


Fig (3.6): Formation of a Coordinate Covalent Bond Between H_2O and H^+

The positive charge covers whole of the hydroxonium ion. After the formation of hydroxonium ion, there does not remain any difference between a coordinate covalent bond and a covalent bond. All the three bonds of oxygen behave exactly alike.

Reaction Between ammonia and boron trifluoride

A reaction between ammonia (NH_3) and boron trifluoride (BF_3) is another example of the formation of a coordinate covalent bond. During the reaction, an electron pair from nitrogen of ammonia fills the partially empty outer shell of boron present in boron trifluoride Fig (3.8).



Other Examples

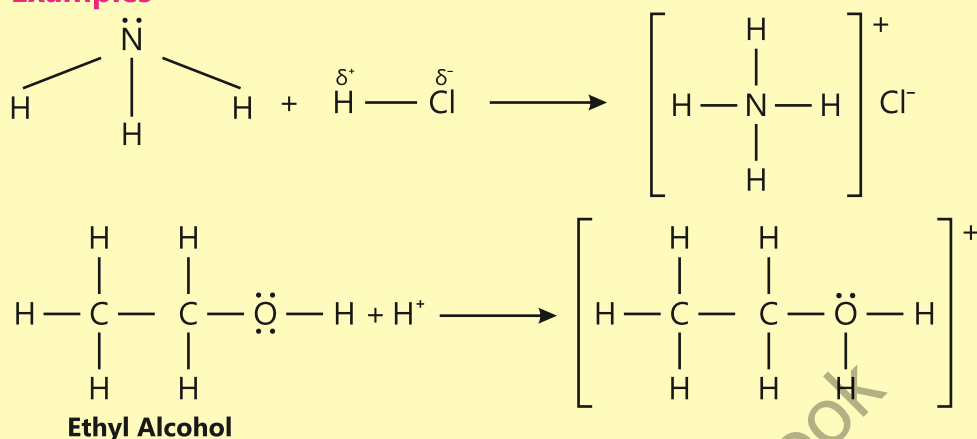


Fig (3.7): Formation of Boron trifluoride ammonia, ammonium chloride and protonated ethyl alcohol

In the above example, a coordinate covalent bond in ammonium chloride links nitrogen of ammonia and the proton. The positive charge is spread all over ammonium ion. All the four bonds between nitrogen and hydrogen in ammonium ion behave exactly alike. This proves the point that the difference between a covalent bond and a coordinate covalent bond lies in the way they are formed. Once such bonds are formed, there does not remain any difference. When a coordinate covalent bond breaks both the bonded electrons go to the donor species.

Exercise

- Draw the pictures of coordinate covalent bond formed between:
 - BF_3 and AlCl_3
 - CH_3OCH_3 and H^+
- Which compound is not able to form a coordinate covalent bond?

The characteristics shown by metals are very different from those of ionic and covalent compounds. This suggests the presence of different types of binding forces among the metallic atoms.

Properties of Metals

- Metals usually show metallic lustre.
- Metals usually have high melting and boiling points.
- Metals are good conductors of heat and electricity.

4. Metals are usually hard and heavy.
5. Metals can be made into different shapes by applying pressure.

These characteristics of metals can be explained if we know the nature of binding forces present between their atoms.

Usually metals have low values of ionization energy. Their atoms can therefore, lose their outer electron or electrons easily. In other words, the nuclei of metallic atoms cannot hold their outer electrons firmly. For example, in sodium metal, each sodium atom is surrounded by eight other sodium atoms. The outer electrons of these atoms move freely between the vacant spaces present between atoms because of the loose linkage they have with their nuclei. No electron remains attached with any particular nucleus. Instead, all the electrons, at the same time, get attached with all the nuclei. When all the atoms attract all the electrons collectively, obviously they will be bound together. A metal will appear to have a sea of electrons in which all the nuclei of atoms are submerged. A metallic bond, is therefore a type of chemical bond which has positively charged ions bound together by the mobile electrons. Fig (3.9)

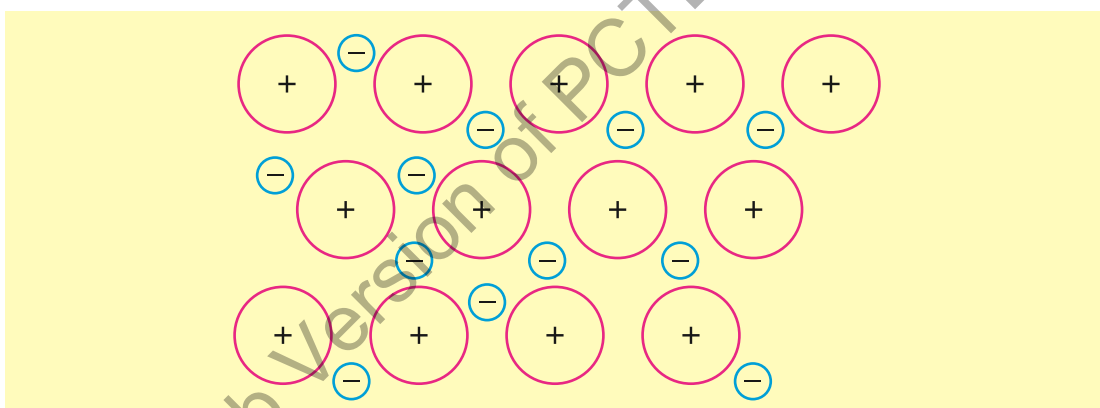


Fig (3.8): Metallic Bond in Sodium Metal

The strength of a metallic bond depends upon two factors: the number of positive charges present on the positive ions and the number of mobile electrons set free by each atom. In sodium metal, for example, each sodium atom sets free only one electron. The metallic bond in sodium metal is, therefore, not very strong. In magnesium metal, each magnesium atom releases two electrons to acquire two positive charges. The metallic bond in magnesium metal will evidently be stronger than that in sodium metal. This explains why the magnesium metal melts at a higher temperature than sodium metal. Fig (3.10)

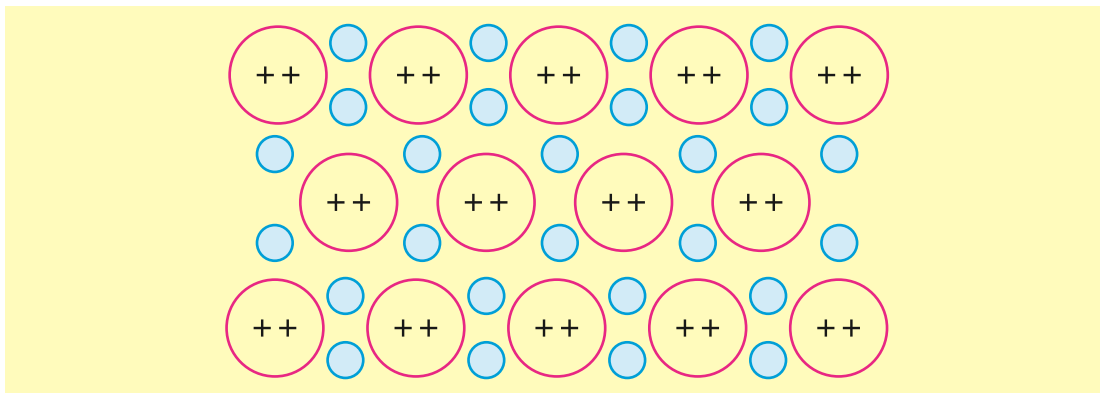


Fig (3.9): Metallic Bond in Magnesium Metal

The presence of freely moving electrons in metals makes them good conductor of heat and electricity. Moreover, in metals, the atoms are strongly held and arranged in the form of rows one above the other. This arrangement makes them hard and heavy. When pressure is applied on the metals, the upper rows of atoms slip past the lower rows. As a result, their shapes are changed. Metals can, therefore, be easily drawn into wires and sheets. Fig (3.11)

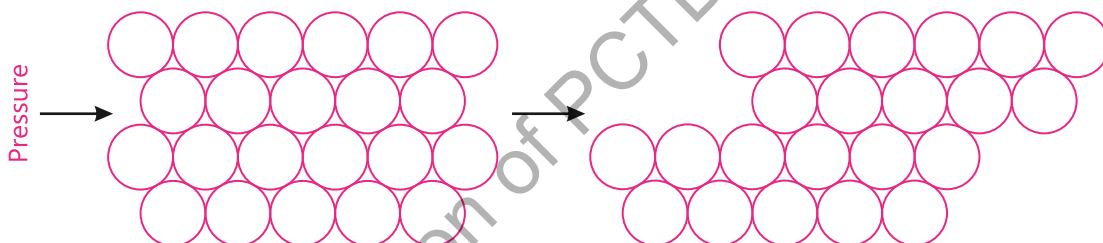


Fig (3.10) Metal under Pressure

Exercise

What type of atoms form metallic bond?

Give a comparison of metallic bond with an ionic bond.



Interesting Information!

Metals are extensively used in many industries. They are used in machinery, automobiles, railways, air crafts, rockets, in construction industry, in electronics industry, in jewellery, in electric wires and many more.

3.4 Electropositive Character of Metals

Metals generally have a tendency to lose electrons to form positive ions called cations. This property is called the electropositive character of metals. This property is also related to the reactivity of the metals. Metals which lose electron or electrons easily are considered more reactive. For example, alkali metals (Na, K) are highly electropositive elements and thus they undergo reactions very

easily. Sodium and potassium react vigorously with water and halogens to give their respective hydroxides and halides. They also react with acids to give salts and hydrogen.

Alkaline earth metals (Mg, Ca), on the other hand, lose their outer electrons less easily and thus they are less electropositive than alkali metals. Their reactions towards water and halogens are also less vigorous.

Aluminum is also highly electropositive metal. It reacts readily with mineral acids to form salts and hydrogen.

3.5 Electronegative character of Non-metals

Non-metals have an affinity towards electrons. They tend to gain electrons and become negatively charged ions called anions. They are therefore, named as electronegative elements. Fluorine is the most electronegative element in the periodic table followed by oxygen, nitrogen and chlorine. Non-metals readily react with metals forming ionic bonds. Non-metals also combine with other non-metals to form a wide variety of molecular substances.

3.6 Compare the properties of ionic and covalent compounds.

Ionic Compounds.

1. In ionic compounds oppositely charged ions are properly arranged to give a crystalline structure. As a whole the compound is neutral. There exists a strong electrostatic force between their ions.
2. Ionic compounds are usually solids having high melting and boiling points. The melting point of sodium chloride is 801°C because it is difficult to break the strong electrostatic forces of attraction between the oppositely charged ions.
3. Ionic compounds are generally soluble in polar solvents like water.
4. They are usually good conductor of electricity in molten state or in solution form. Their conductance is due to the presence of free ions in these forces.

Covalent Compounds.

1. Covalent compounds mostly exist as discrete neutral molecules. There exists a strong electrostatic attraction between the nuclei and the shared electrons.
2. Covalent compounds are made of two or more non-metals. Lower molecular mass covalent compounds are gases or low boiling liquids. High molecular mass covalent compounds exist as solids. Generally, they have lower melting and boiling points.
3. They are usually insoluble in water but soluble in non-polar solvents like ether, benzene and acetone.
4. They are usually bad conductor of electricity.

3.7 Nature of Bonding and Properties

In ionic compounds, the oppositely charged ions are held together by the strong electrostatic force of attraction in the form of a crystal lattice. Since the ions are rigid in ionic compounds, such compounds therefore exist in the form of very stable solids with significantly high melting points. Since ions are spherical and oppositely charged they can surround each other from all the sides, ionic bonds are non-directional. This arrangement of ions is called crystal lattice.

If an external force is applied on the crystal lattice, it breaks easily. It shows that ionic solids are highly brittle. In the solid form, ionic compounds do not conduct electricity because ions are tightly held and cannot move. However, in the molten state, the ions get free and start conducting electricity. Ionic solids are also generally soluble in water. Water not only breaks the electrostatic force of attraction but also hydrates the resulting free ions Fig (3.14). In the process of hydration water molecules surround and interact with ions or molecules.

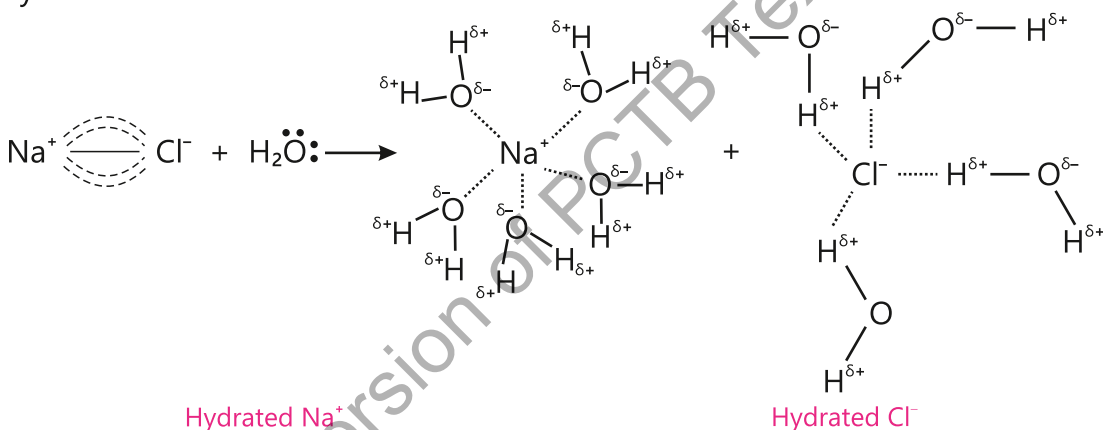
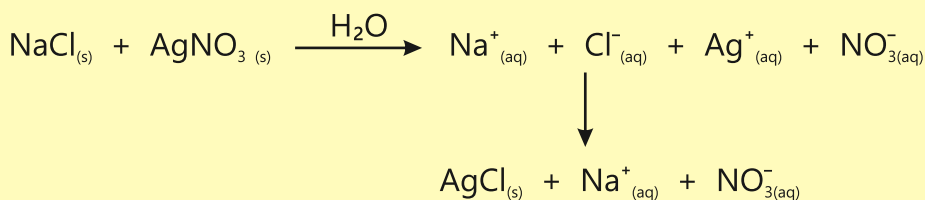


Fig (3.11): Hydration of Ions

Ionic compounds in an aqueous solution also conduct electricity because the free ions can now move towards their respective electrodes. Ionic compounds generally react in an aqueous solution. When we mix two solutions of ionic compounds, the positive ions of one compound may react with the negative ions of the other to form a new compound.



White precipitate of silver chloride comes out of the aqueous solution.

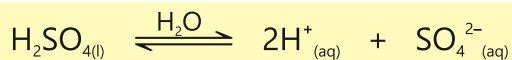
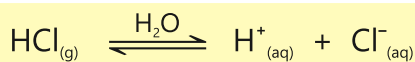
Elements and compounds in which atoms are covalent bonded, behave very differently from ionic compounds. Elements present at the right side of the periodic table exist as covalently bonded diatomic molecules except inert gases, for example, nitrogen(N_2), oxygen(O_2), fluorine(F_2) and chlorine(Cl_2). Due to very weak forces of attraction between their molecules, their densities and boiling points are very low. Bromine (Br_2) exists as volatile fuming liquid while elements like carbon, phosphorous and sulphur exist as covalent solids. All these solid elements exist both in amorphous and crystalline forms.

Coal is the amorphous form of carbon whereas diamond and graphite are its crystalline forms. Coal is used as a fuel in electricity generating plants. In diamond, each carbon atom is surrounded by four other carbon atoms linked together by strong covalent bonds. Due to this rigid structure, diamond is the hardest thing on this planet. It is used as a cutting, polishing and drilling tool.

Graphite consists of a layered structure, made of hexagonal rings of carbon. Since layers are not bonded strongly, they can slip past each other. Graphite is thus used as a lubricant in industry. Further, these layers in graphite have mobile electrons in between them. Graphite is a good conductor of electricity and it is also used as an electrode.

Binary covalent compounds generally exist as low temperature boiling gases except water. Methane (CH_4), ammonia (NH_3), hydrogen sulphide (H_2S), hydrogen chloride (HCl), nitrogen dioxide (NO_2), carbon dioxide (CO_2) and sulphur dioxide are all covalent compounds which are gases at room temperature.

Water and hydrogen fluoride, on the other hand, are liquids at room temperature. Liquid water has a high boiling point because strong intermolecular forces are present between its molecules. Covalent molecules like hydrogen chloride, sulphuric acid and nitric acid ionize completely in water behaving as very strong acids.



Exercise



1. Tick (✓) the correct answer.

- i. When molten copper and molten zinc are mixed together, they give rise to a new substance called brass. Predict what type of bond is formed between copper and zinc.
- (a) Coordinate covalent bond (b) Ionic bond
(c) Metallic bond (d) Covalent bond
- (ii) Which element is capable of forming all the three types of bonds; covalent, coordinate covalent and ionic?
- (a) Carbon (b) Oxygen
(c) Magnesium (d) Silicon
- (iii) Which of the following bonds is expected to be the weakest?
- (a) C – C (b) Cl – Cl
(c) O – O (d) F – F
- (iv) Which form of carbon is used as a lubricant?
- (a) Coal (b) Diamond
(c) Graphite (d) Charcoal
- (v) Which metal has the lowest melting point?
- (a) Li (b) Na
(c) K (d) Rb
- (vi) Which ionic compound has the highest melting point?
- (a) NaCl (b) KCl
(c) LiCl (d) RbCl
- (vii) Which compound contains both covalent and ionic bonds?
- (a) MgCl_2 (b) NH_4Cl
(c) CaO (d) PCl_3
- (viii) Which among the following has a double covalent bond?
- (a) Ethane (b) Methane
(c) Ethylene (d) Acetylene

2. Questions for Short Answers

- i. What type of elements lose their outer electron easily and what type of elements gain electron easily?
- ii. Why do lower molecular mass covalent compounds exist as gases or low boiling liquids.

- iii. Give one example of an element which exists as a crystalline solid and it has covalent bonds between its atoms.
- iv. Which property of metals makes them malleable and ductile?
- v. Is coordinate covalent bond a strong bond?
- vi. Write down dot and cross formula of HNO_3 .

3. Constructed Response Questions

- i. Why HF is a liquid while HCl is a gas?
- ii. Why covalent compounds are generally not soluble in water?
- iii. Why does iodine exist as a solid while Cl_2 exist as a gas?

4. Descriptive Questions

- i. Explain the formation of an ionic bond and a covalent bond.
- ii. How do ions arrange themselves to form NaCl crystal.
- iii. Explain the properties of metals keeping in view the nature of metallic bond.
- iv. Compare the properties of ionic and covalent compounds.
- v. How will you explain the electrical conductivity of graphite crystals?
- vi. Why are metals usually hard and heavy?

Stoichiometry

Student Learning Outcomes

After studying this chapter, students will be able to:

- State the formulae of common elements and compounds.
- Define molecular formula of a compound as the number and type of different atoms in one molecule
- Define empirical formula of a compound as the simplest whole number ratio of different atoms in a Molecule.
- Deduce the formula and name of a binary ionic compounds from ions given relevant information
- Use the relationship amount of substance = mass / molar mass to calculate number of moles, mass, molar mass, relative mass (atomic/molecular/formula) and number of particles
- Define mole as amount of substance containing Avogadro's number 6.02×10^{23} of particles
- Explain the relationship between a mole and Avogadro's constant.
- Construct chemical equations and ionic equations to show reactants forming products, including state symbols.
- Deduce the symbol equation with state symbols for a chemical reaction given relevant information.

Stoichiometry is an important concept in chemistry which helps us to calculate the amounts of reactants and products by using a balanced chemical equation. It is based on the law of conservation of mass which states that matter can neither be created nor destroyed. Therefore, the total mass of all the reactants must be equal to the total mass of all the products. The stoichiometric coefficient used to balance a chemical equation provides the mole ratio between reactants and products.

Stoichiometry is used in industry quite often to determine the amount of raw materials required to produce the desired amount of the products.

4.1 Chemical formula

Elements exist in different forms in this world. There are elements which exist in the form of aggregate of atoms. These elements are represented by their symbols alone, for example, Na, Ca, C, Fe, etc. On the other hand, elements like O_2 , N_2 , H_2 exist as discrete molecules in which their atoms are chemically bonded to each other. In ozone, three atoms of oxygen are bonded to each other, so its chemical formula is O_3 .

Similar to elements, chemical compounds also exist in different forms. Common salt i.e. sodium chloride exists in the form of ions which are bonded together in the form of a crystal. Since ratio between its ions is 1:1, sodium chloride is represented by a formula unit NaCl. Similarly, the other ionic compounds are represented by their formula units which show the minimum ratio present between their ions. Examples are $CaCl_2$, KBr and $BaCl_2$, etc.

Covalent compounds generally exist in discrete molecules in which atoms are bonded together. For example, water exists as molecules which are represented by the chemical formula H_2O . It shows that in one molecule of water two atoms of hydrogen are bonded to one atom of oxygen. Similarly, chemical compound, ammonia is represented by NH_3 and methane gas is represented by CH_4 . A chemical compound is thus, represented by a chemical formula which is called the molecular formula of that compound and which shows all the types of atoms bonded together in one molecule of that compound. Examples of covalent compounds are HCl, HF, H_2S , PH_3 , H_2O_2 , H_2SO_4 , CO_2 , CO, C_6H_6 , etc.

Exercise

How would you differentiate between the chemical formula of an element and that of a compound? Give examples. Write down the names of ionic and covalent compounds whose formulas have been given in this article.

4.2 Empirical Formula

Empirical Formula of a compound shows the simplest whole number ratio of atoms present in that compound. All the ionic compounds are represented by their empirical formulas. These formulas show the simplest ratio present between their ions. The empirical formula of calcium fluoride is CaF_2 which shows the ratio present between calcium and fluoride ions in its crystal.

For covalent chemical compounds, which exist as molecules, the empirical formulas may be different from their molecular formulas. For example, hydrogen

peroxide is represented by its molecular formula H_2O_2 ; its empirical formula will be HO. Similarly, a benzene molecule has C_6H_6 as its molecular formula; so its empirical formula will be CH. For water, the molecular formula and the empirical formula are both the same i.e. H_2O because there does not exist any minimum ratio between hydrogen and oxygen atoms.

Since an empirical formula does not tell us the actual number of atoms present in that compound, rather it represents the simplest ratio between atoms, it is possible that some compounds may have the same empirical formula. For example, both benzene (C_6H_6) and acetylene (C_2H_2) have the same empirical formula CH.

Exercise

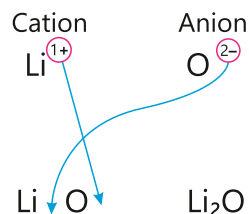
Give two examples of the compounds which have same empirical and molecular formulas.

4.3 Chemical Formulas of Binary Ionic Compounds

In order to write down the formula of an ionic compound, first identify the cations and anions and the number of charges present on them. Finally combine the two ions together to form an electrically neutral compound.

If you know the name of binary ionic compound, you can write its chemical formula. Start by writing symbol of cation with its charge. Then write the symbol of anion with its charge and find out how many of these ions are needed to give an electrically neutral compound. For example, write down the formula of lithium oxide. The symbol of lithium cation with its single positive charge is Li^+ . The symbol of anion is O^{2-} .

Let us now apply crisscross method to write the formula. In this method, the numerical value of each of the ion charges is crossed over to become the subscript of the other ion. Signs of the charges are then dropped.



Write down the formula of Aluminium oxide.



Write down the formula of Magnesium nitride.



Table (4.1) shows some important atoms, their ions and the charges on the ions

**Interesting Information!**

The composition of all the chemical products we use in our lives, such as shampoos, perfumes, soaps and fertilizers are formed using stoichiometric calculations. Without stoichiometry the chemical industry does not exist.

Table (4.1) Atoms and their Cations and Anions

Atoms and their Cations with charges		Atoms and their Anions and Cations with charges	
Atom	Charge	Atom	Charge
H	H ¹⁺	O	O ²⁻
Na	Na ¹⁺	N	N ³⁻
Li	Li ¹⁺	Cl	Cl ¹⁻
K	K ¹⁺	Br	Br ¹⁻
Mg	Mg ²⁺	I	I ¹⁻
Ca	Ca ²⁺	Cu	Cu ¹⁺ , Cu ²⁺
Ba	Ba ²⁺	Fe	Fe ²⁺ , Fe ³⁺
Zn	Zn ²⁺	Sn	Sn ²⁺ , Sn ⁴⁺
Al	Al ³⁺	Ni	Ni ²⁺

4.4 Chemical Formulas of Compounds

Molecular formula of a compound can be found out if we know its empirical formula. To calculate the empirical formula of a compound, you need to determine the simplest whole-number ratio of atoms in the compound. This can be done by using experimental data on the mass percent composition of the compound. Molecular formula is then calculated by the following relationship.

$$\text{Molecular formula} = n (\text{Empirical Formula})$$

$$\text{Where } n = \frac{\text{Molar Mass}}{\text{Empirical Formula mass}}$$

For example, the empirical formula of hydrogen peroxide is HO. Its molar mass is 34. Its molecular formula will then be

$$n = \frac{34}{17} = 2$$

$$\text{Molecular formula} = (\text{HO})_2 = \text{H}_2\text{O}_2$$

If for a compound the value of n is one, then its molecular formula is the same as its empirical formula.

Exercise

Write down the names of three such compounds which have different empirical and molecular formulas.

Sample Problem:

Empirical formula of a compound is CH. Its molecular mass is 78 g mol^{-1} . Find out its molecular formula.

Solution:

Empirical Formula	= CH	Molecular mass = 78 g mol^{-1}
Molecular formula	= $n(\text{Empirical formula})$	
n	= molar mass / empirical formula mass	
	= $\frac{78}{13} = 6$	
Molecular Formula	= $(\text{CH})_6 = \text{C}_6\text{H}_6$	

Exercise

1. The empirical formula of a compound is CH_2O . Its molar mass is 180 g mol^{-1} . Determine its molecular formula.
2. The empirical formula of a compound is CH_2O . Its molar mass is 60 g mol^{-1} . Determine its molecular formula.

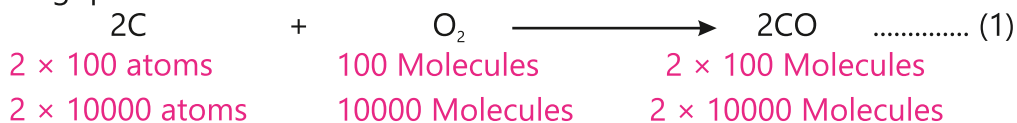
4.5 Avogadro's Number (N_A)

In a chemical reaction, large number of atoms or molecules of reactants react to give the products. We would very much like to know the mass ratio in which the reactants react. For this purpose, we would also like to express these masses of reactants in grams. To achieve this objective, we need to transform the concepts of chemical formula and atomic mass units into such concepts which may lead us to know the masses of reacting elements and compounds in grams. Avogadro, an Italian scientist, helped us to achieve this objective in the following way.

Let us consider the following reaction in which two atoms of carbon react with a molecule of oxygen to produce two molecules of CO.



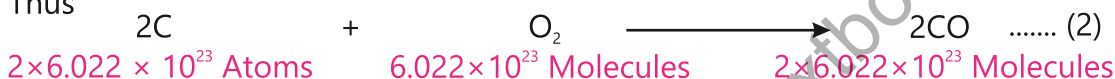
Since it is not possible to account for the masses of individual atoms or molecules because these are very small particles, we increase the number of reacting species as written below.



Increasing the number of reacting atoms or molecules will not change the ratio in which these are reacting or are being formed.

Increasing the number of reacting species, however, has not solved the problem because this number is still very small. We should increase this number to such a value whereby it is convenient for us to calculate their masses.

Thus



6.022×10^{23} is a huge number and we have selected this because

$$1.00 \text{ g} = 6.022 \times 10^{23} \text{ amu}$$

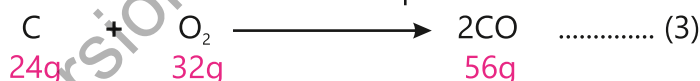
Now the amounts of reactants and products in the forementioned equations can be written as follows.

$$2 \times 6.022 \times 10^{23} \times (12.0 \text{ amu}) = 24.00 \text{ g carbon atoms}$$

$$6.022 \times 10^{23} \times (32.0 \text{ amu}) = 32.00 \text{ g oxygen molecules}$$

$$2 \times 6.022 \times 10^{23} \times (28.0 \text{ amu}) = 2 \times 28.00 \text{ g carbon monoxide molecules}$$

The mass ratio between the reactants and those of products will then become



You must have realized that starting from a simple equation we have developed such ratio of masses of the reacting species which can conveniently be used in the laboratory.

According to the above-mentioned equation, 24 g of carbon contains $2 \times 6.022 \times 10^{23}$ atoms of carbon, 32 g of oxygen contains 6.022×10^{23} molecules of oxygen and 56 g of carbon monoxide contains $2 \times 6.022 \times 10^{23}$ of its molecules.

The number 6.022×10^{23} is called Avogadro's number after the name of an Italian chemist Amedeo Avogadro. Avogadro's number is the number of units in one mole of a substance. This number is represented as N_A .

4.6 The Mole and Molar Mass

Avogadro's number has an immense significance in Chemistry and the quantity of a substance containing Avogadro's number of particles (N_A) is called **a Mole**. Mole is a number like a dozen or a gross. A dozen of oranges means 12

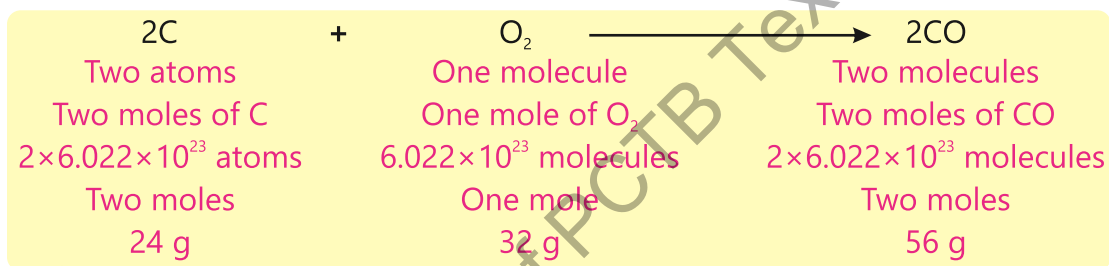
oranges. Similarly, a mole of a substance means its 6.022×10^{23} particles which can be atoms, molecules or ions. When we use the term mole of a substance, we must also refer to what type of particles are present in this substance. The following examples will help you to understand the concept clearly.

A mole of carbon atoms contains 6.022×10^{23} atoms and weighs 12 g.

A mole of oxygen molecules (O_2) contains 6.022×10^{23} molecules and weighs 32g. A mole of sodium chloride ($NaCl$) consists of 6.022×10^{23} of its formula units and its mass is 58.5 g.

The mass of one mole of a substance is called Molar mass. The molar mass of hydrogen atoms refers to the mass of one mole of hydrogen atoms and its value is 1.008 g. Similarly the molar mass of hydrogen molecules will be 2.016 g.

The chemical equation discussed in the previous article will now be understood in the following way.

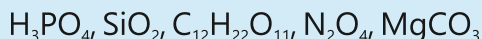


Interesting Information!

Mole is important because atoms and molecules are so small that they can not be counted. The mole concept allows us to count atoms and molecules by weighing macroscopically small amounts of matter.

Exercise

Calculate the molar masses of the following compounds



4.7 Chemical Equations and Chemical Reactions

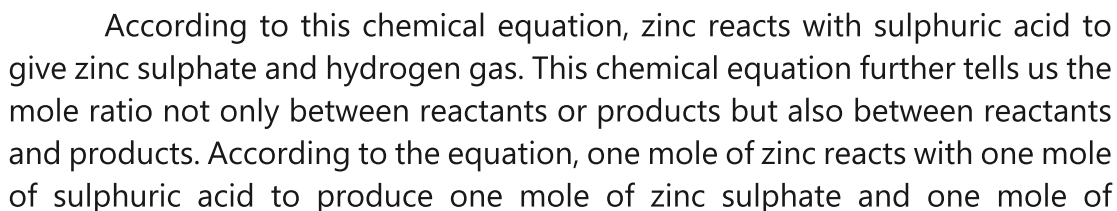
To understand a chemical change and to study the different factors which control it, has always been a focal point in the efforts of chemists. To achieve these objectives, we should have an appropriate way of representing a chemical change. Fortunately, the chemists have developed a very suitable way of representing a chemical change in terms of symbols of elements and formulas of compounds. **Representing a chemical change in this way is called a chemical equation.** A chemical equation tells us the elements or compounds which are reacting and those which are being produced as a result of a chemical change.

$$4\text{Al}_{(s)} + 3\text{O}_{2(g)} \longrightarrow 2\text{Al}_2\text{O}_{3(s)}$$

Reactants Products

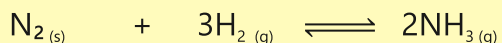


- A chemical equation can, however, only be written if all the previous mentioned points are experimentally verified. For example, the nature of the products and their correct formulas must first be ascertained before writing a chemical equation. After the experimental verification of all the information about a chemical change, we should now be able to write a correct chemical equation. The following equation has been written keeping in view all the points mentioned above.



hydrogen gas.

Sometimes a chemical reaction moves both ways. In other words, the reactants react to give the products and the products, in turn, react to give the reactants back. Such reactions are called as **reversible reactions** and are indicated by ($\rightleftharpoons$) sign, e.g.:



Reaction involving ions may also be shown in the form of chemical equation. Both AgNO_3 and NaCl are ionic compounds and are soluble in water. When they are mixed in water, they react to form products.

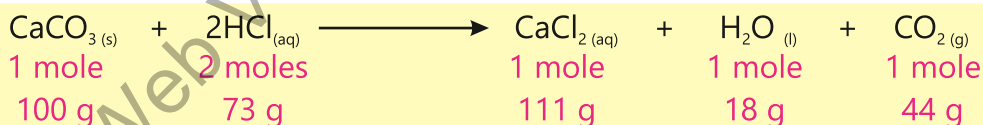


AgCl being insoluble comes out of the aqueous solution in the form of a precipitate.

4.8 Calculations Based on Chemical Equation

A complete and balanced chemical equation tells us the mole ratio or molar mass ratio between the reactants and the products. With the help of this ratio, we can find out the molar masses of the products provided we know the molar masses of the reactants. Similarly the molar masses of the reactants can also be found out if we know the molar masses of the products.

For example, the following equation tells us that one mole (100 g) of calcium carbonate reacts with two moles (73 g) of hydrochloric acid to produce one mole (111 g) of CaCl_2 , one mole (18 g) of water and one mole (44 g) of carbon dioxide.



The total masses of the reactants are equal to the total masses of the products.

Example 1

25g of limestone (CaCO_3) reacts with an excess of hydrochloric acid according to the above equation. How much calcium chloride (CaCl_2) will be produced?

Solution

Mass of $\text{CaCO}_3 = 25 \text{ g}$

Molar mass of $\text{CaCO}_3 = 100 \text{ g mol}^{-1}$

Mass of CaCl_2 product = ?

Molar mass of $\text{CaCl}_2 = 111 \text{ g mol}^{-1}$

Solution Mass of $\text{CaCO}_3 = 25 \text{ g}$ Molar mass of $\text{CaCO}_3 = 100 \text{ g mol}^{-1}$
 Mass of CaCl_2 product = ? Molar mass of $\text{CaCl}_2 = 111 \text{ g mol}^{-1}$

According to the equation

100 g of limestone reacts to produce calcium chloride = 111 g

1 g of limestone will react to produce calcium chloride = $\frac{111}{100} \text{ g}$

25 g of limestone will react to produce calcium chloride = $\frac{111}{100} \times 25$
 = 27.75g

Example 2

1.80 moles of ethyl alcohol, when burnt in air completely, will utilize how many moles of oxygen gas? Also calculate the number of moles of CO_2 produced.

Solution No. of moles of ethyl alcohol = 1.80

No. of moles of oxygen needed = ?

The balanced chemical equation for the reaction will be



According to this equation

One mole of ethyl alcohol, when completely burnt, needs oxygen = 3 moles

1.8 moles of ethyl alcohol, upon burning, will need oxygen = $\frac{3}{1} \times 1.8$
 = 5.4 moles

1 mole of ethyl alcohol produces moles of CO_2 = 2.0

1.8 mol of ethyl alcohol will produce = $\frac{2}{1} \times 1.8$
 = 3.6 moles

Example 3

Aluminium metal reacts with oxygen to produce aluminium oxide. How many grams of oxygen will be required to react completely with 0.3 moles of aluminium?

Solution Moles of Al = 0.3 Grams of O₂ used = ?

The balanced chemical equation for the reaction.



According to this equation

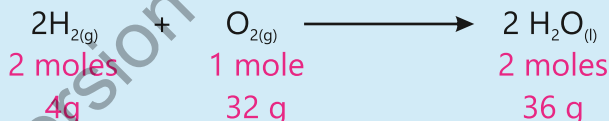
4 moles of aluminium need oxygen	= 3.0 moles
1.0 mole of aluminium will need oxygen	= $\frac{3}{4}$
0.3 moles of aluminium will need oxygen	= $\frac{3}{4} \times 0.3$
	= 0.225 moles
1 mole of oxygen (O ₂) has molar mass	= 32 g
0.225 mole of oxygen (O ₂) will have molar mass	= 32 × 0.225
	= 7.2 g

Example 4

How many molecules of water will be produced if we react 5 g of hydrogen gas with excess of oxygen gas?

Solution Mass of H₂ used = 5g Molecules of water produced = ?

The balanced chemical equation for the reaction:



According to this equation

4 g of hydrogen produce H ₂ O	= 36 g of H ₂ O
5 g of hydrogen will produce H ₂ O	= $\frac{36}{4} \times 5$ = 45g of H ₂ O
18g (1 mole) of H ₂ O contain molecules	= 6.022 × 10 ²³
36g of H ₂ O contain molecules	= 6.022 × 10 ²³ × 2
	= 12.04 × 10 ²³
45g of H ₂ O contain	= $\frac{45}{36} \times 12.04 \times 10^{23}$
	= 1.505 × 10 ²⁴ molecules

Exercise



1. Tick(✓) the correct answer.

- (i) How many atoms are present in one gram of H_2O ?
(a) 1002×10^{23} atoms (b) 6.022×10^{23} atoms
(c) 0.334×10^{23} atoms (d) 2.004×10^{23} atoms
- (ii) Which is the correct formula of calcium phosphide?
(a) CaP (b) CaP_2
(c) Ca_2P_3 (d) Ca_3P_2
- (iii) How many atomic mass units (amu) are there in one gram?
(a) 1 amu (b) 10^{23} amu
(c) 6.022×10^{23} amu (d) 6.022×10^{22} amu
- (iv) Structural formula of 2-hexene is $\text{CH}_3 - \text{CH} = \text{CH} - (\text{CH}_2)_2 - \text{CH}_3$. What will be its empirical formula?
(a) C_2H_2 (b) CH
(c) C_6H_{12} (d) CH_2
- (v) How many moles are there in 25 g of H_2SO_4 ?
(a) 0.765 moles (b) 0.51 moles
(c) 0.255 moles (d) 0.4 moles
- (vi) A necklace has 6g of diamonds in it. What are the number of carbon atoms in it?
(a) 6.02×10^{23} (b) 12.04×10^{23}
(c) 1.003×10^{23} (d) 3.01×10^{23}
- (vii) What is the mass of Al in 204g of aluminium oxide, Al_2O_3 .
(a) 26g (b) 27g (c) 54g (d) 108g
- (viii) Which one of the following compounds will have the highest percentage of the mass of nitrogen?
(a) $\text{CO}(\text{NH}_2)_2$ (b) N_2H_4 (c) NH_3 (d) NH_2OH
- (ix) When one mole of each of the following compounds is reacted with oxygen, which will produce the maximum amount of CO_2 ?
(a) Carbon (b) Diamond
(c) Ethane (C_2H_6) (d) Methane (CH_4)
- (x) What mass of 95% CaCO_3 will be required to neutralize 50 cm^3 of 0.5M HCl solution?
(a) 9.5g (b) 1.25g (c) 1.32g (d) 1.45g

2. Questions for Short Answers

- Write down the chemical formula of barium nitride.
- Find out the molecular formula of a compound whose empirical formula is CH_2O and its molar mass is 180.
- How many molecules are present in 1.5 g H_2O ?
- What is the difference between a mole and Avogadro's number?
- Write down the chemical equation of the following reaction.

Copper + Sulphuric acid $\rightarrow$ Coppersulphate + Sulphur dioxide + Water

3. Constructed Response Questions

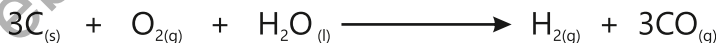
- Different compounds will never have the same molecular formula but they can have the same empirical formula. Explain
- Write down the chemical formulas of the following compounds.
Calcium phosphate, Aluminium nitride, Sodium acetate, Ammonium carbonate and Bismuth sulphate.

4. Descriptive Questions

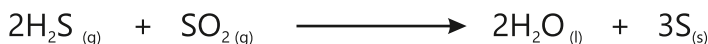
- Which conditions must be fulfilled before writing a chemical equation for a reaction?
- Explain the concepts of Avogadro's number and mole.
- How many grams of CO_2 will be produced when we react 10 g of CH_4 with excess of O_2 according to the following equation?



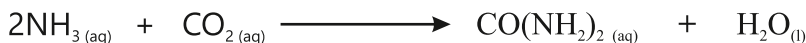
- How many moles of coal are needed to produce 10 moles of CO according to the following equation?



- How much SO_2 is needed in grams to produce 10 moles of sulphur?



- How much ammonia is needed in grams to produce 1 kg of urea fertilizer?



- Calculate the number of atoms in the following.

(a) 3g of H_2 (b) 3.4 moles of N_2 (c) 10g of $\text{C}_6\text{H}_{12}\text{O}_6$.



Energetics

Student Learning Outcomes

After studying this chapter, students will be able to:

- Explain the idea of a chemical system and its connection with its surroundings influences energy transfer during a chemical reaction.
- Differentiate between exothermic and endothermic reactions by giving examples.
- State that thermal energy is called enthalpy change and recognize its sign as negative for exothermic and positive for endothermic reactions
- Define activation energy as the minimum energy that colliding particles must have for a successful collision.
- Explain that activation energy depends on reaction pathway which can be changed using catalysts or enzyme (detailed pathways not required)
- Draw, label and interpret reaction pathway diagram for exothermic and endothermic reaction which includes enthalpy change, activation energy (uncatalyzed and catalyzed), reactants and products
- Recognize that bond breaking is endothermic and bond making is exothermic processes.
- Explain that enthalpy change is sum of energies absorbed and released in bond breaking and bond forming
- Explain that enthalpy change is sum of energies absorbed and released in bond breaking and bond formation.

Energy exists in different forms which are often interconvertable. In chemical energetics we are mainly concerned with two forms of energy.

1. **Chemical Energy:** This energy is stored in a molecule in which atoms are bonded to each other.
2. **Heat Energy:** This form of energy is released when a bond is formed and absorbed when it is broken.

In energetics we study the energy changes that take place during a chemical reaction. These changes are caused by the making and breaking of bonds during a reaction. In most of the reactions the weak bonds of reactants break while in products new strong bonds are formed. Since energy is needed to break a bond while energy is evolved when a bond is formed, such reactions take place always with the evolution of heat. If a reaction is accompanied with the evolution of heat it is called an **exothermic reaction** and if heat is absorbed during a reaction it is called an **endothermic reaction**.

In energetics we not only encounter heat which comes out of a chemical reaction but also another quantity which is called enthalpy.

Enthalpy (H) or heat content, is defined as the total amount of thermal energy stored in a compound. The unit of its measurement is kJ mol^{-1} .

When the energy is absorbed during a reaction, the total enthalpy of the system increases. When energy is evolved during a reaction, the total enthalpy of the system decreases.

Chemical energetics is part of a broader field of chemistry called Thermodynamics. In energetics we study the flow of energy in a chemical reaction. Thermodynamics deals with the energy changes during chemical reactions and how these changes affect the properties of a chemical system.

Thomas Young was the first to use the word 'energy' to the field of physics in 1802.

5.1 System and Surrounding

In Chemistry, any physical or chemical change under study may also be called a system. The chemical reaction includes reactants, products, catalyst, solvent and anything else which is important to study this reaction. Everything else which does not fall in this system is called the surrounding. For example, if you are boiling water in the beaker, the water molecules will be called a system while everything surrounding this like beaker, burner, etc. will be called the surrounding. When energy is transferred from surrounding to the system, the change is called endothermic and it has a positive sign. When energy is transferred from system to surrounding, the change is called exothermic and it carries a negative sign, Fig. (5.1).



Interesting Information!

Energy evolved during a chemical reaction is used in everyday life for cooking, heating, lighting, transportation and much more.

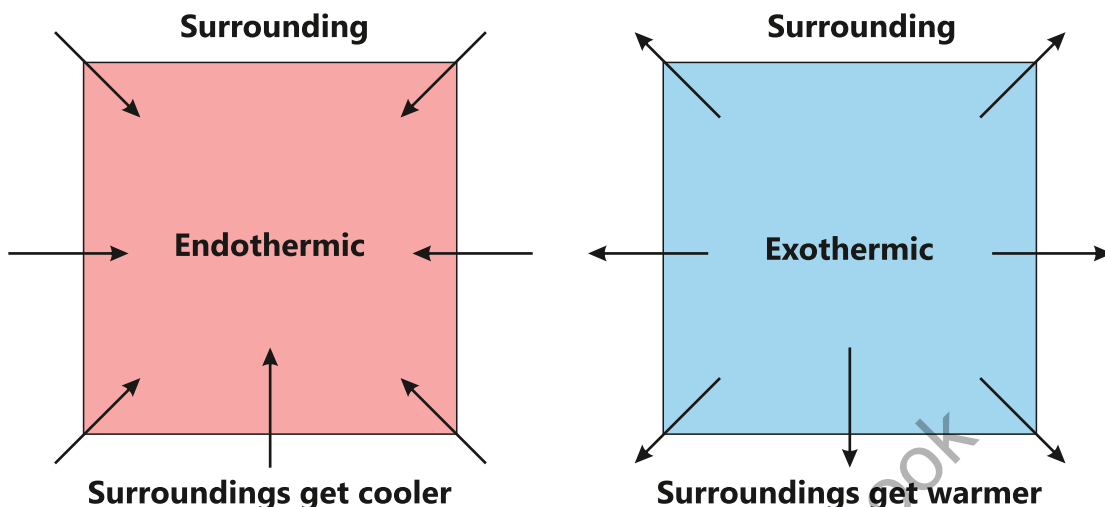


Fig (5.1) System and surrounding exchanging energy

Exercise

Does boiling water in a beaker endothermic or exothermic change? Which form of energy is being transferred in this system.



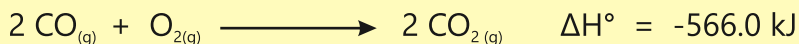
Interesting Information!

Enthalpy is important because it tells us how much heat is present in a system. Heat is important because we can extract useful work from it.

5.2 Enthalpy

The total amount of heat energy present in a molecule under standard conditions (0°C temperature and 760 mm pressure) is also called its heat content. Enthalpy is the measurement of energy in a thermodynamic system. The quantity of enthalpy is equal to the total heat content of a system. Enthalpy of a system is represented by (H) while the change in enthalpy which a system undergoes is represented by ΔH . The total enthalpy (H) of a system cannot be measured directly. However, the change in enthalpy ΔH brought about in a system can be measured comparatively easily.

In Chemistry, the standard enthalpy of reaction (ΔH°) is the enthalpy change when reactants in their standard states undergo reaction to produce products in their standard states. This quantity is called the standard enthalpy change or heat of reaction at constant pressure.



The reaction in this system is thus exothermic evolving 566 kJ of heat energy which is given to the surrounding.

How is enthalpy different from heat?

Heat is a form of energy that flows from hot body to a cold body because of a difference in temperature. We measure heat in joules. Heat is what we call the transfer of thermal energy. Contrary to this, enthalpy is an essential part of a system since it depends on the number of molecules present in that system, its chemical composition and its structure. Heat is not essential part of a system, it just comes and goes. When heat leaves or enters a system, it results in a change of enthalpy. At a constant pressure the enthalpy change is equal to heat evolved or absorbed.

Exercise

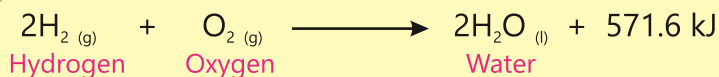
1. Can energy be transferred in a form other than heat during a chemical reaction?
2. Why it is not possible to calculate the enthalpy of a system?

5.3 Exothermic and Endothermic Reactions

A physical or a chemical change is almost always accompanied with either absorption or evolution of heat. Heat, which is evolved or absorbed during a chemical reaction, is called the heat of that reaction.

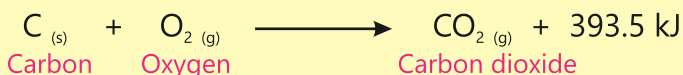
Chemical reactions in which heat energy is evolved are called exothermic reactions while those in which heat energy is absorbed are called endothermic reactions. Heat, which is evolved during an exothermic reaction, goes to the surrounding and the container in which such a reaction is being carried out, gets hot. Conversely, in an endothermic reaction, the absorption of heat from the surrounding will decrease the temperature of the container.

Example: Hydrogen gas and oxygen gas react to give liquid water in an exothermic reaction.



571.6 kJ heat energy is evolved during this reaction. If the energy evolved is shown separately it is expressed as $\Delta H = -571.6 \text{ kJ}$. The same amount of energy will be absorbed when the reaction will move in the backward direction i.e. water will decompose to give hydrogen and oxygen back.

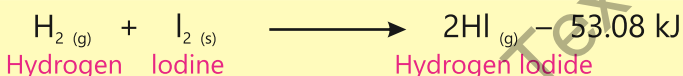
Example: Carbon dioxide gas is produced when solid carbon burns in oxygen gas.



It is also an exothermic reaction and 393.5 kJ heat energy is evolved during this reaction. When this reaction is carried out in the reverse direction, the same amount of energy i.e. 393.5 kJ will be absorbed. The enthalpy change (ΔH) for this reaction is $-393.5 \text{ kJ mol}^{-1}$.

Both the reactions mentioned above are the examples of exothermic changes.

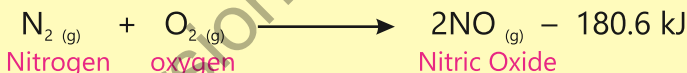
The following reactions represent endothermic changes.



The enthalpy change for the reaction is $\Delta H = 53.08 \text{ kJ}$

Hydrogen gas reacts with solid iodine only at high temperature and 53.08 kJ of heat energy is absorbed.

Formation of NO in air due to lightning in the clouds.



The enthalpy change for the reaction is $\Delta H = 180.6 \text{ kJ}$.



Interesting Information!

Heat evolved or absorbed during a reaction is used in self-heating or self-cooling packs. These packs contain reactants that undergo an exothermic or an endothermic reaction providing high or low temperature.

Our present-day living conditions depend heavily on the availability of energy in its various forms. Exothermic chemical reactions are extensively used to fulfill this requirement. In such reactions, chemical energy is converted into heat energy. We burn fuels like gas, oil and coal to cook food and for other heating purposes in our homes and industry. During this burning process called

combustion, substances present in fuels react with oxygen of the air to produce a certain amount of heat.

Foods such as fats and carbohydrates are important biological fuels. During metabolism, the chemical energy present in this food is converted to heat to keep us warm.

A large portion of electricity is produced at power stations by burning fuels such as natural gas, oil and coal. The heat which comes out from their combustion is used to produce steam at high pressure. This high pressure steam is then used to rotate turbines, which in turn generate electricity.

While driving a vehicle, it is the combustion of petrol or diesel that gives off energy and drives it forward. The one example of exothermic reactions people seem to enjoy the most is that of fireworks. Fireworks are the result of combustion reactions that yield heat, light and sound. Different metal salts alongwith oxidising agents produce a variety of colours when burnt, Fig. (5.2).



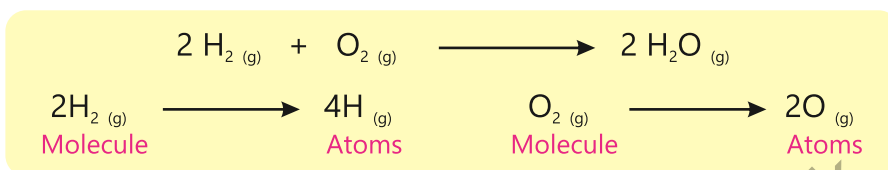
Fig (5.2) Fireworks

Exercise

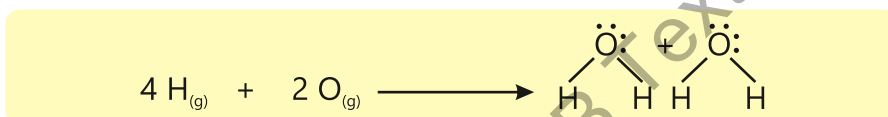
1. Why the chemical reaction between sodium metal and water proceeds violently?
2. Is melting of ice an exothermic or endothermic change?
3. Can exothermic reaction be reversed?

Let us now examine the reason why the chemical reactions are either exothermic or endothermic.

A chemical reaction mainly involves the processes which involve bond breaking and bond formation. In the following reaction, the chemical bonds between the atoms present in the molecules of H_2 and O_2 first break to give their atoms.



These atoms of hydrogen then form bonds with oxygen atoms to form two molecules of gaseous H_2O .



Breaking of bonds of H_2 and O_2 absorb energy (endothermic process) while making of bonds between H and O evolve energy (exothermic process). In this reaction, weaker bonds are broken i.e. of H_2 . Hence less energy is absorbed in the system. While the bonds which are formed in water molecule are stronger and thus greater energy is evolved. Hence, the energy which is evolved is more than the energy which is absorbed. The overall reaction is thus exothermic.

5.4 How does a Reaction take place?

A reaction takes place when the reactant molecules collide with each other to give a transition state. Let us study the following hypothetical reaction.



Before mixing, the molecules of reactants A and B are in a state of random motion separately colliding with each other and with the walls of container. Kinetic energies possessed by these molecules are not the same. Majority of these molecules possess average kinetic energy but a few possess more than average energy while yet others possess less than average kinetic energy. The molecules which possess more than average kinetic energy may also be called excited molecules.

When the two reactant molecules are mixed together, all these molecules start colliding with each other. The collisions which result by

colliding molecules having average or less than average kinetic energies may not be able to produce any result. But when the two excited molecules from both the reactants collide with each other they may be able to produce the transition state as shown in the following Fig (5.3).

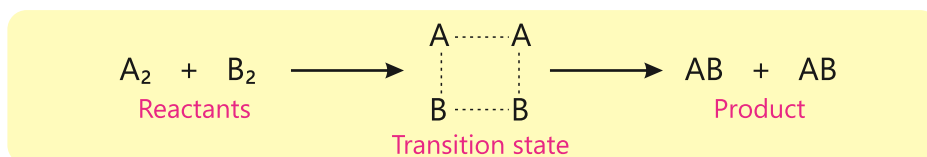


Fig (5.3) Formation of Transition state

The transition state is shown at the peak of curve in Fig (5.4). After a very short period of time the transition state either returns to the reactants or to the products. The progress of the reaction can be shown in the form of the following energy profile diagram drawn between path of the reaction and the energy of the reactants and the products.

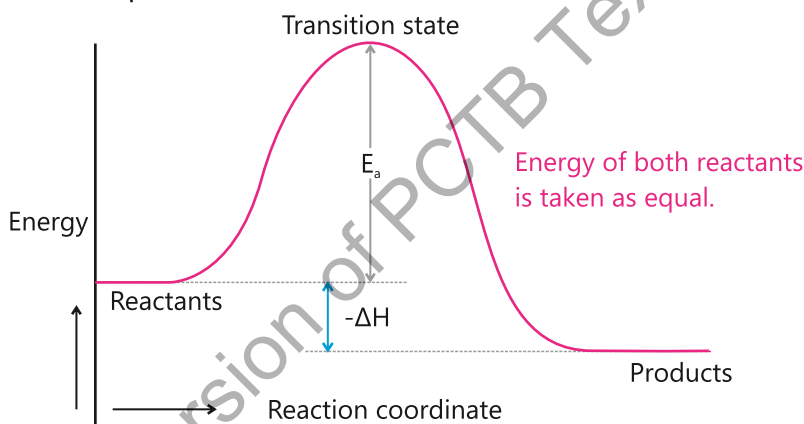


Fig (5.4) Path of Exothermic Reaction



Interesting Information!

Washing clothes at 60°C uses almost twice the energy as at 30°C wash. 90% of the energy used by the traditional electric bulb is wasted in producing heat.

The energy of the transition state is higher than that of reactants or products because the bonds between the reactant or product molecules are being cleaved progressively. The energy absorbed by the reactant or product molecules in order to be converted into the transition state is called the activation energy (E_a) of the reaction. The difference between the energy of reactant and that of the product comes out in the form of heat representing enthalpy (ΔH) of the reaction. This graph represents the path of an exothermic reaction Fig (5.4). A similar graph can be drawn for an endothermic reaction Fig (5.5).

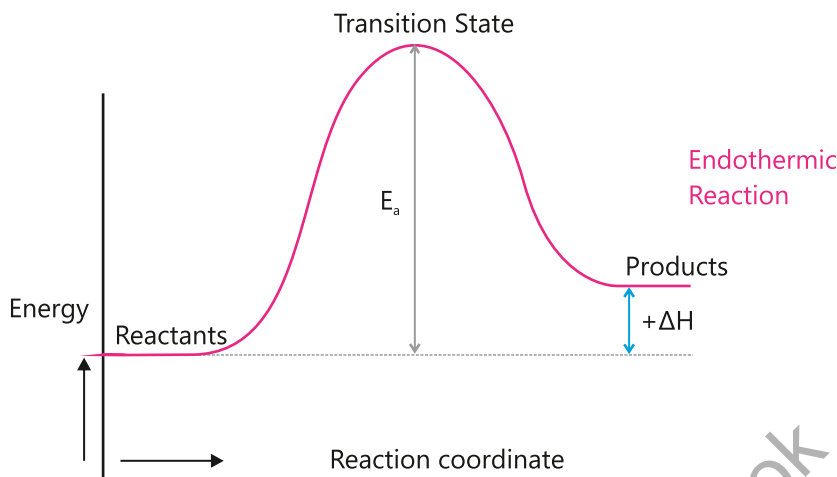


Fig (5.5) Path of Endothermic Reaction

Exercise

1. Are energy diagrams useful?
2. Draw an energy profile diagram for a hypothetical reaction which does not evolve or absorb heat.

An addition of the catalyst in a reaction increases the rate of reaction because it changes the path adopted by the reactants whereby the activation energy value of the reaction is substantially decreased. As a result, more reactants are now able to be converted into product molecules and hence the rate of reaction will increase Fig (5.6).

A catalyst is thus, defined as a substance that increases the rate of a chemical reaction without itself undergoing any permanent chemical change. **For example**, Ni acts as a catalyst in the hydrogenation of oil to give banaspati ghee. Platinum acts as a catalyst in the production of H_2SO_4 . Chlorine acts as a catalyst promoting the breakdown of ozone.

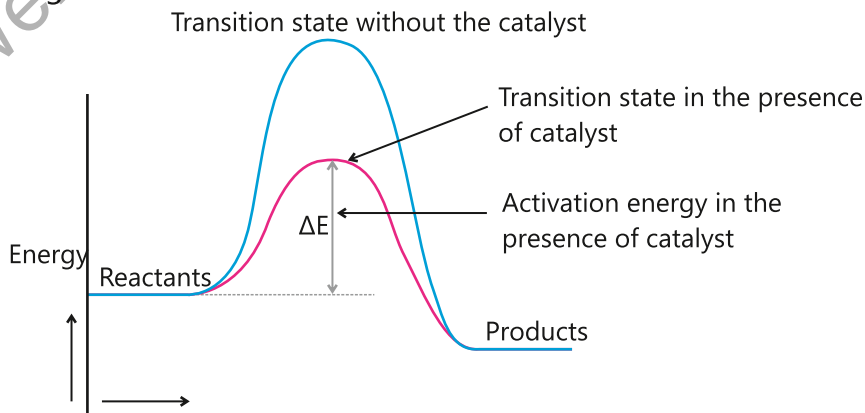


Fig (5.6) Path of Reaction in the absence and presence of a catalyst

Exercise



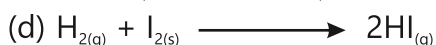
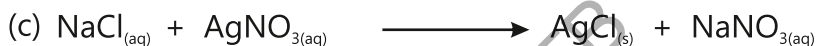
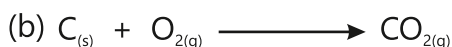
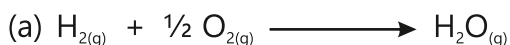
1. Tick (✓) the correct answer.

- (i) The following reaction is an exothermic reaction.



From where does the energy come to break the bonds of H_2 and Cl_2 ?

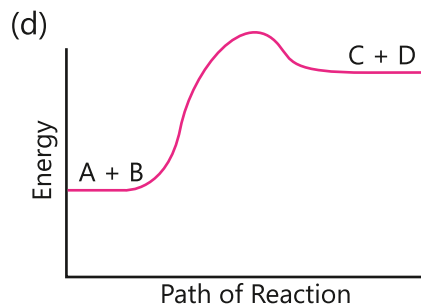
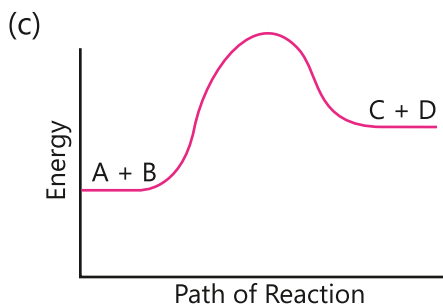
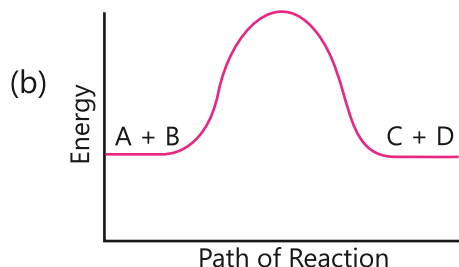
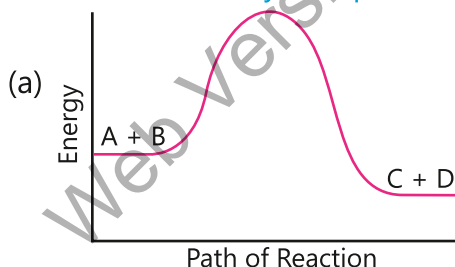
- (a) By collisions between the molecules
 (b) From sunlight
 (c) From the surrounding
 (d) By collisions of the molecules with the walls of the container
- (ii) Which of the following reactions has the least value of activation energy?



- (iii) Formation of which hydrogen halide from the elements is an endothermic reaction?

- (a) HCl
 (b) HF
 (c) HBr
 (d) HI

- (iv) Which reaction do you expect to be a reversible reaction?



- (v) When NaOH and HCl are mixed the temperature increases. The reaction is:
- endothermic with a positive enthalpy change.
 - endothermic with a negative enthalpy change.
 - exothermic with a positive enthalpy change.
 - exothermic with a negative enthalpy change.
- (vi) Why does the following exothermic reaction not occur?
- $$\text{C (Diamond)} \longrightarrow \text{C (Graphite)} \quad \Delta H = -3\text{kJ mol}^{-1}$$
- Structure of diamond is more stable than that of graphite.
 - Diamond has strong covalent bonds than does the graphite.
 - The change from diamond to graphite has high activation energy.
 - Density of graphite is less than that of diamond.

2. Questions for Short Answers

- What is the difference between enthalpy and enthalpy change?
- Why is breaking of a bond an endothermic process?
- Depict the transition state for the following reaction.
$$\text{H}_2 + \text{Cl}_2 \longrightarrow 2\text{HCl}$$
- Draw the reaction profiles for two exothermic reactions one of which moves faster than the other.

3. Constructed Response Questions

- Explain why the reaction between atmospheric gases oxygen and nitrogen does not take place under normal conditions? But in the presence of lightning these gases react to give NO. The reaction stops as soon as the lightning stops.
- A reaction between natural gas (CH₄) and atmospheric oxygen does not take place when you mix them. As soon as you show a burning match stick, the reaction starts immediately and then it continues until one or both of the reactants is/are used up. Explain.

4. Descriptive Questions

- Explain the difference between the terms heat and enthalpy.
- Explain giving examples of exothermic and endothermic reactions.
- How does a reaction take place?
- Explain the following terms:
Activation energy, Transition state, system and surrounding.

Equilibria

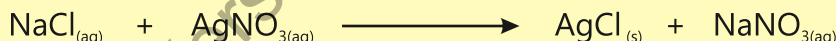
Student Learning Outcomes

After studying this chapter, students will be able to:

- Recognize that reversible reactions are shown by symbol and may not go to completion
- Describe how changing the physical conditions of a chemical equilibrium system can redirect reversible reactions (Some examples can include: **a.** effect of heat on hydrated compounds **b.** addition of water to anhydrous substances in particular copper (II) sulfate and cobalt (II) chloride)
- State that reversible reactions can achieve equilibrium in a closed system when rate of forward and backward reactions are equal.

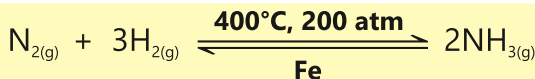
6.1 Reversible and Irreversible Changes

In a chemical reaction, the reactants react to give the products. The reaction will continue until all the reactants or one of the reactants is converted into product. For example, the following reaction takes place immediately in an aqueous solution to give the products.



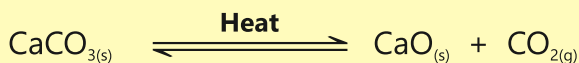
The reaction goes to completion and if stoichiometric amounts of the reactants are used then no reactant species are present at the end of the reaction. Such a reaction is called an **irreversible reaction**. It moves in the forward direction only.

In the majority of chemical reactions, however, the reaction does not go to completion. The products of the reaction react among themselves to give back the reactants under the same conditions.. Such a reaction moves in both the forward and backward directions under the same conditions. The reactants react to give the products and the products, in turn, react to give back the reactants. The reaction is called a **reversible reaction** and it is denoted by a symbol $\rightleftharpoons$ which has two half arrowheads one pointing in each direction. For example,



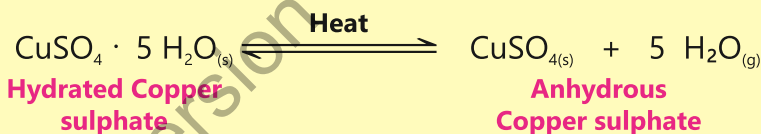
In this reaction, one mole of nitrogen gas reacts with three moles of hydrogen gas under the given reaction conditions in a closed container to give two moles of ammonia gas. After its formation, the ammonia gas decomposes to give the reactants back. The reaction never goes to completion. At any time, all the three species are simultaneously present in the reaction mixture.

A reversible reaction, however, goes to completion if either one of the products is withdrawn from the reaction mixture or being a gas, it escapes into the atmosphere. For example, calcium carbonate is decomposed by heating at a particular temperature.



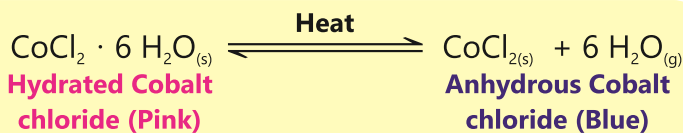
If the above reaction is carried out in an open container, the carbon dioxide gas will escape into the atmosphere as soon as it is formed and the reaction is forced to go to completion. If, on the other hand, the reaction is performed in a closed container, the carbon dioxide will react with calcium oxide to give back the reactants.

Like chemical changes, physical changes may also be reversible in nature. Copper sulphate pentahydrate ($\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$) is an important salt which is blue in colour. When this salt is heated strongly, its colour changes to white. This physical change involves the following equilibrium.

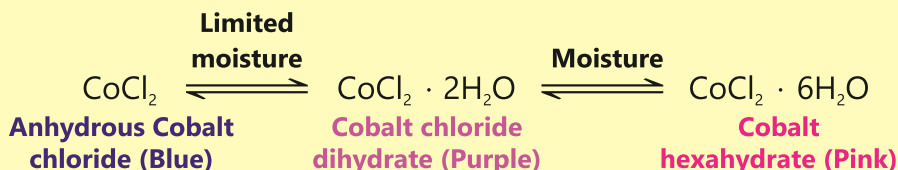


When white anhydrous copper sulphate absorbs moisture from the atmosphere, it will turn blue again.

Similar to this, when cobalt chloride hexahydrate ($\text{CoCl}_2 \cdot 6 \text{H}_2\text{O}$), which is pink in colour, is heated it is converted to anhydrous CoCl_2 which is blue in colour.



In the reverse reaction, the anhydrous cobalt chloride absorbs less moisture, it is first converted into a dihydrate which is purple in colour. This dihydrate then further absorbs four more water molecules to become a hexahydrate which is pink in colour.



Activity

Take a few grams of blue coloured copper sulphate in a dry test tube. First heat it gently and then strongly for some time.

Note down the observation. Let the test tube cool down. Again note down the observation.

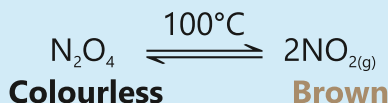
6.2 Dynamic Equilibrium

If a reversible reaction is started by mixing the reactants, the reaction moves in the forward direction only. After some time when enough concentrations of the products are built up, they react to give back the reactants in the reverse reaction.

The reaction will keep on going in both the directions until the rate of forward reaction becomes equal to the rate of backward reaction. In other words, the number of reactant molecules which will disappear as a result of forward reaction becomes equal to the number of reactant molecules which will form as a result of the **reverse** reaction. The same will be true for the product molecules. At this stage, the reaction is said to be in a state of **chemical equilibrium**. It appears as if nothing is going on in the reaction vessel as the concentrations of both reactant and product molecules do not undergo any change at this stage. Since the reaction did not cease at this **state of equilibrium**, rather it keeps on going in both the directions, this state is called **dynamic equilibrium**. The concentrations of reacting species (reactants and products) remain constant at equilibrium.

Exercise

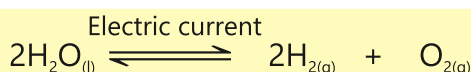
1. Elaborate an example of dynamic equilibrium which exists in this world between the three physical states of water.
2. Dinitrogen tetraoxide (N_2O_4) is a colourless gas. It slowly changes to brown coloured nitrogen dioxide (NO_2) at 100°C . Predict how the colour of the mixture will change if N_2O_4 is kept in a sealed flask at 100°C .



The time a reaction will take to attain the state of dynamic chemical equilibrium depends upon the nature of the reaction and the conditions at which the given reversible reaction is performed.

For a dynamic equilibrium to be set up, the rate of the forward reaction must be equal to the rate of backward reaction. This does not happen instantly and for very slow reaction, it may take years. Both the rates of formation and the decomposition of ammonia are reasonably fast at around 400°C in the presence of a catalyst. This reaction will reach the equilibrium state within minutes after the start of reaction.

The following equilibrium reaction takes 4 – 5 seconds to reach at the point of equilibrium.

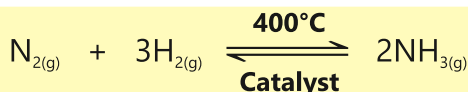


6.3 Changing the Physical Conditions of a Chemical Reaction

If a given reversible reaction has attained the state of dynamic equilibrium, it will remain in this state for infinite time unless it is somehow disturbed. A reversible chemical system may be disturbed in the following possible ways.

1. Adding or withdrawing one or more of the reacting species
2. Adding or withdrawing one or more of the product species
3. Changing the temperature of the reaction
4. Effect of the presence of a catalyst on a reversible reaction
5. Changing the pressure of the reaction if it involves reactants or the products in the gaseous state

Consider the following reversible reaction at equilibrium.

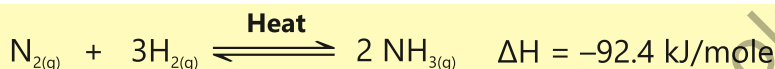


The concentrations of all the participating chemicals will be constant at the state of equilibrium. At this stage if we add more N₂ gas in the mixture, its concentration will increase and the reaction will no longer maintain its state of equilibrium. To restore the equilibrium state again, nitrogen will react with hydrogen to produce more ammonia. This change will go on until a new state of equilibrium is reached at which the concentrations of all the species will again become constant. These new concentrations will, however, be different from the concentrations of the earlier equilibrium state.

Now let us disturb the equilibrium again by withdrawing some of the ammonia gas formed. As a result, its concentration will decrease. To restore the equilibrium state, more nitrogen and hydrogen will react to produce ammonia. When the state of equilibrium is reached again, the concentrations of all the species shall again become constant.

Effect of Changing the Temperature on the State of Equilibrium

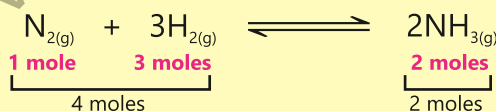
The formation of ammonia is exothermic in the forward direction and hence this reaction will be endothermic in the reverse direction.



If this reaction is at equilibrium and its temperature is increased, the state of the equilibrium will be disturbed again. The ΔH of this reaction is negative. This means that the total energy of the system containing N_2 and H_2 is higher than that of ammonia. The increase in temperature of this reaction at equilibrium will push the reaction in the backward direction i.e. the reactants side. Decreasing the temperature will drive the equilibrium to the forward direction.

Effect of Change of Pressure on the Reaction at Equilibrium

Change of pressure will disturb the equilibrium state of only those gaseous reactions in which the number of moles of the reacting gases will be different from the number of moles of the gases being produced. Since the formation of ammonia gas meets such a condition, the state of its equilibrium will be disturbed by changing the pressure exerted on the reaction mixture.



In this reaction, 4 moles of reacting gases are producing two moles of product gas. If this reaction is at equilibrium and the pressure is increased, the equilibrium will be disturbed. To restore this, the reaction will move to that side in which the number of moles are less i.e. forward direction. The formation of ammonia gas is thus favoured at high pressure.

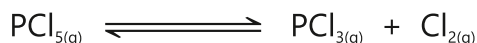
Effect of the Presence of a Catalyst on the Reversible Reaction

At the stage of dynamic equilibrium, the rates of both forward and backward

reactions are equal.

A catalyst increases both the rates of forward and back reactions of a reversible reaction. So if a reversible reaction is carried out in the presence of a catalyst it will decrease the time taken by the reaction to attain the state of equilibrium.

Exercise: Let us consider another example of a reversible reaction at equilibrium. Phosphorous pentachloride decomposes according to the following equation.



According to the equation one mole of the gaseous reactant is giving two moles of the product gases. The reaction is an endothermic reaction. Keeping in view the above description of the reaction, answer the following questions.

- What will happen if the gas mixture is compressed?
- What will happen if we add Cl_2 gas to the equilibrium mixture?
- What will happen if the temperature of the reaction is increased?

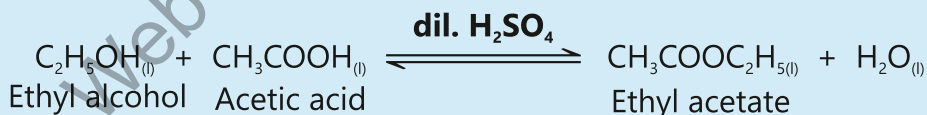


Interesting Information!

Industrial production of ammonia in Haber Process is a very useful application of the phenomenon of chemical equilibrium. Ammonia gas leads to the formation of an important fertilizer urea. The ability of ammonia gas to be converted into its liquid form easily is used to drive the reaction to completion. In this way, practically 100% conversion of N_2 and H_2 to NH_3 is achieved.

Exercise

The preparation of ethyl acetate is commercially very important because it is used as a thinner in paint industry.



One way to get the better yield of the product ethyl acetate is to remove water from the reaction mixture as soon as it is formed. Suggest a suitable method to withdraw water from the reaction mixture.

Exercise



1. Tick (✓) the correct answer.

- (i) What will happen if the rates of forward and reverse reactions are very high?
- (a) The equilibrium point will reach very soon.
 - (b) The equilibrium point will reach very late.
 - (c) The reaction will not attain the state of dynamic equilibrium.
 - (d) The reaction will be practically irreversible.
- (ii) Predict which components of the atmosphere react in the presence of lightning.
- (a) N_2 and H_2O
 - (b) O_2 and H_2O
 - (c) CO_2 and O_2
 - (d) N_2 and O_2
- (iii) CaO or lime is used extensively in steel, glass and paper industries. It is produced in an exothermic reversible reaction by the decomposition of lime (CaCO_3). Choose the conditions to produce maximum amount of lime.
- (a) Heating at high temperature in a closed vessel
 - (b) Heating at high temperature in an open vessel
 - (c) Cooling it in a closed vessel
 - (d) Cooling it in an open vessel
- (iv) What condition should be met for the reversible reaction to achieve the state of equilibrium?
- (a) All the reactants should be converted into the products.
 - (b) 50% of the reactants should be converted into products.
 - (c) The concentrations of all the reactants and the products should become constant.
 - (d) One of the products should be removed from the reaction mixture.
- (v) Why the gas starts coming out when you open a can of fizzy drink?
- (a) Because the solubility of the gas increases
 - (b) Because the gas is insoluble in water
 - (c) Because the gas is dissolved under pressure hence it comes out when pressure is decreased
 - (d) Because the solubility of the gas decreases at high pressure
- (vi) The following reaction is performed in a closed vessel.



How the equilibrium will be affected if you increase the pressure?

- (a) The forward reaction will be favoured
- (b) The backward reaction will be favoured
- (c) No effect on backward reaction
- (d) No effect on forward or backward reaction

(vii) Is reversible reaction useful for preparing compounds on large scale?

- (a) No
- (b) Yes
- (c) They are useful only when equilibrium lies far to the right side
- (d) They are useful only when equilibrium lies far to the left side

2. Questions for Short Answers

- i. How is dynamic equilibrium different from the static equilibrium?
- ii. How the following reversible reaction will be affected if its temperature is increased?



- iii. How can you get the maximum yield in a reversible reaction?
- iv. How can you decrease the time to attain the position of equilibrium in a reversible reaction?
- v. What is the effect of increasing pressure on the following reaction at equilibrium?



3. Constructed Response Questions

- i. Why are some reactions irreversible while others are reversible?
- ii. Why are combustion reactions generally irreversible?
- iii. How do you know if a reaction is reversible or irreversible?

4. Descriptive Questions

- i. How can you drive the reversible reaction at equilibrium?
 - (a) in the forward direction
 - (b) in the backward direction
- ii. Explain how the rates of forward and backward reactions change when the system approaches equilibrium.
- iii. Describe the effect of a catalyst on the reversible reaction.
- iv. How can a reversible reaction be forced to go to completion?
- v. How does change in temperature affect the reaction at equilibrium?

Acid Base Chemistry

Student Learning Outcomes

After studying this chapter, students will be able to:

- Define Bronsted-Lowry acids as proton donors and Bronsted-Lowry bases as proton acceptors
- Recognize that aqueous solutions of acids contain H^+ ions and aqueous solutions of alkalis contain OH^- ions
- Define a strong acid and base as an acid or base that completely dissociates in aqueous solution and weak acid and base that partially dissociates in aqueous solution. (Some examples include: Student writing symbol equations to show these for hydrochloric acid, sulphuric acid, nitric acid, and ethanoic acid.)
- Formulate dissociation equations for an acid or base in aqueous solution.
- Recognize that bases are oxides or hydroxides of metals and that alkalis are water-soluble bases
- Describe the characteristic properties of acids in terms of their reactions with metals, bases and carbonates
- Identify the characteristic properties of bases in terms of their reactions with acids and ammonium salts

7.1 Acids and Bases

Acids and bases have been known to mankind since centuries. Acids have been known for their sour taste like the taste of lemon and ability to change the colour of litmus paper from blue to red. There are many substances which contain acids and hence taste sour, such as curd, tamarind, lemon and lime, etc.

Common examples of acids include acetic acid, hydrochloric acid, nitric acid, sulphuric acid and tartaric acid. Most of these acids, we come across in everyday life, are available in the form of aqueous solutions as they can be easily dissolved in water.

Acids are divided into two types on the basis of their occurrence: natural acids and mineral acids. Acids which are obtained from natural sources are called natural or organic acids. Mineral or man-made acids are prepared from minerals like sodium chloride or sodium nitrate. Common examples of mineral acids are hydrochloric acid, sulphuric acid and nitric acid.

Some common organic acids and their natural sources are given in Table (7.1)

Table (7.1) Some Common Organic Acids and their Sources	
Organic acid	Natural Source
Acetic acid	Vinegar
Ascorbic acid	Amla, Guava
Citric acid	Lemon, Orange
Lactic acid	Sour milk, Curd
Formic acid	Ant-sting
Oxalic acid	Tomato
Tartaric acid	Tamarind

In comparison to acids, bases are known for their bitter taste, slippery touch and ability to change the colour of litmus paper from red to blue. An alkali is a base that dissolves in water. Some common examples of alkalis are sodium hydroxide (caustic soda), potassium hydroxide (caustic potash) calcium hydroxide (lime water) and aqueous ammonia (NH_4OH). Both sodium hydroxide and potassium hydroxide are extremely corrosive and can burn your skin.

Metal oxides are also basic in nature because they react with acids to form salts and water. Na_2O is basic oxide because it contains oxide ion, O^{2-} , which is a very strong base with a strong tendency to react with water to produce hydroxide ions.



Other examples of basic oxides are calcium oxide (CaO), zinc oxide (ZnO) and magnesium oxide (MgO).

Both acids and bases are known to cancel the properties of each other when mixed together in equal amounts. The reaction is called neutralization reaction. A salt and water are formed as a result of this reaction.

Exercise Name some fruits which contain citric acid.



Interesting Information!

Oxalic acid ($\text{H}_2\text{C}_2\text{O}_4$) is the simplest organic diprotic acid. Its commercial uses include bleaching straw and leather and removing rust and ink stains from fabrics.

Activity 7.1

The following compounds are provided in the form of liquid or in the form of their aqueous solutions. Use blue and red litmus paper strips to show whether these substances are acidic or basic in their aqueous solutions.

Substance	Acidic	Basic
Tap water		
Battery water		
Rain water		
Soap solution		
Tooth paste		
Shampoo		
Bleach		

Exercise

In what ways are mineral acids useful for us?

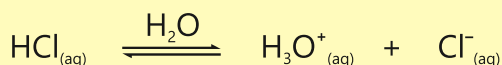
7.2 Different Concepts of Acids and Bases

Arrhenius Acids and Bases

Although the earlier definitions of acids and bases describe some distinctive features of these substances, yet new and broader definitions were needed to explain their chemical behavior on the molecular level.

Svante Arrhenius, a Swedish Chemist, suggested that acids and bases may be classified in terms of their behavior in water. According to him;

An acid is that substance which dissociates in water to give proton (H^+) or hydroxonium ion (H_3O^+). Some typical Arrhenius acids are HCl , HNO_3 , H_2SO_4 and HCN .

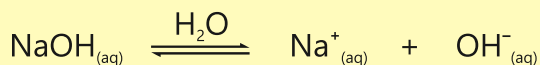


Similarly, a base is that substance which dissociates in water to give hydroxyl ions (OH^-).

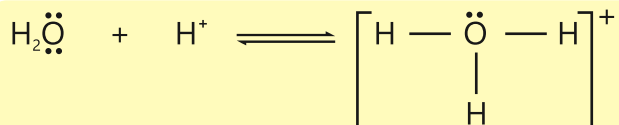


Svante Arrhenius
(1859-1927)

Some typical Arrhenius bases are NaOH, KOH and Ba(OH)₂.

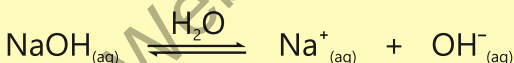
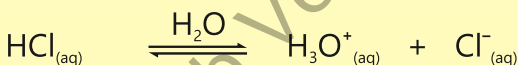


Water has an essential role to play in Arrhenius concept of acids and bases. Whenever an acid or a base dissociates in water, its molecules participate in reaction by surrounding the resultant proton (H⁺) and hydroxyl ion(OH⁻). Since proton is very small in size and its charge density is very high, it forms a strong bond with the lone pair of water molecule to give hydroxonium ion, H₃O⁺.



According to some reports upto four water molecules may surround the proton. Similarly, hydroxyl ions are also surrounded by water molecules as shown in the Figure (7.1).

Arrhenius also explained the process of neutralization. According to him when a strong acid and a strong base are dissolved in water, they completely dissociate into ions.



Exercise

How do chloride ions exist in water?

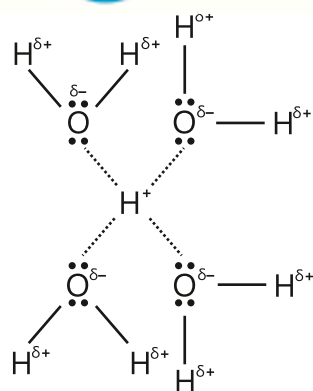
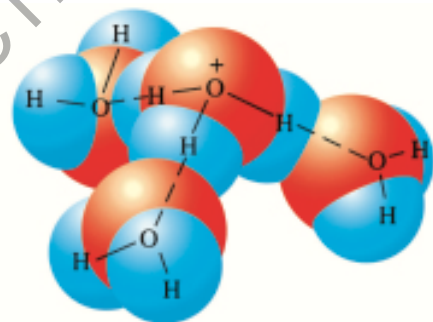
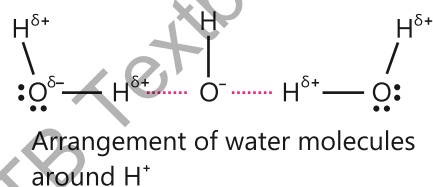
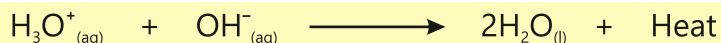


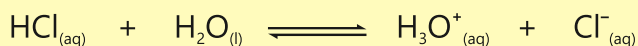
Fig (7.1)

The hydroxonium ion and the hydroxyl ion then react to form water with the evolution of heat.

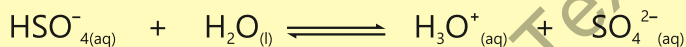
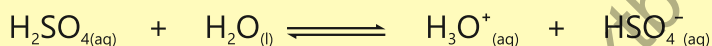


The salt NaCl that is formed with water does not exist as solid crystals. It remains present in the solution in the form of hydrated spectator sodium ions (Na^+) and chloride ions (Cl^-).

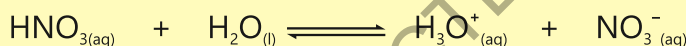
Mineral acids are generally very strong acids. The strength of an acid depends upon the extent to which it is ionized in water. Hydrochloric acid ionizes in water completely giving a large amount of hydroxonium ions in water.



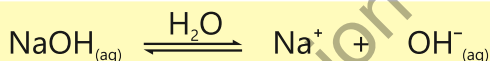
Sulphuric acid being a diprotic acid ionizes in two steps.



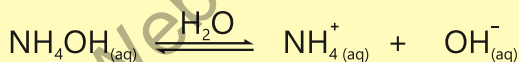
Nitric acid ionizes in single step.



NaOH and KOH are the examples of strong bases because they also ionize completely in water.



NH_4OH and $\text{Al}(\text{OH})_3$ are the examples of weak bases because they only partially ionize in water.



Exercise

Why does ammonium hydroxide only partially ionize in water?

7.3 Bronsted – Lowry concepts of Acids and Bases

Arrhenius concept of acids and bases is very simple and easily understandable but it has its own limitations. For example, Na_2CO_3 , K_2CO_3 and NH_3 do not contain any hydroxyl group which will get ionized by water but all these compounds behave as bases and yield OH^- in water. Arrhenius definitions also require that water must be present as a solvent in order for a compound to behave as an acid or a base. It was realized that a broader definition for acids and bases was needed to cover all the aspects of the concept.

To remove the limitations of Arrhenius concept, Bronsted and Lowry gave the following definitions of acids and bases:

An acid is a substance that donates a proton.

This definition requires that to behave as an acid a compound must have a proton to donate. The condition of the presence of water during this donation was, however,

eliminated. All Arrhenius acids are, thus, Bronsted-Lowry acids as well for example, HCl. It dissociates in water to give H^+ and Cl^- . It also donates H^+ to H_2O forming H_3O^+ .

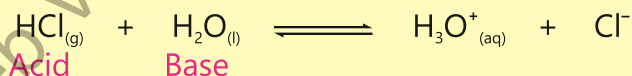
A base is a substance that accepts a proton.

For example, OH^- , CN^- , NH_3 and Cl^- are all bases because they have the ability to accept a proton. Note that except OH^- all other species are not Arrhenius bases. All Arrhenius bases are, however, Bronsted-Lowry bases as well.

All Bronsted-Lowry acids and bases are not Arrhenius acids and bases. NH_4^+ is not Arrhenius acid and NH_3 is not Arrhenius base.

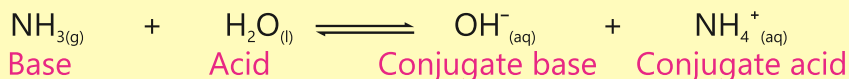
According to Bronsted-Lowry, an acid base reaction is that reaction in which a proton is transferred from a proton donor to its acceptor. This reaction may take place in gas phase or in the presence of any solvent.

Consider the following reaction between hydrogen chloride gas and liquid water.



In this reaction, HCl gas acts as an acid because it donates its proton to water which acts as a base.

Similarly when ammonia gas dissolves in water, a proton is transferred from water to ammonia and ammonium ion is formed.



Ammonia is a base while water is an acid in this reaction. Water has the ability to act both as an acid or a base depending upon the other compound with which it reacts.



Johannes N. Bronsted
(1879-1947)

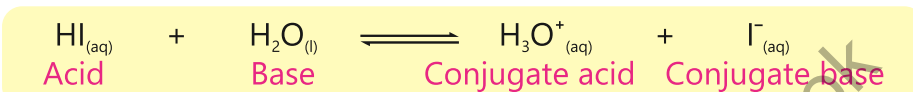
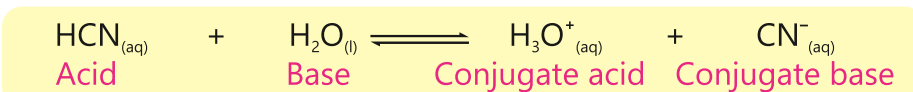


Thomas M. Lowry
(1874-1936)

Water is therefore called an **amphoteric** compound which means a compound that can behave both as an acid and a base.

In the reverse reaction, OH^- is a base because it accepts a proton donated by the acid NH_4^+ . In order to differentiate, OH^- is called the conjugate base while NH_4^+ the conjugate acid.

Some other examples of Bronsted-Lowry acids and bases.



Exercise

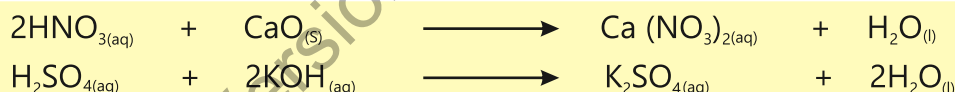
Give two examples of Bronsted-Lowry bases which are not bases by Arrhenius definition.

7.4 Properties of Acids and Bases

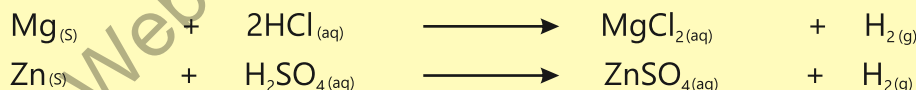
Properties of acids and bases vary depending upon the substances they react with. For example, acids react with metals to produce hydrogen gas whereas bases do not react with metals.

a. Acids give the following three types of reactions.

1. With alkalis or metal oxides, they form salts and water.



2. With reactive metals (Mg, Zn) they form salts and evolve hydrogen gas.

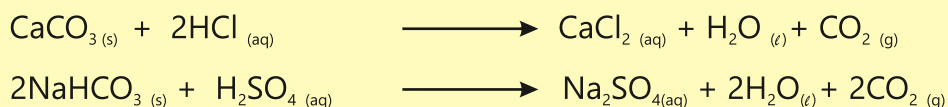


The less reactive metals Cu, Ag, Au, and Pt do not evolve hydrogen gas with acids.

Activity 7.2

Take a few granules of zinc in a clean test tube. Add dilute H_2SO_4 in it and heat it gently. Identify the gas evolved in this reaction by taking a burning match stick near the mouth of test tube. Note the observations in your notebook.

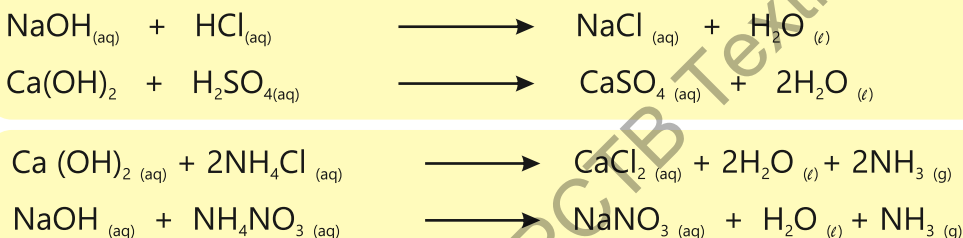
3. They decompose metal carbonates and hydrogen carbonates evolving carbon dioxide gas.



b. Properties of Bases

As already mentioned alkalis are those bases which are soluble in water. Examples of alkalis are NaOH, KOH etc. $\text{Ca}(\text{OH})_2$ is sparingly soluble in water while $\text{Cu}(\text{OH})_2$ is insoluble.

Apart from reacting with acids in neutralization reactions, alkalies also react with ammonium salts and liberate ammonia gas



Exercise



1. Tick (✓) the correct answer.

- (i) Which acid is not used as a food or mixed with food?
 - (a) Tartaric acid
 - (b) Ascorbic acid
 - (c) Citric acid
 - (d) Formic acid
- (ii) While baking, which gas is responsible for raising the bread and making it soft?
 - (a) Oxygen
 - (b) Carbon dioxide
 - (c) Nitrogen
 - (d) Carbon monoxide
- (iii) How many hydroxide ions, calcium hydroxide will release in water?
 - (a) 1
 - (b) 2
 - (c) Zero
 - (d) 3
- (iv) In a neutralization reaction between KOH and H_3PO_4 , how many molecules of KOH will react with one molecule of H_3PO_4 ?
 - (a) 2
 - (b) 1
 - (c) 3
 - (d) 4
- (v) Which compound is formed when SO_2 is dissolved in water?
 - (a) SO_3
 - (b) H_2SO_3
 - (c) H_2SO_4
 - (d) $\text{H}_2\text{S}_2\text{O}_7$

- (vi) Which compound in the following reaction is behaving as a conjugate base?



2. Questions for Short Answers

- Choose Arrhenius Acids among the following compounds.
 HF , NH_4^{+} , H_2SO_3 , SO_3 , H_2S , H_2O
- How does calcium metal react with dilute H_2SO_4 ?
- Which salt is formed when HCl reacts with BaCO_3 ?
- How will you justify that HSO_4^{-} is a Bronsted – Lowry acid?
- Why is HCl not edible although it is present in the stomach and responsible for digestion of food?

3. Constructed Response Questions

- What chemical name will you give to soap as a compound?
- In the presence of a drop of an acid, water is known to ionize as follows:



In your opinion, which name will be suitable for water: an acid, a base or both?

- Why does Na_2CO_3 behave like a base in water?

4. Descriptive Questions

- Explain Arrhenius concept of acids and bases.
- Compare Arrhenius and Bronsted – Lowry concepts of acids and bases.
- How does sulphuric acid react with the following compounds?
 NH_4Cl , NH_4NO_3 , MgO , MgCO_3
- What happens when a base reacts with a non-metallic oxide. What do you infer about the nature of non-metallic oxide?
- State the reason of showing acidic character by both dry HCl gas and HCl solution in water.
- Differentiate between an acid and its conjugate base.

Periodic Table and Periodicity

Student Learning Outcomes

After studying this chapter, students will be able to:

- Define the periodic table as an arrangement of elements in periods and groups, in order of increasing proton number/atomic number
- Identify the group or period or block of an element using its electronic configuration (only the idea of subshells related to the blocks can be introduced)
- Explain the relationship between group number and the charge of ions formed from elements in the group in terms of their outermost shells
- Explain similarities in the chemical properties of elements in the same group in terms of their electronic configuration
- Identify trends in group and periods, given information about the elements, including trends for atomic radius, electron affinity, electronegativity, ionization energy, metallic character, reactivity and density
- Use terms alkali metals, alkaline earth metals, halogens, noble gases, transition metals, lanthanides and actinides in reference to the periodic table
- Predict the characteristic properties of an element in a given group by using knowledge of chemical periodicity
- Deduce the nature, possible position in the Periodic Table and the identity of unknown elements from given information about their physical and chemical properties

8.1 Modern Periodic Table

The modern periodic table is based upon the arrangement of elements according to increasing atomic number. When the elements are arranged according to ascending order of their atomic numbers from left to right in a horizontal row, properties of elements are found repeating after regular intervals. This results in the form of a table in which elements of similar properties are placed in the same vertical columns.

The horizontal rows of elements in the periodic table are called periods while the vertical columns are called groups.

In the modern periodic table, the electronic configuration of the elements continues changing when we move from left to right in a period. Due to this, the elements in a period show a gradual change in the properties. Against this, elements present in a group possess similar electronic configurations in their outermost shells. Therefore, the elements in a group show similar chemical properties. As a whole the periodic table shows the repetition of the properties of elements after regular intervals. This trend is called periodicity in the properties.

8.2 Salient Features of Modern Periodic Table

Periods

The horizontal rows in the modern periodic table are called periods. There are seven periods in total. Each period except the first starts with an alkali metal and ends at a noble gas. Each period also represents the completion of a shell. Since the number of electrons to be accommodated in a particular shell is fixed, the number of elements in a period is also fixed. The detail about the periods is given in Table (8.1).

Table (8.1) Periods in the Periodic Table

Period No.	Name of the Period	Number of Elements	Number of Shell being filled
1st	Short	2	1st
2nd	Normal	8	2nd
3rd	Normal	8	3rd
4th	Long	18	4th
5th	Long	18	5th
6th	Very long	32	6th
7th	Very long	32	7th

In 6th and 7th periods, two series of fourteen elements each have been accommodated. Because of the space problem, these two series were placed at the bottom of the periodic table to keep it in a manageable and presentable form. The first series starts after lanthanum (La = 57) and it is called **lanthanides**. The second series starts after actinium (Ac = 89) and it is called **actinides**. Together the elements present in these two series are also called **rare earths or f-block elements**.

Groups

The vertical columns present in the periodic table are called groups. There are in total eighteen groups. Elements present in a group resemble one another in their chemical properties since they contain the same number of electrons in their outermost shells. Elements present in a group are also called a family and each group has also been given a family name. The distribution of electrons in the outermost shells (electronic configuration) and other information about the groups are given in the following Table (8.2).

Table (8.2) Electronic Configurations of Elements In The Outermost Shell

Group No.	Family Name	Electronic Configuration In the Outermost Shell
1	Alkali metals	ns^1
2	Alkaline earth metals	ns^2
3 to 12	Transition metals	$nd^x ns^2$
13	Boron family	$ns^2 np^1$
14	Carbon family	$ns^2 np^2$
15	Nitrogen family	$ns^2 np^3$
16	Oxygen family	$ns^2 np^4$
17	Halogen family	$ns^2 np^5$
18	Noble gases	$ns^2 np^6$

The groups 1 to 2 and 13 to 17 contain the normal elements. In the normal elements, all the inner shells are completely filled while the outermost shell is incomplete. The groups 3 to 12 are called transition elements and in these elements the inner sub-shells are in the process of completion.

Elements present in the periodic table are also classified into blocks. It depends upon the type of the subshell which is being filled; s, p, d and f. Elements of group 1 and 2 are called **s-block elements** because in them s-subshell of the outermost shell is being filled. Similarly, elements present in groups 13 to 18 are called **p-block elements** because p-subshell is filled in these elements. The **d-block elements** lie between the s and p blocks, while **f-block elements** in the form of two rows lie at the bottom of the periodic table.



Interesting Information!

Mendeleev arranged elements in his periodic table in the ascending order of their atomic masses. Later on, Moseley arranged the elements in ascending order of their atomic numbers.

PERIODIC TABLE OF THE ELEMENTS

Atomic Number Element Symbol Element Name										30 Zn Zinc										2 He Helium 4.0026									
4 Be Beryllium										9 F Fluorine										10 Ne Neon									
11 Na Sodium										16 S Sulfur										17 Cl Chlorine									
12 Mg Magnesium										33 As Arsenic										34 Se Selenium									
19 K Potassium										48 Cd Cadmium										50 Sn Tin									
20 Ca Calcium										79 Au Gold										80 Hg Mercury									
21 Sc Scandium										109 Mt Meitnerium										110 Ds Darmstadtium									
22 Ti Titanium										108 Hs Hassium										111 Rg Roentgenium									
23 V Vanadium										107 Bh Bohrium										112 Cn Copernicium									
24 Cr Chromium										106 Sg Seaborgium										113 Nh Ununtrium									
25 Mn Manganese										105 Db Dubnium										114 Fl Flerovium									
26 Fe Iron										104 Rf Rutherfordium										115 Mc Ununpentium									
27 Co Cobalt										103 Lr Lutetium										116 Lv Livermorium									
28 Ni Nickel										102 Sg Seaborgium										117 Ts Ununseptium									
29 Cu Copper										101 Db Dubnium										118 Og Ununoctium									
30 Zn Zinc										100 Fm Fermium										119 Uue Ununennium									
31 Ga Gallium										99 Es Einsteinium										120 Uub Unubium									
32 Ge Germanium										98 Cf Californium										121 Uut Ununtrium									
33 As Arsenic										97 Bk Berkelium										122 Uuq Ununquadium									
34 Se Selenium										96 Cm Curium										123 Uuh Ununhexium									
35 Br Bromine										95 Am Americium										124 Uus Ununseptium									
36 Kr Krypton										94 Pu Plutonium										125 Uuo Ununoctium									
37 Rb Rubidium										93 Np Neptunium										126 Uuo Ununoctium									
38 Sr Strontium										92 U Uranium										127 Uuo Ununoctium									
39 Y Yttrium										91 Pa Protactinium										128 Uuo Ununoctium									
40 Zr Zirconium										90 Th Thorium										129 Uuo Ununoctium									
41 Nb Niobium										89 Ac Actinium										130 Uuo Ununoctium									
42 Mo Molybdenum										88 Ra Radium										131 Uuo Ununoctium									
43 Tc Technetium										87 Fr Francium										132 Uuo Ununoctium									
44 Ru Ruthenium										86 Ba Barium										133 Uuo Ununoctium									
45 Rh Rhodium										85 At Astatine										134 Uuo Ununoctium									
46 Pd Palladium										84 Po Polonium										135 Uuo Ununoctium									
47 Ag Silver										83 Bi Bismuth										136 Uuo Ununoctium									
48 Cd Cadmium										82 Pb Lead										137 Uuo Ununoctium									
49 In Indium										81 Tl Thallium										138 Uuo Ununoctium									
50 Sn Tin										80 Hg Mercury										139 Uuo Ununoctium									
51 Sb Antimony										79 Au Gold										140 Uuo Ununoctium									
52 Te Tellurium										78 Pt Platinum										141 Uuo Ununoctium									
53 I Iodine										77 Ir Iridium										142 Uuo Ununoctium									
54 Xe Xenon										76 Os Osmium										143 Uuo Ununoctium									
55 Cs Caesium										75 Re Rhenium										144 Uuo Ununoctium									
56 Ba Barium										74 W Tungsten										145 Uuo Ununoctium									
57 La Lanthanum										73 Ta Tantalum										146 Uuo Ununoctium									
58 Ce Cerium										72 Hf Hafnium										147 Uuo Ununoctium									
59 Pr Praseodymium										71 Yb Ytterbium										148 Uuo Ununoctium									
60 Nd Neodymium										70 Er Erbium										149 Uuo Ununoctium									
61 Pm Promethium										69 Tm Thulium										150 Uuo Ununoctium									
62 Sm Samarium										68 Gd Gadolinium										151 Uuo Ununoctium									
63 Eu Europium										67 Yb Ytterbium										152 Uuo Ununoctium									
64 Gd Gadolinium										66 Lu Lutetium										153 Uuo Ununoctium									
65 Tb Terbium										65 Tb Terbium										154 Uuo Ununoctium									
66 Dy Dysprosium										64 Nd Neodymium										155 Uuo Ununoctium									
67 Ho Holmium										63 Pr Praseodymium										156 Uuo Ununoctium									
68 Er Erbium										62 Sm Samarium										157 Uuo Ununoctium									
69 Tm Thulium										61 Pm Promethium										158 Uuo Ununoctium									
70 Yb Ytterbium										60 Nd Neodymium										159 Uuo Ununoctium									
71 Lu Lutetium										59 Eu Europium										160 Uuo Ununoctium									
72 Hf Hafnium										58 Ce Cerium										161 Uuo Ununoctium									
73 Ta Tantalum										57 La Lanthanum										162 Uuo Ununoctium									
74 W Tungsten										56 Ba Barium										163 Uuo Ununoctium									
75 Re Rhenium										55 Cs Caesium										164 Uuo Ununoctium									
76 Os Osmium										54 Xe Xenon										165 Uuo Ununoctium									
77 Ir Iridium										53 I Iodine										166 Uuo Ununoctium									
78 Pt Platinum										52 Te Tellurium										167 Uuo Ununoctium									
79 Au Gold										51 Sb Antimony										168 Uuo Ununoctium									
80 Hg Mercury										50 Sn Tin										169 Uuo Ununoctium									
81 Tl Thallium										49 In Indium										170 Uuo Ununoctium									
82 Pb Lead										48 Cd Cadmium										171 Uuo Ununoctium									
83 Bi Bismuth										47 Ag Silver										172 Uuo Ununoctium									
84 Po Polonium										46 Pd Palladium										173 Uuo Ununoctium									
85 At Astatine										45 Rh Rhodium										174 Uuo Ununoctium									
86 Rn Radon										44 Ru Ruthenium										175 Uuo Ununoctium									
87 Fr Francium										43 Tc Technetium										176 Uuo Ununoctium									
88 Ra Radium										42 Mo Molybdenum										177 Uuo Ununoctium									
89 Ac Actinium										41 Nb Niobium										178 Uuo Ununoctium									
90 Th Thorium										40 Zr Zirconium										179 Uuo Ununoctium									
91 Pa Protactinium										39 Y Yttrium										180 Uuo Ununoctium									
92 U Uranium										38 Sr Strontium										181 Uuo Ununoctium									
93 Np Neptunium										37 Rb Rubidium										182 Uuo Ununoctium									
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95 Am Americium										35 Br Bromine										184 Uuo Ununoctium									
96 Cm Curium										34 Se Selenium										185 Uuo Ununoctium									
97 Bk Berkelium										33 As Arsenic										186 Uuo Ununoctium									
98 Cf Californium										32 Ge Germanium										187 Uuo Ununoctium									
99 Es Einsteinium										31 Ga Gallium										188 Uuo Ununoctium									
100 Fm Fermium										30 Zn Zinc										189 Uuo Ununoctium									
101 Md Mendelevium										29 Cu Copper										190 Uuo Ununoctium									
102 No Nobelium										28 Ni Nickel										191 Uuo Ununoctium									
103 Lr Lutetium										27 Co Cobalt										192 Uuo Ununoctium									
104 Rf Rutherfordium										26 Fe Iron										193 Uuo Ununoctium									
105 Db Dubnium										25 Mn Manganese										194 Uuo Ununoctium									
106 Sg Seaborgium										24 Cr Chromium										195 Uuo Ununoctium									
107 Bh Bohrium										23 V Vanadium										196 Uuo Ununoctium									
108 Hs Hassium										22 Ti Titanium										197 Uuo Ununoctium									
109 Mt Meitnerium										21 Sc Scandium										198 Uuo Ununoctium									
110 Ds Darmstadtium										20 Ca Calcium										199 Uuo Ununoctium									
111 Rg Roentgenium										19 K Potassium										200 Uuo Ununoctium									
112 Cn Copernicium										18 Ar Argon										201 Uuo Ununoctium									
113 Nh Ununtrium										17 Cl Chlorine										202 Uuo Ununoctium									
114 Fl Flerovium										16 S Sulfur										203 Uuo Ununoctium									
115 Mc Ununpentium										15 P Phosphorus										204 Uuo Ununoctium									
116 Lv Livermorium										14 Si Silicon										205 Uuo Ununoctium									
117 Ts Ununseptium										13 Al Aluminium										206 Uuo Ununoctium									
118 Og Ununoctium										12 Mg Magnesium										207 Uuo Ununoctium									
119 Uue Ununennium										11 Na Sodium										208 Uuo Ununoctium									
120 Uub Unubium										10 Ne Neon										209 Uuo Ununoctium									
121 Uut Ununtrium										9 F Fluorine										210 Uuo Ununoctium									
122 Uuq Ununquadium										8 O Oxygen										211 Uuo Ununoctium									
123 Uuh Ununhexium										7 N Nitrogen										212 Uuo Ununoctium									
124 Uuo Ununoctium										6 C Carbon										213 Uuo Ununoctium									
125 Uuo Ununoctium										5 B Boron										214 Uuo Ununoctium									
126 Uuo Ununoctium										4 Be Beryllium										215 Uuo Ununoctium									
127 Uuo Ununoctium										3 Li Lithium										216 Uuo Ununoctium									
128 Uuo Ununoctium										2 H Hydrogen 1.008										217 Uuo Ununoctium									
129 Uuo Ununoctium										1 H Hydrogen 1.008										218 Uuo Ununoctium									
130 Uuo Ununoctium										0										219 Uuo Ununoctium									
131 Uuo Ununoctium										0										220 Uuo Ununoctium									
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148 Uuo Ununoctium										0										237 Uuo Ununoctium									
149 Uuo Ununoctium										0										238 Uuo Ununoctium									
150 Uuo Ununoctium										0										239 Uuo Ununoctium									
151 Uuo Ununoctium										0										240 Uuo Un									

Atomic Number
Element Symbol
Element Name

30
Zn
Zinc

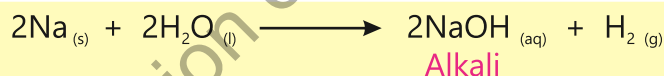
Exercise

- (1) The electronic configuration of the outermost shell of an element is s^2p^3 . Find out period number and the group number of the element. In which block will you place this element?
- (2) What is the group of the element having eight electrons in its outermost shell? In which physical state does this element exist?
- (3) An element belongs to 16th group and it is a gas. To which period does it belong?

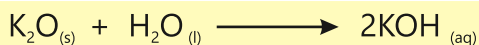
8.3 Similarities in the Chemical Properties of Elements in the Same Group

The chemical properties of elements depend largely upon the number of electrons present in their outermost shells. Since in a group of the periodic table all the elements have the same number of electrons in the outermost shell, they are expected to show similar chemical properties.

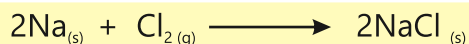
All elements of group I have one electron in their outermost shells, so they show a strong tendency to lose their electron forming cations. They are thus known as electropositive metals. These metals react vigorously with water producing hydrogen and giving alkalis in the solution.



Oxides of these metals are also strongly basic in nature. They are readily soluble in water giving alkalis.



Alkali metals also react with halogens giving halides.

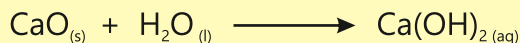


The reactivity of alkali metals gradually increases down the group.

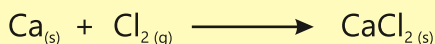
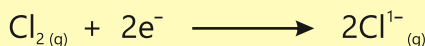
The second group elements also show a tendency to lose both of their outermost electrons forming dipositive ions.



The tendency to lose electrons down the group increases due to gradual increase in their atomic sizes. The oxides of these metals are also basic in nature and form alkalies in water.



All elements in group 17 have s^2, p^5 configuration in their outer shells. They show a strong tendency to gain one electron to become an anion. They are called electronegative elements. These elements are very reactive non-metals and form salts with metals. Hence they are called halogens (salt forming). Unlike alkali metals, the reactivity of halogens decreases down the group.



Elements present in group 16 have s^2, p^4 configuration in their outermost shells, so they have a tendency to accept two electrons to form a dinegative ion. Oxygen behaves as a strong electronegative element.



Relationship between Group Number and Charge of Ions

When we move from left to right in the periodic table, the main-group elements tend to form cations having a charge equal to the group number. For example, group 1 elements form $1+$ ions, group 2 elements forms $2+$ ions and group 3 elements form $3+$ ions. The number of charges on the cations also correspond to the number of electrons present in their outermost shells.

When we move from right to left in the periodic table, elements often form anions with a negative charge equal to the number of group towards the left side of the noble gases. For example, group 17 elements (which are located one group towards left to the noble gases) form ions with one negative charge. Similarly the group 16 elements form ions with two negative charges. The negative charges present on these ions correspond to number of electrons which these groups need to complete their octets.

8.4 Variation of Periodic Properties in Periods and Groups

Periodicity in the properties of the elements occur due to recurrence of similar electronic configuration in the outermost shells. These properties include:

- | | |
|------------------------|-----------------------|
| i) Atomic radius | ii) Ionization Energy |
| iii) Electron affinity | iv) Electronegativity |

Atomic Radius

The electron cloud of an atom has no definite limit. Because of this, the size of an atom cannot be defined exactly. However, it is possible to measure the radius of an atom when it is bonded to an identical atom.

Atomic radius is defined as half the distance between the nuclei of the two identical bonded atoms. It is expressed in pm ($1\text{pm} = 10^{-12}\text{m}$). For example, the distance between the nuclei of two bonded carbon atoms is 154 pm. Half of this distance i.e. 77pm is therefore the radius of carbon atom. This is also called covalent radius of carbon atom.

Variation of Atomic Radius in Periods

When we move from left to right in a period, the size of atoms decreases generally. It is because as we go from lithium (Li) to neon (Ne) in the second period, we are adding electrons to the outermost shell. The charge on the nucleus also increases from +3 to +10. This tends to pull the electrons closer to the nucleus and hence the sizes of atoms decrease from lithium to neon Table (8.4).

Table(8.4) Atomic Radii of Second Period Elements

2nd period elements	Li	Be	B	C	N	O	F	Ne
Atomic Radii (pm)	152	113	88	77	75	73	71	69

Variation of Atomic Radius in Groups

The atomic radii of atoms increase from top to bottom in a group. It is because a new shell is being added in the successive period down the group and the inner electrons shield the outermost electrons from the nucleus Table (8.5).

Table (8.5) Atomic Radii of First Group Elements

First Group Elements	No. of electrons in the Inner Shells	Atomic Radius (pm)
Li	2	152
Na	10	186
K	18	227
Rb	36	248
Cs	54	265



Interesting Information!

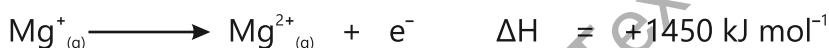
Although you might expect atoms to become larger with the increase in their atomic numbers, this does not always occur because the size of an atom is determined by the diameter of its electron shells.

Ionization Energy

Ionization energy is the amount of energy required to remove the most loosely bound electron from the valence shell of an isolated gaseous atom. When there is only one electron present in the valence shell, the energy required to remove it will be called first ionization energy. It is expressed in kJ mol^{-1} . For example, the first ionization energy of sodium atom is $+496 \text{ kJ mol}^{-1}$.



When there are more than one electrons in the valence shell, they can be removed one by one providing more and more energy. For example, Mg has two electrons in its outermost shell. It is easier to remove the first electron from magnesium than the second one.



Ionization energy value is related to the atomic size. The smaller the radius of an atom, the stronger the attraction between the nucleus and the outer electrons and higher the value for ionization energy. The ionization energy values thus increase from left to right in a period and decrease from top to bottom in a group. Table (8.6) and Table (8.7).

Table (8.6) Ionization Energies of Elements of Second Period

2nd period elements	Li	Be	B	C	N	O	F	Ne
Ionization energy (kJ mol^{-1})	520	899	801	1086	1402	1314	1681	2081

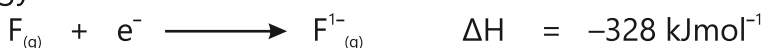
Table (8.7) Ionization Energies of Elements of First Group

First Group Elements	Ionization Energy (kJ mol^{-1})
Li	520
Na	496
K	419
Rb	403
Cs	377

Electron Affinity

Electron affinity is the amount of energy change when an electron is added up in the outermost shell of an isolated gaseous atom. For example, 328

kJ mol^{-1} energy is released when an electron enters in the fluorine atom.



Electron affinity values are also related to the sizes of the atoms. The smaller the size of an atom, the higher the force of attraction with which the nucleus will attract the entering electron and hence higher is the value of electron affinity. Table (8.8)

Table (8.8) Electron Affinities of Group 17 Elements

Group 17 Elements	Electron Affinity (kJ mol^{-1})
F	-328
Cl	-349
Br	-324.6
I	-295.2

In a group, the electron affinity values decrease from top to bottom because the sizes of atoms increase down the group. In a bigger atom, the nucleus will attract the incoming electron with a weaker force and hence the electron affinity will also be low.

Things to Know

While first electron affinities can be negative or positive, second electron affinities are always positive.

Electronegativity

Electronegativity of an atom is its electron-attracting ability. It is defined as the force with which an atom attracts the shared pair of electrons towards itself in a bond. Electronegative atoms are those whose outer electrons are attracted by a large nuclear charge. It increases from left to right in a period and decreases from top to bottom in a group. Thus the most electronegative atoms are found at the top right-hand corner of the periodic table. The most electronegative atoms are F, O, N and Cl. Table (8.9) and Table (8.10). The most electropositive elements lie at the bottom left of the periodic table.

Table (8.9) Electronegativities of Elements of Second Period measured on Pauling scale

2nd period elements	Li	Be	B	C	N	O	F
Electronegativity	1.0	1.6	2.0	2.6	3.0	3.4	4.0

Table (8.10) Electronegativities of Elements of Group 17 measured on Pauling scale

17th Group Elements	Electronegativity
F	4.0
Cl	3.2
Br	3.0
I	2.7



Interesting Information!

Electronegativity is one of the most well-known property for explaining why chemical reactions take place.

8.5 Metallic Character and Reactivity

The metallic character is the tendency of an element to lose electrons and form positive ions or cations.

Since the ionization energy decreases down the group, the elements have increased ability to lose electrons. For this reason both the metallic character and reactivity increase down the group.

As we move from left to right in a period, the nuclear charge increases due to a gradual increase in the number of protons in the nucleus. Owing to this the valence electrons are pulled strongly by the nucleus making it difficult for the atoms to lose electrons. Hence the metallic character decreases in a period from left to right.

The chemical reactivity gradually decreases as we move from left to right in a period. For example, the second period starts with a very reactive element sodium which is followed by the less reactive magnesium. Aluminium and silicon are less reactive than both sodium and magnesium. This is because the number of valence electrons increases, making it difficult to lose electrons. Moving further right in a period towards non-metals, the chemical reactivity gradually increases again.

Things to Know

Metallic character of a metal generally determines its level of reactivity.

Density

It is mass of a substance in a unit volume. Density of elements generally increases from top to bottom in a group but varies less significantly from left to right in a period. It is expressed in g/cm^3 .

Table (8.11) Densities of First Group Elements

First group Elements density	Li	Na	K	Rb	Cs
	0.53	0.97	0.89	1.63	1.879

Exercise

- Barium (Ba) is present in 2nd group and 6th period. Answer the following questions about this element.
 - Is it a metal or a non-metal?
 - Will it be electropositive or electronegative?
 - What is the nature of its oxide?
 - In which physical state you expect this element to exist?
- In which group and period you expect to find an element with the largest atomic radius?
- Can you predict the group number of the most electropositive and the most electronegative elements?
- Choose among the following the element having the lowest ionization energy and the element with highest electron affinity. Also assign its group number and period number Li, K, O, F, Cl.
- Which two elements of the periodic table react to give
 - A basic oxide and
 - An acidic oxide?

Exercise



1. Tick (✓) the correct answer.

- Identify the electronic configuration of the outermost shell of a transition metal.
 - ns^2np^4
 - nd^xns^2
 - ns^2np^6
 - ns^2np^5
- Which is the softest metal?
 - Na
 - Ca
 - Al
 - Zn
- A yellow solid element exists in allotropic forms which is also present in fossil fuels. Indicate the name.
 - Carbon
 - Iodine
 - Aluminium
 - Sulphur
- How many electrons can nitrogen accept in its outermost shell?
 - 2
 - 3
 - 4
 - 5
- Which element is the most reactive element?
 - Oxygen
 - Chlorine
 - Fluorine
 - Nitrogen
- Which element has the highest melting point?
 - Na
 - K
 - Rb
 - Cs

- (vii) Which of the following best describe the correct order of oxygen, fluorine, and nitrogen's atomic radii?

(a) $O < F < N$ (b) $N < F < O$ (c) $F < O < N$ (d) $O < N < F$

2. Questions for Short Answers

- Why was atomic number chosen to arrange the elements in the periodic table?
- What is the significance of the word periodic?
- Why does the size of a period increase as we move down the periodic table?
- In a group, the elements have the same number of electrons in the outermost shell. Why is it so?
- Which element has the maximum atomic radius and which element has the minimum atomic radius in third period?
- The first ionization energy value of magnesium is less than the second one. Give reason.
- Which element has the least value of ionization energy and which element has the highest value of electronegativity?

3. Constructed Response Questions

- Suppose a new element is discovered. Where would you like to accommodate this element in the periodic table?
- What is the first element of the periodic table? Will it lose an electron or gain it?
- Why is it not possible for oxygen atom to accept three electrons to form O^{3-} ion like nitrogen which can accept electrons to form N^{3-} ?

4. Descriptive Questions

- Which information is needed to locate the elements in the periodic table if you do not know its atomic number? Is atomic mass helpful for this purpose?
- How many blocks of elements are present in the periodic table? Are these blocks helpful in studying the properties of elements?
- Explain the variation in the following properties in the periods giving reasons.
(a) Atomic radius (b) Ionization energy
- Which physical properties of elements may lead us to know what type of bond it will form?
- Write down the names of four non-metals which exist in solid state at normal temperature.

Group Properties and Elements

Student Learning Outcomes

After studying this chapter, students will be able to:

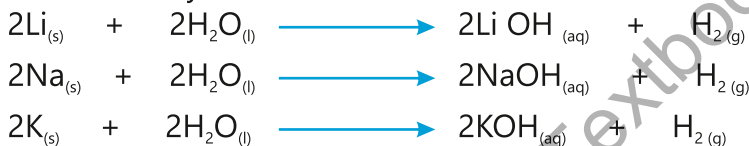
- Define group 1 Alkali metals as relatively soft metals with general trends down the group limited to decreasing melting point, increasing density and increasing reactivity.
- Predict properties of other elements in group 1, given information about the elements.
- Predict properties of elements in group 1 in order of reactivity given relevant information.
- Define group 17 halogens as diatomic non-metals with general trends limited to increasing density, and decreasing reactivity.
- Identify the appearance of halogens at rtp as fluorine as pale-yellow gas, chlorine as yellow -green gas, bromine as red -brown liquid, iodine as grey- black solid
- Explain the displacement reactions of halogens with other halide ions and also as reducing agents
- Predict the properties of elements in group 17, given information about the elements
- Analyse the relative thermal stabilities of the hydrogen halides and explain these in terms of bond strengths
- Define the Group 18 noble gases as unreactive, monatomic gases
- Explain this in terms of electronic configuration properties of metals
- Compare the general physical properties of metals and non-metals (Specifically in terms of:
 - a. thermal conductivity
 - b. electrical conductivity
 - c. malleability and ductility
 - d. melting points and boiling points)

Elements present in a group of the periodic table show similar chemical properties owing to the presence of same number of electrons in their outermost shells. However, a small variation in the chemical properties of elements is expected because the atomic size increases down the group.

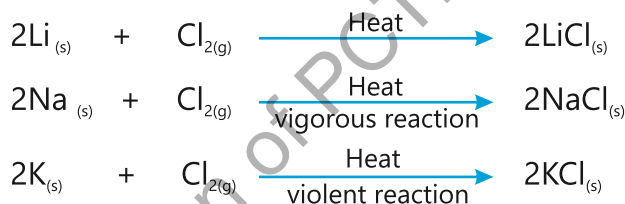
9.1 Properties of Group 1 Elements

All the elements present in group 1 have ns^1 configuration in their outermost shells. They are also called alkali metals. This single electron can be removed easily which makes these metals very reactive except the first element hydrogen which is a gas and a non-metal. When we move from top to bottom in this group, the atomic size increases. Owing to this, it becomes easier for elements to lose electron down the group which is reflected in the increased reactivities of the lower members of the group.

Lithium reacts with water steadily giving hydrogen and lithium hydroxide. Sodium reacts vigorously while potassium reacts violently with water giving their respective water soluble hydroxides.



Similarly, reaction of these metals with chlorine becomes more vigorous as we go down the group.



Exercise

Keeping in view the trends of reactivity in first group elements how would they react with oxygen?



Interesting Information!

Li, Na and K are lighter than water but rubidium sinks in water. Cesium explodes on contact with water, possibly shattering the container.

Increase in the atomic size down the first group also weakens the interatomic attraction of the atomic metals. This fact makes them softer down the group and their melting points decrease.

As we go down the first group, both the size and volume of the atoms increase as the number of electrons and protons increases. But the increase in mass of the elements is greater than the increase in volume; so the density which is defined as the mass per unit volume increases gradually down the group. Table (9.1)

Table (9.1) Physical Properties of First Group Metals

Metal	Li	Na	K	Rb	Cs
Melting point °C	180	98	64	39	28
Density g/cm ³	0.53	0.97	0.86	1.53	1.87

9.2 Properties of Group 17 Elements

All the elements in the group 17 have seven electrons in their outermost shells (n^2, s^5). They are electronegative non-metals because they have strong tendency to accept one electron to become an anion. They exist as diatomic molecules and behave as very reactive non-metals. Atomic radii, melting and boiling points of halogens increase when you go down the group. This is because the atoms get larger as they have more electrons. As a result the molecules which are formed from these atoms will attract each other with a stronger force. This is the reason why the melting and boiling points of bromine and iodine are higher than those of chlorine and fluorine.

Halogens react with alkali and alkaline earth metals to give salts. These elements are thus named as **Halogens** which means salt-forming elements. Unlike metals, the reactivity of halogens decreases from top to bottom in the group. This is due to the fact that atomic size increases down the group and tendency to accept electron from other atoms decreases making them less reactive.

Fluorine gas is pale yellow, chlorine gas yellowish green while bromine is fuming red-brown liquid. Iodine exists as shiny grey crystals which easily turn into dark purple vapours when they are warmed up.

Metal halides are formed when halogens react directly with alkali and alkaline earth metals. Metal halides behave usually as ionic compounds.

Oxidation is a process in which an electron is lost. The substance which loses an electron is called a reducing agent.

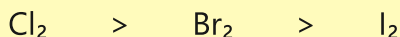


On the contrary, reduction is a process in which an electron is gained and the substance which accepts the electron is called an oxidizing agent.



Oxidation-reduction are simultaneous reactions. In other words an electron is lost only when there is a substance to accept it.

Halogens are oxidising agents and their oxidising power decreases down the group.



This fact gives a unique property to halogens when a halogen having more oxidising power displaces an ion of another halogen from its compound.



Halogens react with hydrogen to give hydrogen halides. Hydrogen halides behave as strong acids in water.



All halides exist in gaseous state at ordinary temperature except hydrogen fluoride which is a liquid. Bond length between hydrogen and halogen increases down the group because as the halogen atom gets bigger the bonding pairs of electron get further away from the halogen nucleus. The bond between hydrogen and halogen therefore gets weaker. The weaker the bond, the less heat energy it will need to break it. Hence the thermal stability of hydrogen halides decreases down the group.



Interesting Information!

Water in swimming pool is sterilized with chlorine.

Exercise

How do halogens react with water?

9.3 Properties of Noble Gases

Elements present in group 18 of the modern periodic table are called noble elements. All noble elements are monoatomic gases having very low boiling points, Helium (He), Neon (Ne), Argon (Ar), Krypton (Kr), Xenon (Xe) and Radon (Ra). All these gases have eight electrons (s^2p^6) in their outermost shells except He which has s^2 electronic configuration. Since their outer shells are complete, they show very little chemical reactivity.

9.4 Physical Properties of Metals and Non-metals

Metals and non-metals can be distinguished based on their physical and chemical properties.

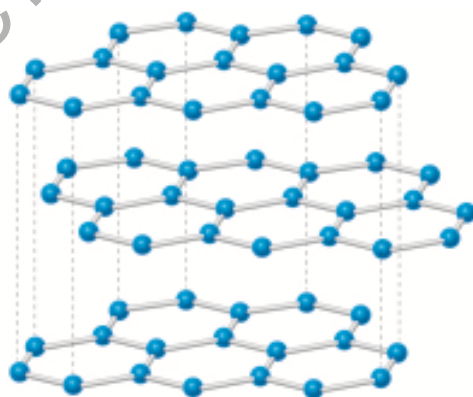
Metals are defined as the elements which can generally form cations easily. They also tend to form metallic bond. Metals can be hammered into thin

sheets. This property is called malleability. Metals can also be drawn into wires and this property is named as ductility. Metallic bonding in metals allows metals to be the best conductor of heat and electricity. Metals are lustrous which means that they have a shiny appearance. Due to high tensile strength metals can hold heavy weights. When metals are hit by an object, they make a ringing sound. Metals cannot be cut easily because they are hard substances.

Due to the presence of strong metallic bond metals generally have high melting and boiling points. Their densities are also very high. Alkali metals being soft metals are treated as exceptions. Example of metals include copper, silver, iron, aluminium, gold, platinum, zinc, etc.

Non-metals generally gain electrons easily. Non-metals show a greater variety of colours and physical states compared to metals. Non-metals cannot be beaten into thin sheets because being brittle they break into pieces when hammered. Sulphur and phosphorous exist in powdered form and cannot be made into sheets. Non-metals cannot be drawn into wires without breaking. Non-metals do not have free electrons due to which the bonds between their atoms are weak and they break down when stretched.

As there are no free electrons so non-metals cannot conduct heat and electricity. Graphite is the only exception. It conducts electricity because of its special crystalline arrangements. The electrons which are present between the layers of graphite crystal are loosely held and hence they can become mobile. The conduction of electricity of graphite is due to the mobility of these electrons.



Non-metals cannot be polished because they either exist in powder or gaseous form. Most of the powders are dull in texture. Due to non-ductile and non-malleable properties, non-metals are not strong at all. Their bonds being weak, break easily. All non-metals have low melting and boiling points. The melting point of sulphur is 115°C . Graphite and diamond have high melting points and these are exceptions. Non-metals have low densities as compared to metals. This means that in non-metals atoms are not strongly bound with each other. Examples of non-metals are oxygen, nitrogen, chlorine, sulphur, carbon and bromine, etc.

**Interesting Information!**

About 75% of all the elements in the periodic table are metals. There are total 20 non-metals which exist in solid or liquid or gaseous state at room temperature.

Comparison of the Physical Properties of Metals and Non-metals

Metals	Non-metals
1. Metals usually have high melting and boiling points.	1. Non-metals may be solids, liquids or gases at room temperature. They show wide range of melting and boiling points.
2. Metals are good conductors of heat and electricity.	2. Non-metals are poor conductors of heat and electricity (except graphite).
3. Metals can be made into different shapes by applying pressure. Metals can therefore be easily drawn into wires and sheets.	3. Non-metals are brittle.
4. Metals are usually lustrous solids (except mercury).	4. Non-metals are dull and cannot be polished (except iodine).
5. Metals are generally tough and strong.	5. Non-metals are neither tough nor strong.

Exercise**1. Tick (✓) the correct answer.**

- (i) Which halogen will have the least reactivity with alkaline earth metals?
 (a) Chlorine (b) Iodine (c) Bromine (d) Fluorine
- (ii) Which compound do you expect to be coloured?
 (a) KCl (b) BaCl₂ (c) AlCl₃ (d) NiCl₂
- (iii) In which element there exists the strongest forces of attraction between atoms?
 (a) Mg (b) Ca (c) Sr (d) Ba
- (iv) Which halogen acid is unstable at room temperature?
 (a) HBr (b) HI (c) HCl (d) HF

(v) Which oxide is the most basic oxide?

(a) Na_2O

(b) Li_2O

(c) MgO

(d) CO

(vi) The following solutions of a halogen and a sodium halide are mixed together. Which solution will turn dark because of a reaction?

(a) Br_2 and NaCl

(b) Br_2 and NaF

(c) Cl_2 and NaF

(d) Cl_2 and NaI

(vii) X is a monoatomic gas, which statement about this is correct?

(a) X burns in air

(b) X is coloured

(c) X is unreactive

(d) X will displace iodine from it

2. Questions for Short Answers

- Why does it become easier to cut an alkali metal when we move from top to bottom in a group I?
- Predict the reactivity of potassium towards halogens.
- In the following reaction, chlorine acts as an oxidising agent. Which is the reducing agent?



- Why does iodine exist in the solid state at room temperature?

3. Constructed Response Questions

- Which noble gas should have the lowest boiling point and why?
- Why are almost all the metals solids while non-metals generally exist as gases and solids?
- Compare the reactivity of chlorine and bromine as an oxidising agent.

4. Descriptive Questions

- Explain the role of catalytic converter in an automobile.
- Why do the chemical reactivities of alkali metals increase down the group whereas they decrease down the group in case of halogens?
- Why are metals generally tough and strong whereas non-metals are neither tough nor strong?
- Both alkali metals and halogens are very reactive elements with roles opposite to each other. Explain.
- Compare the properties of metals and non-metals.



Environmental Chemistry

Student Learning Outcomes

After studying this chapter, students will be able to:

- State that composition of clean, dry air is approximately 78% nitrogen, N_2 , 21% oxygen, O_2 , and the remainder as a mixture of noble gases and carbon dioxide, CO_2 • State the major sources of air pollutants (Some examples include:
 - a. carbon dioxide from the complete combustion of carbon-containing fuels
 - b. carbon monoxide and particulates from the incomplete combustion of carbon-containing fuels
 - c. methane from the decomposition of vegetation and waste gases from digestion in animals
 - d. oxides of nitrogen from car engines
 - e. sulfur dioxide from the combustion of fossil fuels which contain sulfur compounds
 - f. ground level ozone from reactions of oxides of nitrogen, from car engines, and volatile organic compounds, in presence of light)
- State the adverse effects of air pollutants (Some examples include:
 - a. carbon dioxide: higher levels of carbon dioxide leading to increased global warming, which leads to climate change
 - b. carbon monoxide: toxic gas
 - c. particulates: increased risk of respiratory problems and cancer
 - d. methane: higher levels of methane leading to increased global warming, which leads to climate change
 - e. oxides of nitrogen: acid rain, photochemical smog and respiratory problems
 - f. sulfur dioxide: acid rain and haze)
- Explain how the greenhouse gases carbon dioxide and methane cause global warming, (Some examples include:
 - a. the absorption, reflection and emission of thermal energy
 - b. reducing thermal energy loss to space

- Describe the role of sulfur in the formation of acid rain and impact on the environment. Describe the strategies to reduce the effects of major environmental issues (some examples include:
 - climate change: planting trees, reduction in livestock farming, decreasing use of fossil fuels, increasing use of hydrogen and renewable energy, e.g. wind, solar.
 - Acid rain: Use of catalytic converters in vehicles, reducing emissions of sulfur dioxide by using low sulfur fuels and flue gas desulfurization with calcium oxide).
- Describe the role of NO and NO₂ (subscript) in the formation of acid rain, both directly and through their catalytic role in the oxidation of atmospheric sulfur dioxide.
- Explain how oxide of nitrogen form in car engines and describe their removal by catalytic converters, e.g. $\text{CO} + 2\text{NO} \rightarrow 2\text{CO}_2 + \text{N}_2$
- Define photosynthesis as the reaction between carbon dioxide and water to produce glucose and oxygen in the presence of chlorophyll and using energy from light.
- Analyze how to use tools to reduce personal exposure to harmful pollutants (some examples include the usage of masks, air quality indices and CO detectors).
- Identify high risk situation in life including those where long-term exposure to these pollutants can lead to respiratory issues and reduction in quality and longevity of life.

Introduction

The branch of Chemistry which deals with the study of chemicals and other pollutants in the environment is called **environmental chemistry**. It also covers the adverse effects of these chemicals on living and non-living things.

Environmental chemistry is a part of environmental education. The objective of which is to enlighten the people particularly the students, about the importance of protection and conservation of our environment. The need for this environmental education, both formal and non-formal, is keenly felt at the national level.

Since the start of industrial revolution, human activities have played havoc with the atmosphere of the Earth. The gases which are released due to the increasing use of fossil fuels (natural gas, coal and petroleum) have polluted the atmosphere upto such an extent that it is difficult to breathe air in some areas of metropolitan cities.

The excessive use of fertilizers, insecticides and pesticides, etc. for agriculture purposes has proved to be harmful for animals, birds and human beings. The situation is turning serious for every passing day and there is an urgent need to control the emission of pollutants to the atmosphere.

10.1 Air Pollutants

Any substance (solid, liquid or gas) in the air which has adverse effect on human health and quality of life is called an air pollutant.

The concentration of a pollutant is expressed in parts per million (ppm). A

concentration of one ppm means one part of pollutant per million part of solid, liquid or gas mixture in which the pollutant is formed.

Every Individual should try to.....

Pour liquid waste into sewers not in open drains, river and sea. Stop using environmentally hazardous substances.

Major Air Pollutants

Air is not always clean as it should be. There are seven types of harmful substances which account for more than 90% of air pollution. These are fast growing sources of air pollution created by our day-to-day activities. The detail of these substances is mentioned below.

- | | | |
|-------|--------------------|--|
| (i) | Carbon dioxide | (CO ₂) |
| (ii) | Carbon monoxide | (CO) |
| (iii) | Oxides of nitrogen | (NO, NO ₂) collectively referred to as NO _x |
| (iv) | Oxides of sulphur | (SO ₂ , SO ₃) collectively referred to as SO _x |
| (v) | Hydrocarbons | (Methane, ethane, etc.) |
| (vi) | Particulates | (Dust, pollens, metallic compounds) |
| (vii) | Ozone | (O ₃) |

Sources of Air Pollutants

Millions of tonnes of pollutants are emitted into the atmosphere each year as a result of human activities. The major activity among them is the complete and incomplete combustion of fossil fuels which alone is responsible for most of our pollution problem.

Burning of fossil fuels (oil natural gas, coal) produce carbondioxide, carbon monoxide, NO_x, SO_x, CH₄, ash, smoke and suspended particles.

Many of these pollutants are also released into the air by natural processes e.g., volcanic eruption releases large quantities of CO₂, SO₂ and particulates. Methane is released in the air by the decomposition of vegetation. It is also present in waste gases produced during digestion in animals.

Rapid growth of population, urbanization, industrialization and transportation are the main factors which are responsible for environmental pollution. All these factors are increasing in every city of the world especially in the last half century. These pollutants are affecting the environment very badly.

Another pollutant ozone (O₃) is formed when heat and sunlight cause chemical reaction between oxides of nitrogen (NO_x) and volatile organic compounds (hydrocarbons).

In winter the smoke present in the atmosphere is mixed with fog to form

what is called **smog**. Many cities of Pakistan are completely covered with the blanket of smog. Its suspension is caused by a combination of factors which include industrial pollution, vehicle emission and crop burning. These factors are responsible for the accumulation of nitrogen oxides, sulphur dioxide, particulate matter and volatile organic compounds in air.

The following Table (10.2) shows the major air pollutants and their harmful effects on human beings and on the environment.

Table (10.2) Pollutants and Their Harmful Effects

No.	Pollutants	Harmful Effects
1	Carbon dioxide (CO ₂)	Higher levels of carbon dioxide lead to increased global warming which can cause ice caps to melt and oceans to warm, causing sea levels to rise. Extreme weather changes such as heat waves, heavy rains and wild fires also occur.
2	Carbon monoxide (CO)	It is extremely poisonous gas that can cause suffocation and death. Carbon monoxide is very toxic gas that stops the red blood cells in animal's blood from carrying oxygen that body needs.
3	Oxides of nitrogen (NO _x) NO, NO ₂	NO _x can damage lungs, irritate the eyes and damage vegetation. It can also cause acid rain which affects buildings and statues made of limestone.
4	Oxides of sulphur (SO _x) SO ₂ , SO ₃	SO _x irritates the eyes and causes breathing difficulties and acid rain.
5	Hydrocarbons	They can cause pneumonia, coughing, many other breathing and lung diseases. They also cause global warming.
6	Particulate matter	Irritates the eyes and can also cause severe breathing problems for people with asthma. It also makes clothes dirty. Also, visibility is reduced because it produces haze in the air.

7	Ozone	Breathing ozone can cause a variety of health problems including chest pain, coughing, throat irritation and congestion.
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Apart from the effects of these pollutants, the smog present in the atmosphere can lead to health complications like allergies, asthma and lung infections. It also inhibits the growth of plants by reducing the amount of carbon dioxide absorbed during photosynthesis.

Amazing Facts about the Environment

1. 78% of marine mammals are at risk of choking with plastic.
2. Humans can use only 1% of all available water.
3. The world has over 3.04 trillion trees. 27000 out of them are cut down daily to make toilet paper.

Exercise

How does air pollution affect plants?

10.2 Acid Rain

When rain water has pH between 4.2 and 4.4, it is known as acid rain.

In 1852, Robert Angus Smith was the first to show the relationship between acid rain and atmospheric pollution in Manchester, England. He is sometimes referred to as the "Father of Acid Rain". Burning of fossil fuels releases harmful gases into the atmosphere. These gases (SO_2 , SO_3) are produced due to the presence of sulphur in the fossil fuels. SO_2 is converted to SO_3 in the presence of oxides of nitrogen in the atmosphere. Oxides of nitrogen are produced mostly by the direct combination of atmospheric oxygen and nitrogen in the industrial and domestic combustion processes. They are also produced by the combination of atmospheric nitrogen and oxygen in the presence of lightning. Significant amount of nitrogen oxides is produced by the reactions taking place in automobile engines Fig.(10.1).

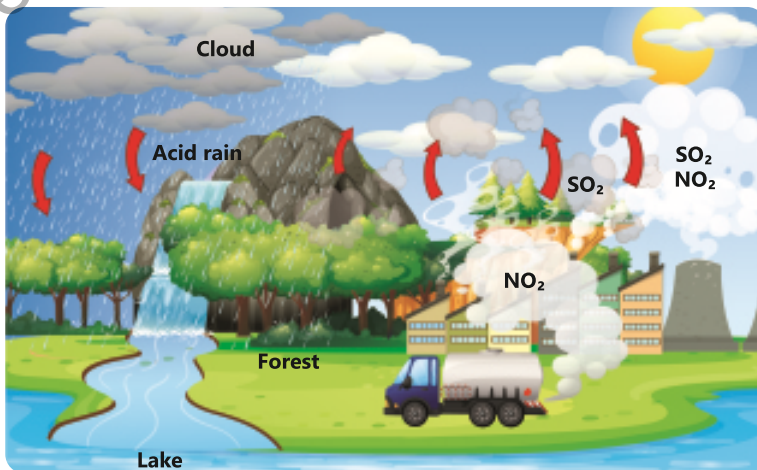
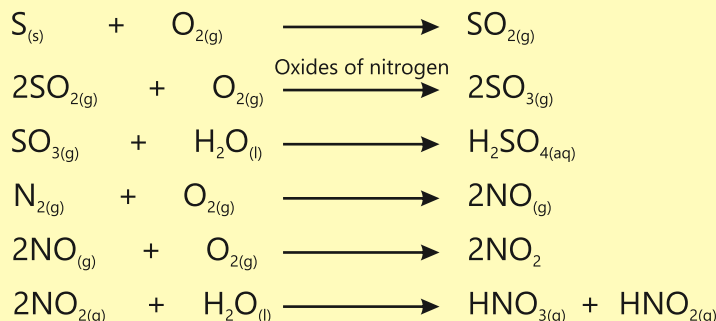


Fig (10.1): Acid rain usually falls far from the site where the acidic oxides are generated.



These gases mix with the moisture that is always present in the air to form acid droplets. Wind can carry these acidic droplets to huge distance.

Finally, these droplets return to the ground as acid rain, acid hail, snow and even fog. Acid rain looks, feels and tastes like clean rain. Its corrosive nature causes widespread damage to the environment.



Interesting Information!

There is a giant floating patch of garbage spread over in the Pacific Ocean. It contains about 100 million tons of garbage.

Effects of Acid Rain

Acid rain causes a number of adverse effects. It tends to increase acidity of the soil, threatens humans and aquatic animals, destroys forests, and reduces agricultural productivity. Besides, it can corrode buildings, monuments, statues, bridges and railings. Most important adverse effects of acid rain are as follows:

(i) Soil

Acid rain makes soil more acidic. It dissolves and washes away nutrients present in the soil which are needed by plants. It can also dissolve toxic substances such as aluminium and mercury which are naturally present in the soil.

(ii) Plants

Acid rain can damage vegetation and plants. Many plants cannot live or grow in acidic soil. Tree roots hold the soil together on hills and mountain areas. If the trees are destroyed then the soil is washed away and new plants cannot grow there.

(iii) Aquatic Life

Acid rain falls into drains, streams, lakes, marshes, rivers and damages the aquatic life. Acid rain can make water too acidic for animals to live in. Due to this, many lakes and rivers no longer have fish.

(iv) Human Health

The acidification of surface water does not affect life directly. However, toxic substances leached from the soil can pollute land water supplies and damage human health.

(v) Agriculture

Crops are less affected by the acid rain than forests. Farmers can prevent acid rain damage by monitoring the conditions of the soil and when necessary adding crushed lime (CaO) to neutralize the acid.

(vi) Human-made Structure

Acid rain and the dry deposition of acidic particles damage buildings, statues, automobiles, and other structures made of stone, metal etc. Historical buildings like Parthenon in Athens (Greece) and the Taj Mahal in Agra (India) are deteriorating due to acid rain.

Exercise

1. Which acids are made when SO_2 and NO_2 dissolve in rain?
2. What happens to the soil if trees are destroyed by acid rain?

10.3 Global Warming (Greenhouse Effect)

The progressive warming up of the Earth's surface due to blanketing effect of man-made carbon dioxide, methane, water vapours and other gases in the atmosphere is called Greenhouse Effect.

The sun emits short wave radiation that passes through greenhouse gases to heat the surface of the Earth. At night, the hot Earth's surface emits longwave radiation that is mostly absorbed by greenhouse gases. This process of absorption prevents the radiation to reach space, reducing the speed at which the Earth can cool off (Fig 10.2). This increases

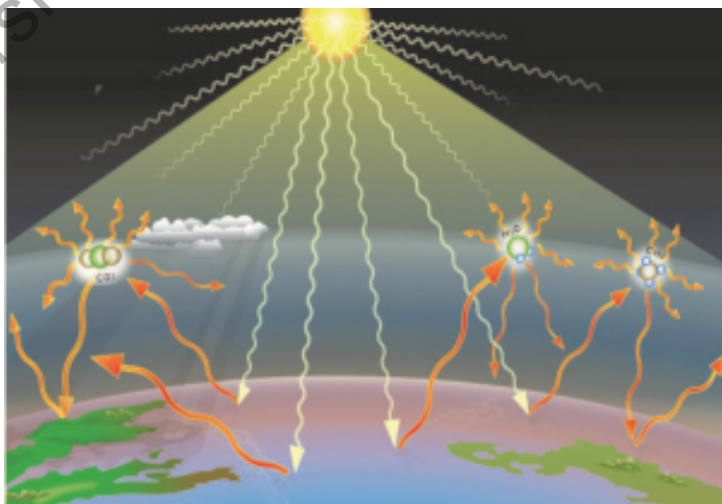


Fig 10.2: Global warming/greenhouse effect

the temperature of the Earth and causes global warming.

Higher the concentration of carbon dioxide and other greenhouse gases, greater will be the absorption of thermal radiation and greater will be the increase in global warming.



Do you Know?

Greenhouse effect, depletion of ozone and acid rain are the global effects of pollution.

Sources of Greenhouse Gases

Due to the burning of the large amount of coal, oil and natural gas, the amount of greenhouse gases, carbon dioxide together with other gases in the atmosphere has increased for the last 300 years. About half of this carbon dioxide is utilized by plant life during photosynthesis. As human beings cut down forests, the capacity of the trees to remove CO_2 from the air is diminished.

Methane is another greenhouse gas which causes adverse effects. The increase in its concentration in air is due to the increased decay of vegetation matter, digestion in animals and increased farming of the rice fields.

Sunlight passes through the greenhouse gases and warms everything on the Earth. The Earth warms up and gives out heat. Some heat passes through the greenhouse gases but some is trapped inside, warming up the Earth.

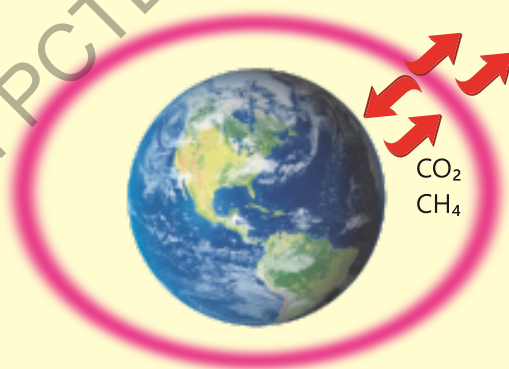


Fig (10.3): Greenhouse gases



Interesting Information!

Water vapours are Earth's most abundant greenhouse gas. It is responsible for about half of Earth's greenhouse effect.

The remaining carbon dioxide which is not utilized in photosynthesis goes on accumulating in the lower areas of the atmosphere and forms a thick dense layer. This layer behaves like glass sheet of greenhouse that allows the incoming solar radiation but does not allow it to escape outside; as a result of this the average temperature of the Earth rises.

A rise of a few degrees in temperature may seem small, but it can be enough to cause significant changes in the climate. At the moment, it is difficult for scientists to say how big the changes will be and where the worse effects will occur. This can damage agriculture and food production as well.

Exercise

1. How do living things add and plants remove carbon dioxide from the air?
2. Which gas do rice plants produce?
3. Which gas is given out by rotting garbage?

The effects of climate change may be physical, ecological, social or economical. Following are four adverse effects of the global warming.

Rise in Sea level

Higher temperature will make the water of the seas and oceans expand. Ice melting in the Antarctic and Greenland will flow into the sea and it results in higher sea levels. This phenomenon will threaten the low-lying coastal areas of the world such as the Netherlands and Bangladesh. Fig (10.4).

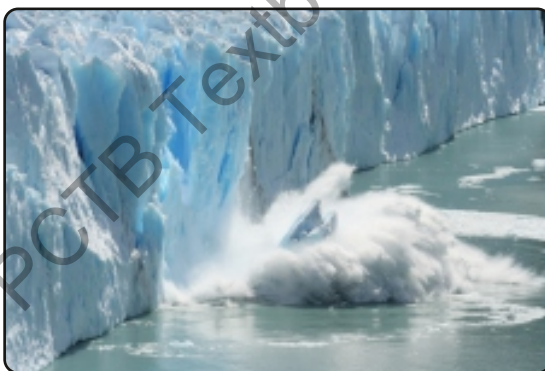


Fig (10.4): Global warming at sea level

Increase in Rainfall

There may be enormous increase in rainfall in a few regions which may increase the sea level. This ultimately will cause worldwide floods endangering survival of living species. Fig (10.5)



Fig (10.5): Effects of Climate Change

Effects on Agriculture

The changes in the weather will affect the types of crops grown in different parts of the world. Some crops such as wheat and rice grow better in higher temperature but other plants such as maize and sugar cane do not Fig (10.6).



Fig (10.6): Global warming risking the agriculture drought.

Summer and Winter Seasons

In moderate region, the winter will be shorter and warmer and the summer will be longer and hotter Fig (10.7).



Interesting Information!

Since 1990, we have lost around 28 trillion tonnes of ice. At presents, its melting rate is 1.2 trillion tonnes per year.



Fig (10.7): Summer and winter

10.4 Strategies to Reduce Environmental Issues

Huge amount of pollutant gases are thrown out in the atmosphere by burning fossil fuels. Automobiles, aeroplanes, industrial machines and coal-fired electricity generating plants, etc. are mainly responsible for the extremely polluted air especially in big cities. Scientists have developed a number of different ways to control this menace of pollution.

Planting trees is thought to be very helpful in removing the air pollution. A well-known process carried out by plants is photosynthesis in which plants clean the air through absorption of carbon dioxide and releasing oxygen. This famous reaction takes place in the presence of sunlight and it is catalysed by chlorophyll, the green pigments present in leaves.

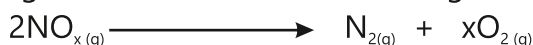


The particulate matter present in the atmosphere is also removed by plants when it deposits on leaves, branches and trunk surfaces.

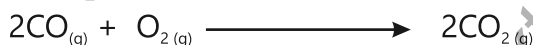
Catalytic converters are used in the exhaust system of modern-day automobiles to reduce the emissions from an internal combustion engine. Due to non-availability of enough oxygen the carbon fuel in engine does not burn completely into carbon dioxide and water. Thus, toxic by-products like CO and hydrocarbons are produced.

A three-way catalytic converter performed the following three functions simultaneously.

1. It reduces nitrogen oxides into elemental nitrogen and oxygen.



2. It oxidizes CO to CO₂



3. It oxidizes hydrocarbons into CO₂ and H₂O



Similarly, the emission of sulphur dioxide can be decreased either by using fuels which have significantly less sulphur contents or using flue gas desulphurization process. This process can remove sulphur dioxide gas from the exhaust gases of fossil fuel. Flue gas is the mixture of waste gases produced when fossil fuels are burnt in power plants. The desulphurization process involves the addition of adsorbents like calcium oxide which can remove upto 95% of the sulphur dioxide from the flue gas.

To discourage the excessive use of fossil fuels in our daily lives it is urgently required to use the renewable resources to meet our energy needs. Renewable resources are those resources that can continue to exist despite being consumed over a period of time even as they are used.

These resources include Sun, wind, water, geothermal and biomass. Solar energy and wind energy have been proved to be very effective ways of generating electricity without damaging the environment.



Interesting Information!

Radon is a natural radioactive gas. When it builds up indoors to high level, it increases the risk of lung cancer.

How to avoid harmful effects of air pollution?

Air quality index (AQI) is a rating system that shows how bad is the atmosphere around you. An (AQI) value under 50 is considered good in quality.

This means that it is safe for you to spend time outdoors without posing a risk to your health. As the (AQI) number increases, so does the risk to health. An (AQI) over 300 is considered hazardous. Children under 18, adults over 65, people with chronic heart or lung disease and diabetic people are high risk groups. Outdoor workers can also be considered at higher risk because of prolonged exposure.

The following steps should be taken when air quality is bad.

1. Reduce the time you spend outdoors. Also reduce the intensity of outdoor activity. According to experts the chances of being affected by unhealthy atmosphere increase if a person stays outdoor for longer periods or involve in more laborious activity outdoor.
2. If you are forced to go out, then consider wearing a mask. Unfortunately, not all the masks provide adequate safety against particulate matter. Cloth or dust mask may not effectively filter out the finer particles. However, well fitted N95 masks have better filtration capabilities and may be safer to use.
3. Keep your indoors healthy by keeping the windows and doors closed. If it is difficult to maintain clean air in the entire room then create a clean room by switching on air conditioner or air cleaner.
4. If you experience such symptoms that worry you, talk to your doctor.
5. Install carbon monoxide detector to detect the increased level of carbon monoxide. These higher levels of CO may occur due to faulty fuel burning appliances.

Breathing in polluted air by these high risk groups may affect their lungs, heart and brain. Air pollutants can enter their blood stream and can cause coughing or itching of eyes which may lead to poor quality of life, hospitalization, cancer or even premature death.

Exercise



1. Tick (✓) the correct answer.

- (i) Which gases are responsible for greenhouse effect?
(a) SO_2 , NO_2 (b) NO_2 , CO (c) CO_2 , CH_4 (d) O_2 , N_2
- (ii) Concentration of which gas in the atmosphere is decreased by photosynthesis in plants?
(a) Oxygen (b) Nitrogen (c) Carbon dioxide (d) Water vapours
- (iii) Which substance remains unaffected in the catalytic converter fixed in the exhaust of the automobiles?
(a) CO_2 (b) CO
(c) NO (d) NO_2
- (iv) People of which age groups are most affected by the air pollution?
(a) Young adults
(b) Middle age people
(c) Children
(d) Both children and aged people
- (v) In which area there is a greater possibility of acid rain?
(a) Around village
(b) Around big cities
(c) Around industrial area
(d) Around water bodies
- (vi) Which catalyst is used in the catalytic converter fixed in the exhaust systems of automobiles?
(a) Ni (b) Cu
(c) Pt , Pd and Rh (d) CaO
- (vii) Which gases contribute towards the formation of acid rain?
(a) Oxides of carbon
(b) Oxides of sulphur
(c) Oxides of nitrogen
(d) Both the oxides of nitrogen and sulfur

2. Questions for Short Answers

- i. What is the main objective of environmental education?
- ii. How is particulate matter released in the atmosphere?
- iii. Which gas is more poisonous, CO₂ or CO?
- iv. How does acid rain affect forests?
- v. In what way sulphur present in fossil fuels becomes dangerous?
- vi. How is wind energy useful for us?

3. Constructed Response Questions

- i. How is the excessive use of insecticides and pesticides harmful for birds?
- ii. Percentage of CO₂ in air is only 0.04%. Then how does it become harmful for the ecosystem?
- iii. Why only some pollutant gases present in the atmosphere cause green house effect while others do not?

4. Descriptive Questions

- i. Describe the harmful effects of the major pollutants present in the air.
- ii. Explain greenhouse effect. How is global warming dangerous for us?
- iii. What is air quality index? What information does it convey?
- iv. Who are high risk groups and why is pollution more dangerous for them?
- v. Describe three strategies to address environmental issues.

Chapter 11

Hydrocarbons

Student Learning Outcomes

After studying this chapter, students will be able to:

- Describe the properties of alkanes as being generally unreactive, except in terms of combustion and substitution by chlorine
- State that in a substitution reaction one atom or group of atoms is replaced by another atom or group of atoms
- Describe the substitution reaction of alkanes with chlorine as a photochemical reaction.
- Describe, using symbol equations, preparation of alkanes from cracking of larger hydrocarbons, hydrogenation of alkenes and alkynes, and reduction of alkyl halides

Introduction

All the organic compounds are known to contain carbon as an essential element. This fact has led us to define organic chemistry as the chemistry of carbon compounds. The ionic compounds like carbonates, cyanides, carbides and cyanates, etc. and the oxides of carbon are, however, classified as inorganic compounds. Apart from carbon, most of the organic compounds contain hydrogen and oxygen as well.

Organic compounds are famous for their large number and diverse behaviour. Several million organic compounds are known to exist naturally or have been synthesized in the laboratory. Organic molecules are usually large and more complex in nature. They include biomolecules like proteins, enzymes, carbohydrates, lipids, vitamins and nucleic acids, pharmaceuticals and synthetic fibres, etc.

The number of compounds formed by the element carbon is far more than the total number of compounds formed by all the rest of elements put together. This is due to some unique properties of carbon.

The element carbon is present at the center of the periodic table and it is generally not possible for it to gain or lose electrons to form ionic bonds. Therefore, it forms four covalent bonds. Because of its small size, these covalent bonds are short and strong enabling carbon to give strong and stable bond with itself and with hydrogen, oxygen and nitrogen. The self-linking property of carbon is called **catenation** and due to this it forms long, straight and branched chains and rings.

Carbon atom mostly forms covalent bonds with other carbon, hydrogen, oxygen, nitrogen and halogens. Compounds in which carbon forms single bond with other atoms are called saturated compounds. These compounds are generally represented by their structural formulas. Methane, a saturated compound, is represented by the following structural formula.



A structural formula thus shows the symbols for the atoms present in the compounds connected by short lines which represent the bonds present in them. Other examples of saturated compounds are C_2H_6 , CH_3Cl , CH_3OH , CH_3NH_2 , etc.

Exercise

What do you understand by the term structural formula of an organic compound?

11.1 Hydrocarbons

The family of hydrocarbons constitutes a very large number of simple organic compounds containing carbon and hydrogen only. Most of the fuels which we use every day, for example, natural gas, LPG (Liquefied Petroleum Gas), CNG (Compressed Natural Gas), petrol, diesel and kerosene oil, are all simple hydrocarbons. These hydrocarbons also serve as a feedstock to prepare more complex and useful compounds like plastics, medicines, synthetic fibres, paints and varnishes.

Hydrocarbons are classified into several structural types called, alkanes, alkenes, alkynes and aromatic hydrocarbons. Only alkanes are discussed in this chapter.

Methane is the simplest alkane and it is mainly used as a fuel but it is also used to make hydrogen gas, carbon black, carbon disulphide, chloroform and hydrocyanic acid, etc.

11.2 Alkanes

Alkanes are the simplest family of hydrocarbons with only carbon – carbon and carbon – hydrogen single bond. Alkanes are also called saturated hydrocarbons because all the four valencies of carbon in them are fully utilized either by hydrogen atoms or by other carbon atoms. For example, in ethane ($\text{CH}_3 - \text{CH}_3$) both the carbon atoms are fully saturated with single bonds.

Alkanes are represented by a general formula $\text{C}_n\text{H}_{2n+2}$ (where n is an integer). Methane being the simplest hydrocarbon, is also called parent hydrocarbon.

As a result of the great complexity and large number of organic compounds, it is not possible to name each and every compound individually. The International Union of Pure & Applied Chemistry has devised a systematic way of naming organic compounds called IUPAC nomenclature.

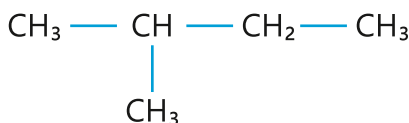
According to IUPAC system of nomenclature, the entire name of an organic compound has three parts:

- 1. Root:** It tells us the number of carbon atoms in the longest continuous chain present in the molecule. The roots upto ten carbon atoms are shown in Table (11.1).
- 2. Suffix:** It is added after the root and tells us about the class of organic compounds.
- 3. Prefix:** It is indicated before the root and tells us about the group or groups attached to the longest chain.

Table (11.1)

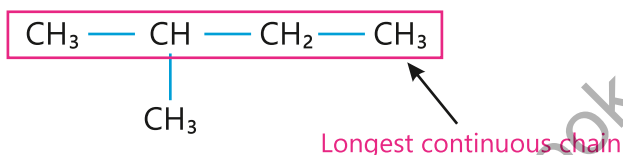
Root	No. of Carbon atoms
Meth-	1
Eth-	2
Prop-	3
But-	4
Pent-	5
Hex-	6
Hept-	7
Oct-	8
Non-	9
Dec-	10

To explain the above system, let us name the following compound.

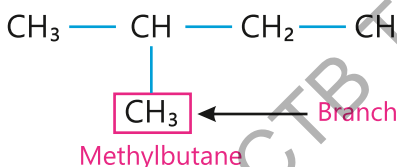


- Identify the longest continuous chain present in the compound.
- Identify the class of organic compounds.
- Identify the substituent or substituents if present.

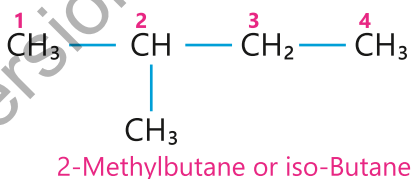
This organic compound contains four carbon atoms in the longest continuous chain and it belongs to the family of organic compounds called Alkanes. The root is therefore But- and the suffix -ane added to this. The organic compound will thus be given the name Butane.



The name of the only branch methyl- will be added to this name as prefix. So the name will become:



To specify where the branch occurs, the longest continuous chain is numbered starting from the end closest to the branch. This number is then attached to this prefix. The name of the above compound will then be:



If butane has no branch it is called normal or n-butane.

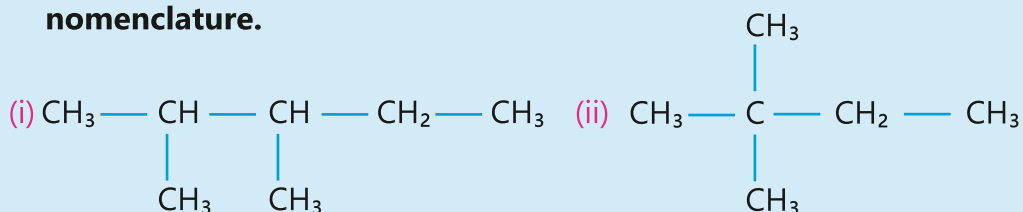


Interesting Information!

The distinguishing feature of alkanes making them distinct from other compounds is their lack of reactivity towards usual chemical reagents.

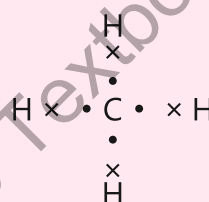
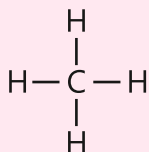
Exercise

Name the following compounds according to IUPAC system of nomenclature.

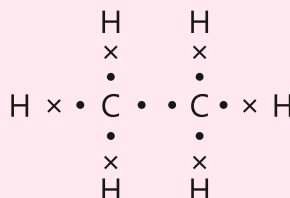
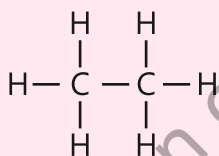


Electron cross and dot structures of Alkanes

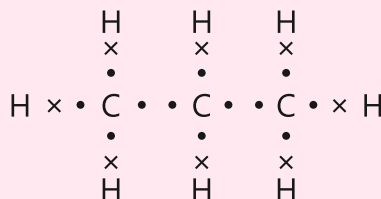
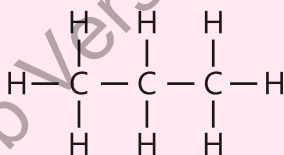
Methane



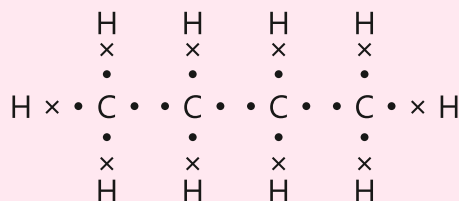
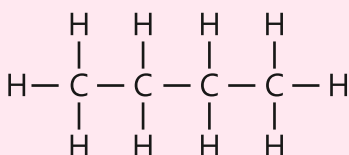
Ethane



Propane



Butane



Exercise

How many methyl and methylene groups are present in each of the above compounds?

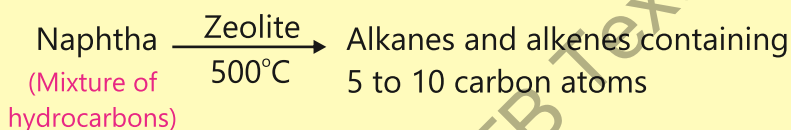
11.3 Preparation of Alkanes

Generally, any member of the alkane series can be prepared by the following methods.

1. Cracking of Higher Hydrocarbons

Cracking is the name of a process in which hydrocarbons with higher molecular masses (which are lesser in demand) are broken up into smaller hydrocarbons which are more in demand. This is done by heating the hydrocarbons at high temperature in the presence of a catalyst.

Fractional distillation of petroleum gives naphtha which consists of a mixture of liquid hydrocarbons. It is then heated at around 500°C in the presence of catalyst called zeolite to give hydrocarbons which have five to ten carbon atoms.

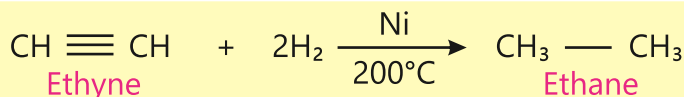
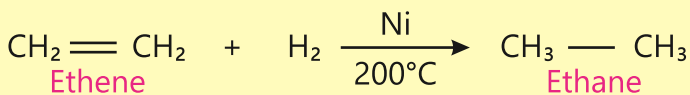


Interesting Information!

Cracking of hydrocarbons helps balance the availability of petroleum fractions with the demand for them. When cracking transforms bigger hydrocarbons into small hydrocarbons, the fuel supply is increased. That helps to balance demand with supply.

2. Reduction of Alkenes and Alkynes

Alkanes can be prepared by reducing alkenes and alkynes with hydrogen gas in the presence of nickel metal as a catalyst. Methane cannot be prepared by this method. The reaction is also called hydrogenation of alkenes and alkynes and is an example of addition reaction. An addition reaction occurs when hydrogen (H_2) is added to an unsaturated compound.

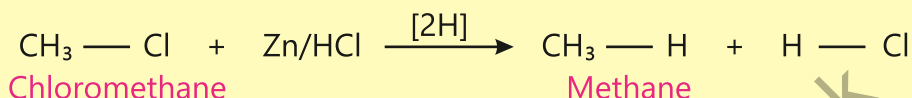
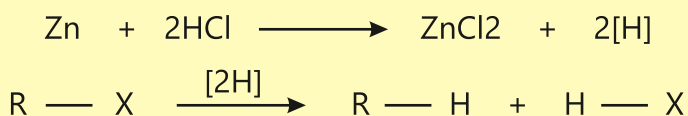


Interesting Information!

Addition of hydrogen to alkenes or alkynes is called reduction reaction. This reaction is used to prepare banaspati ghee and margarine.

3. Reduction of Alkyl Halides

Alkyl halides (R-X) can be reduced to alkanes with nascent hydrogen generated by reaction of zinc metal with hydrochloric acid.

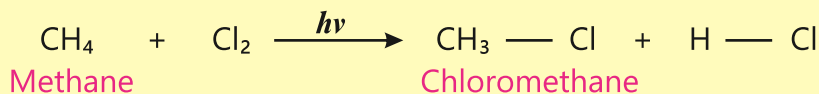


11.4 Important Reactions

Alkanes are sometimes referred to as paraffins which means "little affinity". This term describes their behaviour, for alkanes show little chemical affinity for other substances and are chemically inert to most laboratory reagents. However, alkanes do react with chlorine and oxygen under suitable conditions. The unreactivity of alkanes may be explained on the basis of the non-polarity of the bonds present in them. The electronegativity values of carbon (2.6) and hydrogen (2.1) do not differ appreciably and the bonding electrons between C – H and C – C bonds, are almost equally shared. This fact makes alkanes almost non-polar. In view of this, the reagents like acids, bases, oxidizing agents find no reaction site in alkane molecules to which they could be attached.

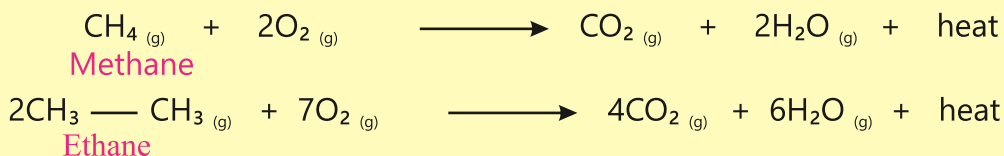
1. Halogenation

Alkanes give substitution reactions. The reactions which involve the replacement of hydrogen of alkanes by an atom or a group of atoms like halogen are called substitution reactions. Alkanes react with halogens especially chlorine to give alkyl halides. Since these substitution reactions are carried out in the presence of UV light, these are called photochemical substitution reactions.



2. Combustion

Alkanes burn in oxygen or air to form CO_2 and H_2O with the evolution of large amount of heat. The reaction is called combustion.



Interesting Information!

A mixture of natural gas (methane) and air may explode when ignited. This is the main cause of explosion at homes where gas leakage occurs.

Exercise

- (1) In the reduction of alkyl halides with Zn / HCl, alkyl halide is being reduced. Which species in this reaction is being oxidized?
- (2) During the combustion reaction of ethane, which bonds are being broken and which are being formed?

Exercise

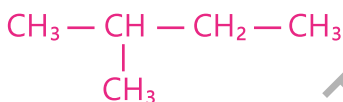


1. Tick (✓) the correct answer.

- (i) Which other atom is almost always present along with carbon atom in all organic compounds?
 - (a) Oxygen
 - (b) Nitrogen
 - (c) Hydrogen
 - (d) Halogen
- (ii) Which other metal can be used to reduce alkyl halides?
 - (a) Al
 - (b) Mg
 - (c) Ni
 - (d) Co
- (iii) If naphtha undergoes a combustion reaction what products do you expect it to form?
 - (a) Alkanes
 - (b) Alkenes
 - (c) CO₂ and H₂O
 - (d) Both alkanes and alkenes
- (iv) Which alkane will evolve the most amount of heat when it is burnt with oxygen?
 - (a) Ethane
 - (b) Propane
 - (c) n-Butane
 - (d) iso-Butane
- (v) Which reaction is not given by alkanes?
 - (a) Substitution
 - (b) Combustion
 - (c) Addition
 - (d) Cracking
- (vi) Which hydrocarbon is responsible for explosions in coal mines?
 - a) Butane
 - b) Pentane
 - c) Methane
 - d) Ethene
- (vii) Which of the following is not a process of halogenation of alkanes?
 - (a) Cracking
 - (b) Chlorination
 - (c) Bromination
 - (d) Iodination

2. Questions for Short Answers

- Differentiate between an organic and an inorganic compound.
- Why are organic compounds found in large numbers?
- Name the products which are obtained when natural gas is oxidized under controlled conditions.
- How naphtha fraction is decomposed to give lower hydrocarbons?
- Write down the molecular formula, structural formula and the condensed structural formula for iso-butane.
- How are organic compounds useful for us?
- Write down the names of five such organic compounds which exist naturally?
- Give IUPAC name to the following compound.



- How do the melting and boiling points of alkanes change when we move from lower members to higher members?

3. Constructed Response Questions

- Why do alkanes show little reactivity towards the other reagents?
- Why does a mixture of natural gas and air explode?
- How do you compare the melting and boiling points of inorganic and organic compounds?

4. Descriptive Questions

- Describe the importance of organic compounds in daily life.
- Why is carbon so important as an element that the whole branch of chemistry is based on it?
- A carbon-carbon single bond (C – C) in alkanes does not behave as a functional group but a carbon carbon double bond (C = C) in alkenes does. Explain.
- Explain IUPAC system of nomenclature for alkanes.
- How combustion reaction of alkanes is useful for us?

Laboratory and Practical Skills

Student Learning Outcomes

After studying this chapter, students will be able to:

- Explain, with examples, the types of chemical hazards in the lab and suggest safety precautions. (Types of chemical hazards to be identified: flammable or explosive hazards, corrosive hazards, toxic hazards)
- Recognize the meaning of different chemical hazard signs in the lab and on chemicals.
- Recognize the importance of personal protective equipment (PPE) by correctly identifying the types of PPE needed for different lab. activities.
- Locate the nearest fire extinguisher and emergency shower.
- Show awareness of emergency procedures in the event of an emergency in the lab.

Introduction

A chemistry laboratory is a place where a student is trained to observe the physical and chemical characteristics of substances by following definite procedures. Before starting the laboratory work, a student should get himself familiarized with the layout of the laboratory and various fittings provided in the laboratory table as well as the side shelves. Students are expected to conduct themselves in a responsible manner at all times in the lab. They are advised not to work alone in the lab. Experiments should be performed in the presence of lab. instructor and other laboratory staff. All equipments should be checked before use whether they are working properly according to the requirements of the experiments. Determine the potential hazards related to any equipment or the experiment before beginning any work. Appropriate safety precautions must be observed at all cost. There must not be any crowding in the lab and students should stick to their work places at a safe distance from each other. Don't bring any food items in the lab. Never taste or smell any compound or a gas. If it is necessary to smell a gas it is always advised to waft the fumes or vapours towards your nose.

Warning signs are displayed when unusual hazards, hazardous materials, hazardous equipment or special conditions are expected. Do not pour chemicals down the drains and do not utilize the sewer for chemical waste disposal. Keep all sink traps and floor drains clean. Laboratory chemical waste can be disposed off in sewer or trash bin if they are non-hazardous materials. Acids and bases are first neutralized followed by sewer disposal. Hazardous waste material is transported to hazardous waste disposal site.

Misuse and mishandling of chemicals may create serious problems for the laboratory workers. A laboratory worker must use the chemicals according to the standard procedures keeping in view the particular hazards and precautions required for the safe use. The chemicals which can create problems for the safety of workers are cleaning agents, disinfectants, solvents, paints, compressed gas cylinders, mineral acids, carcinogenic chemicals etc.

Recognizing hazards which are commonly encountered in the laboratory helps to identify and minimize many of the health and safety problems. Most hazards which we might face while working in the laboratory fall into the following categories.

12.1 Chemical Hazards in the Laboratory

12.1.1 Flammable and Explosive Chemical Hazards

To start working in the laboratory requires great care, responsible behaviour and good attention. It is important to exercise extreme caution while working with delicate instruments, hazardous chemicals and open flames. If flammable and explosive chemicals are not handled in a safe and compliant manner, they can cause acute health problems. These problems may include burns, eye injuries, lung disease and suffocation.

Chemicals that cause a sudden release of pressure, gas and heat when they experience sudden shock are called explosive chemicals. Examples of chemicals which are expected to explode are picric acid, 2,4 -di-nitrophenyl hydrazine, benzoyl peroxide, nitrocellulose etc.

Flammable chemicals or mixtures are those which have a flashpoint around room temperature. Examples of flammable compounds are ethers, methylated spirit, benzene, acetone, petrol etc.

If you ever come across any chemical that you suspect to explode, do not attempt to move the container to avoid shock. Explosives can cause damage to people, windows, tables etc.

Avoid using a chemical that is hazardous. Look for its alternative. If you must use a potentially dangerous chemical, you must follow the underlying safety instructions.

1. Obtain prior approval from your teacher.
2. Always use smallest quantity of the chemicals.
3. Always conduct experiment in fume hood.
4. Remove all other chemicals and apparatus around you.
5. Inform other people working with you.
6. Always wear safety spectacles, gloves and lab. coat.
7. Always keep flammable compounds away from heat source.
8. Pour the flammable liquid very carefully.
9. Properly dispose off any hazardous waste.
10. Do not store flammable liquid in refrigerator.

12.1.2 Corrosive Hazards

Corrosive chemicals attack living tissues when they come in contact with them. They can be in the form of solids, liquids or gases. Such chemicals attack skin, eyes and respiratory tract and in the intestine as well. Whenever you work with corrosive chemicals, wear splash goggles instead of safety glasses and use a face shield.

Safety Precautions

1. Corrosive chemicals must be used in a fume cupboard to avoid breathing corrosive vapours.
2. While mixing concentrated acids with water, always add acid slowly to water and not vice versa.
3. Ensure eye wash and emergency shower is available.
4. Wash the affected area with soap and water and seek medical attention in case of emergency.

Examples of corrosive chemicals are mineral acids including HF, caustic alkalies, acetic acid (glacial) etc.

12.1.3. Toxic Chemical Hazards

A toxic chemical is a poisonous material which is capable of causing serious health problems. Mercury, benzene, chlorine, pesticides, ammonia, hydrogen cyanide are some examples of toxic chemicals. The following safety instructions may be ensured in case you intend to work with toxic chemicals.

1. Wear gloves, masks or other protective devices.
2. Keep the work area well ventilated.
3. Keep the toxic chemicals in original container.
4. Do not work alone.
5. Wash your hands with soap and water after you finished.
6. Always work in fume hood because toxic vapours can be formed during an experiment.
7. Seek immediate medical aid if you think you may have exposed to poisonous substance.

12.2 Hazard Signs

A chemistry laboratory is a strict area where rigorous rules must be practised to avoid a chance of a deadly accident. A dangerous situation may arise not only for the individuals working there but for the whole area. In a laboratory there are several hazardous materials, sensitive equipments and specified areas for specific tasks. Proper warning signs ought to be posted on these areas to ensure that every person entering there must understand and act accordingly to maintain laboratory safety.



Fig (12.1): Different Hazard Signs

Several signs and symbols are posted in different areas of the lab. and bottles containing hazardous chemicals. These signs indicate that specific precautions must be observed according to the requirement of the sign posted there. If you see such signs, you must be alert and take extra care to maintain safety in that area, Fig (12.1).

Exercise

1. What does warning sign “caution” convey the message?
2. Name some explosive chemicals.

12.3 Personal Protective Equipment (PPE) in the Laboratory

Personal protective equipment should be made available to students to face any emergency situation which may arise in the lab. They are also useful to reduce exposures to hazardous chemicals. Proper protective equipment include such items as lab coat, protective glasses, face shields, apron, boots and hearing protection.

12.4 Location of Fire Extinguisher

Chemical laboratories using such materials which are likely to catch fire during experiments must have a portable fire extinguisher. This equipment can quickly be used to control a small fire if it is applied by a student individually. For this purpose all students should be well aware of the location where this fire extinguisher is placed. A training session should be held to train all the students to know how to handle and apply this fire extinguisher to extinguish the fire properly without any panic or harm to anybody.

Similarly the facility of a shower should also be made available in the lab whose location and working must be told to everybody working in the lab. In case of fire or any other emergency students should know how to face that emergency situation.

Exercise

1. Should emergency drills be compulsory or optional?

12.5 Emergency Situation in the Lab.

Students should make themselves aware of the actions that need to be taken in case of an emergency in a laboratory or if a person is affected. For this purpose periodic drills should be held with compulsory participation. Students should not only been given lectures but involve them practically to handle the emergency situations. During drill firefighting and other equipments must be checked whether they are in proper working order or not.

The following points should be kept in mind to cope with the emergency situation.

1. Stay calm and do not panic.
2. Alert people in the area to evacuate.
3. In case of fire, close doors to confine fire. Use fire extinguisher to put down the fire.
4. In case of chemical emergency adopt safety procedures as mentioned in article 13.1.
5. Call and assist emergency staff.

Exercise



1. Tick (✓) the correct answer.

(i) Safety in the chemistry laboratory is:

- (a) the responsibility of the students only
- (b) the responsibility of the professor only
- (c) the responsibility of the lab. incharge only
- (d) a shared responsibility

(ii) Accidents often result from:

- (a) making mistakes
- (b) failure to use common sense
- (c) failure to follow instructions
- (d) all of the above

(iii) The label "Corrosive" on a chemical bottle indicates:

- (a) that the material is an oxidizing agent
- (b) that the material can degrade rapidly upon exposure
- (c) that its contact destroys living tissue
- (d) that the chemical can explode

- (iv) When diluting an acid with water:
- (a) do it quickly
 - (b) do not stir the container
 - (c) always add acid to water
 - (d) always add water to acid
- (v) What should you do in case of a fire drill in the lab.?
- (a) run to safety shower
 - (b) climb into the fume cupboard
 - (c) close gas valves and turn off all equipments
 - (d) carry chemicals out of the lab.

2. Questions for Short Answers

- i. Name some corrosive chemicals.
- ii. What type of safety precautions are adopted to avoid damage due to explosive chemicals?
- iii. Why signs and symbols are posted on lab. and chemical bottles?
- iv. How fire caused by chemicals should be handled?
- v. Why emergency drills are important to face emergency situations?

3. Constructed Response Questions

- i. How will you handle an emergency situation caused by fire due to short circuiting?
- ii. What type of reactions should be carried out in fume cupboard?
- iii. Why sudden shock can cause some chemicals to explode?

4. Descriptive Questions

- i. Explain hazards due to explosive and toxic chemicals.
- ii. Write down five such common safety instructions which are used to avoid all types of hazards.
- iii. Explain the importance of warning signs and symbols to avoid any accident in the lab.
- iv. Name some toxic chemicals. Describe the effects of spreading toxic gas in the lab.
- v. A student has spilled over a corrosive and explosive chemical due to an accident. Which emergency measures you will take to tackle the situation.

Glossary

2n² Formula: Formula used for filling the electrons in shells.

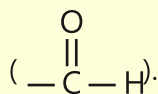
Accuracy: It refers to how close a measurement is to the true value.

Acid Rain: Pollutant gases mixed with rain water come down as acid rain.

Activation Energy: Energy absorbed by the reactants and product molecules in order to be converted into the transition state.

Aerobic Respiration: The process of respiration in the presence of oxygen.

Aldehydes: organic compounds containing the aldehydic functional group



Aliphatic Compound: (Acyclic compounds) Compounds containing open chain of carbon atoms.

Alkanes or Saturated Hydrocarbons: Hydrocarbons in which all the four valencies are fully utilized or they contain single bonds only.

Alkenes: Compounds containing a double bond between two carbons atoms.

Alkyl Amine: Organic compounds containing amino group ($-\text{NH}_2$) as a functional group.

Alkyl Halides: A family of organic compounds containing halogen atom as a functional group.

Alkyl Radical: When an alkyl molecule drops one of its many hydrogen atoms.

Alkynes: Compounds containing a triple bond between two carbons atoms.

Aromatic Compounds: Compounds containing at least one benzene ring.

Arrhenius Acid: A chemical which gives proton (H^+) in water.

Arrhenius Base: A chemical compound which gives hydroxide ion (OH^-) in water.

Astrochemistry: It is the study of molecules and ions occurring in stars and interstellar space.

Atomic Mass: The mass of an atom of an element relative to the unit mass, which is 1/12th of the mass of C-12.

Atomic Number: The number of positively charged particles (protons) present in the nucleus of an atom.

Atomic Size: Average distance between the nucleus of an atom and its outermost electronic shell. Its units are m or pm.

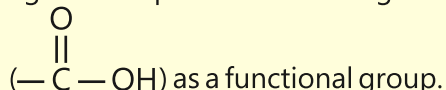
Avogadro's Number: The huge number 6.022×10^{23} is called Avogadro's number.

Biochemistry: Which deals with the study of chemical compounds present in the living things.

Bronsted Acid: A compound which can donate proton.

Bronsted Base: A compound which can accept proton.

Carboxylic Acids: Organic compounds containing carboxyl group



Catalytic Converter: It is a device used in the exhaust of an automobile which converts harmful gases produced in the engine.

Catenation: The self-linking property of carbon.

Chemistry: It deals with the composition and changes in matter and the laws which govern these changes.

Colloidal Solution: A solution in which solute particles are bigger than those present in a true solution and which cannot be filtered.

Combustion: Burning of an organic compound in an excess of oxygen.

Concentrated Solution: A solution that contains a relatively large amount of a dissolved solute.

Concentration of a Solution: The amount of a solute which has been dissolved in a particular amount of solvent.

Coordinate Covalent Bond: When the shared pair of electrons is provided by one of the bonded atoms, a coordinate covalent bond is formed.

Covalent Bond: It is the force of attraction that arises between two atoms due to mutual sharing of an electron pair.

Crystal Lattice: Three-dimensional arrangement of ions.

Cyclic or Ring Compounds: Compounds in which carbon atoms are linked together to give a ring.

Dilute solution: A solution that contains a relatively small amount of dissolved solute.

Discharge Tube: A glass tube containing a gas at a very low pressure and provided with electrodes to study the passage of electricity through the gas.

Dynamic Chemical Equilibrium: When the rate of forward reaction becomes equal to the rate of reverse reaction, at this stage the reaction is said to be in a state of dynamic chemical equilibrium.

Electron: It is the smallest negatively charged particle present in all kinds of atoms. Its mass is 9.1095×10^{-31} kg and carries a charge -1.602×10^{-19} C.

Electronegativity: It is the power of an atom to attract the shared pair of electrons.

Empirical Formula: The formula of a compound, which shows the minimum ratio present between its atoms.

Endothermic Reactions: Those chemical reactions during which heat is absorbed.

Enthalpy of Reaction: Heat of reaction which takes place at constant pressure.

Enthalpy: It is the measurement of energy in a thermodynamic system.

Environmental Chemistry: In this branch, we study the chemicals and other pollutants present in the environment. It also covers the effects of these pollutants on living and non-living things.

Exothermic Reactions: Those chemical reactions during which heat is evolved.

Extranuclear Portion: Area surrounding the nucleus of an atom.

First Ionization Energy: The minimum amount of energy required to remove an electron from the outermost electronic shell of an isolated gaseous atom. Its unit are kJ mol^{-1} .

Formula Mass: Formula mass is the mass of a compound relative to the unit mass which is 1/12th of the mass of C-12.

Functional Group: An atom or group of atoms or a bond whose presence imparts characteristic properties to the organic compounds.

Geochemistry: It covers the study of chemical composition of rocks and minerals.

Global Warming: The progressive warming of the Earth's surface due to the blanketing effect of CO_2 , CH_4 and the water vapour present in the atmosphere.

Heat content: The total amount of heat energy present in a molecule under standard conditions.

Heat of neutralization: The heat given out during a neutralization reaction when one mole water is formed from an acid and base is called the heat of neutralization.

Heat of Reaction: Heat evolved or absorbed during a chemical reaction which takes place at any pressure.

Hydrated Salt: A salt with water molecular of crystallization.

Hydrocarbons: Compound of carbon and hydrogen only.

Inorganic Chemistry: The study of all other elements and their compounds except carbon and its compounds.

Inorganic Compounds: Compounds obtained from non-living things or mineral sources or synthesized in the laboratory by reacting metals with non-metals.

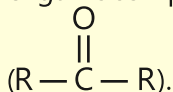
Intermolecular Forces: Forces between two separate molecules.

Ionic Bond: A bond formed due to the electrostatic force of attraction between oppositely charged ions.

Irreversible Reaction: Reaction which moves in one direction only; the reactants react to give the product.

Isomerism: The phenomenon shown by the organic compounds having the same molecular formula but different structural formula.

Ketones: Organic compounds containing the ketonic functional group



Mass number: The total number of protons and neutrons present in the nucleus of an atom.

Medicinal Chemistry: Designing and synthesizing medicines or drugs which are useful to mankind.

Metallic Bond: When positively charged metal ions are held together by freely moving electrons, the bond formed is called a metallic bond.

Metallic Lustre: Shine present on metallic surfaces.

Modern Periodic Table: A table in which elements are arranged in ascending order of their atomic numbers.

Molar Mass: The mass of an element or a compound which contains Avogadro's number of particles.

Molecular Formula: The formula of an element or a compound which tells the actual number of atoms present in the molecule of that element or a compound.

Molecular Mass: Molecular mass is the mass of the molecule of an element or a compound relative to the unit mass, which is 1/12th of the mass of C-12.

Neutralization: Acids and bases react together to form salts and water and, in this way, they neutralize the properties of each other. This reaction is called neutralization reaction.

Neutron: It is the smallest neutral particle present in the nucleus of atoms. Its mass is slightly more than that of a proton.

Nuclear Chemistry: This branch deals with the reactions taking place in the nucleus of an atom.

Nucleus: Central part of an atom where most of its mass is concentrated. Its size is very small as compared to the size of the atom.

Octet Rule: When an atom has eight electrons in its outer most shell, it is said to be stable and does not combine with other atoms to reduce its energy. This is called octet rule.

Orbit: the circular path of an electron around the nucleus.

Organic Chemistry: The branch of chemistry in which we study the compounds of carbon.

Organic Compounds: Compounds obtained from living or plant and animal sources and which can be synthesized in the laboratory. All the organic compounds contain carbon as an essential element.

Oxidation: A process in which an electron or electrons are lost.

Oxidising Agent: A substance which accepts an electron or electrons.

Physical Chemistry: This branch investigates how substances behave on an atomic or molecular level and how physical laws govern the specific characteristics of atoms and molecules.

Pollutant: Any solid, liquid or gaseous substance present in such concentration as may be injurious to health.

Polymer Chemistry: It focuses on the properties, structure and synthesis of polymers and macromolecules.

Proton: It is the smallest positively charged particle present in all kind of atoms. The mass of this particle is equal to the mass of hydrogen nucleus (H^+).

Reducing Agent: A substance which loses an electron or electrons.

Reduction: A process in which an electron or electrons are gained.

Reversible Reaction: A chemical reaction, which takes place in both directions, forward as well as backward.

Saturated Solution: A solution, which contains the maximum amount of a solute at a particular temperature and which is unable to dissolve further amount of solute in it.

Solubility: The amount of a solute in grams which has been dissolved in 100g of a solvent at a particular temperature to prepare a saturated solution.

Strong Acid: An acid which ionizes completely in water.

Strong Base: A base which can ionize completely in water giving excess of hydroxide ions.

Structural Formula: The formula which shows the arrangement of atoms in a compound.

Substitution Reaction: A reaction which occurs when an atom or a group of atoms from the reagent displaces an atom or group of atoms from the organic reactant.

Surrounding: Everything else which does not fall in the system.

System: Anything under consideration for the purpose of study.

Transition Elements: Elements having incomplete penultimate (next inner to the outermost) electronic shell.

Transition State: A state of molecules when they are undergoing breakage or formation of bonds.

Unified Atomic Mass Unit: Unit of a new scale, which is equal to $1/12$ th of the mass of $C-12$.

Unsaturated Hydrocarbons: Hydrocarbons containing double or triple bonds.

Unsaturated Solution: A solution, which can dissolve further amount of a solute at a particular temperature.

Vital Force: The imaginary supernatural force which was supposed to be present in all those compounds which were obtained from living things.

Water of Crystallization: The number of water molecules present in the crystals of a solid.

Weak Acid: An acid which ionizes partially in water.

Weak Base: A base which ionizes partially in water.

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