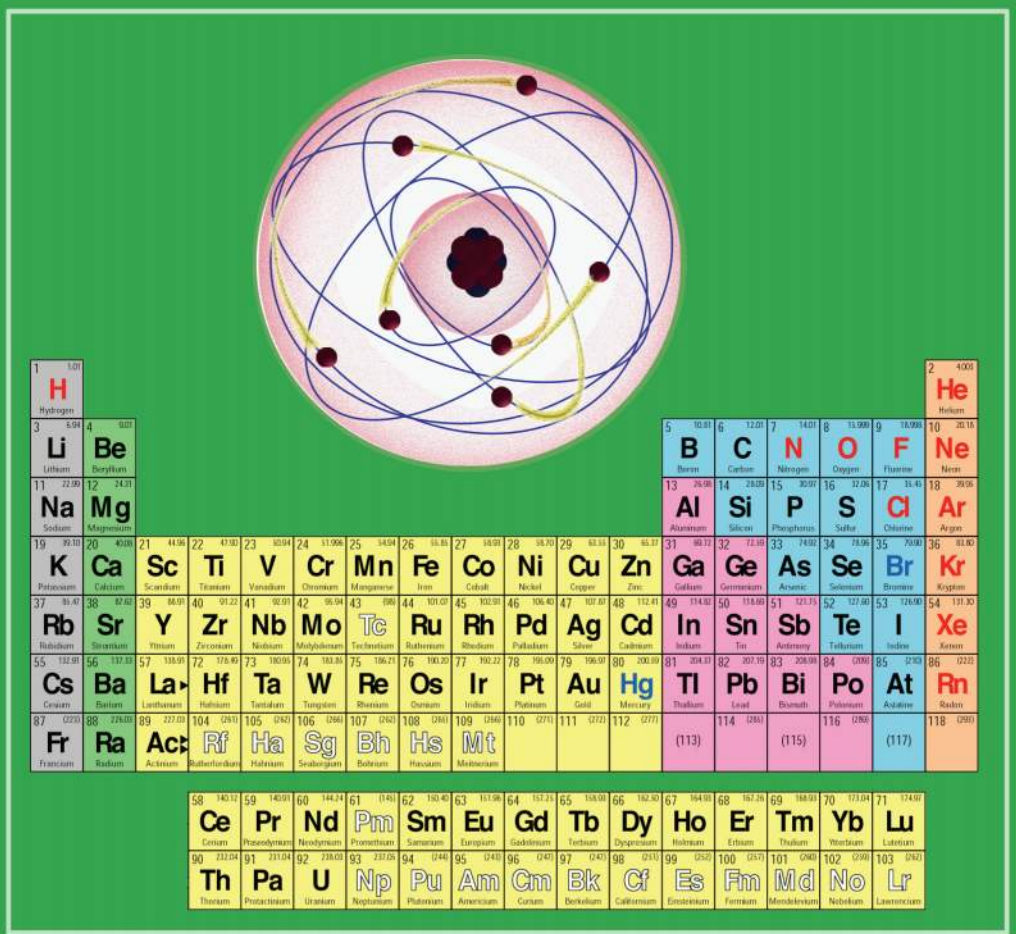
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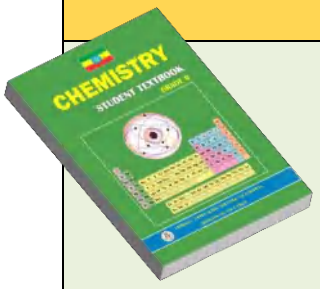
GRADE 9



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CHEMISTRY

STUDENT TEXTBOOK

GRADE 9

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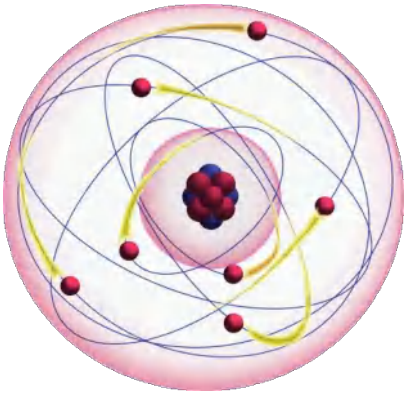
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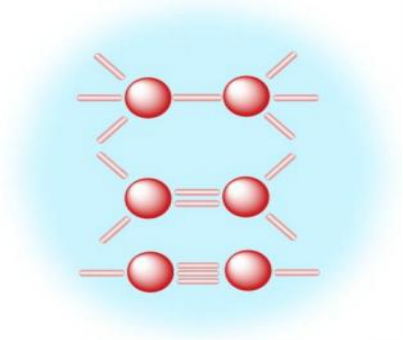
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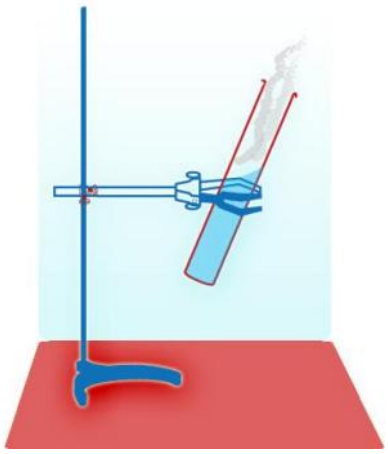


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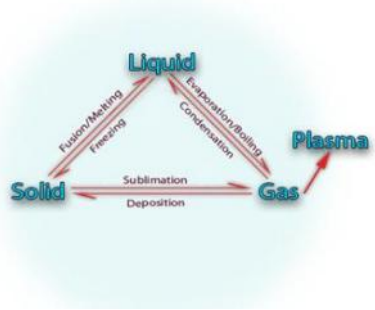
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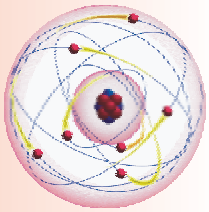
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UNIT 1



Bohr's Model

Structure of the Atom

Unit Outcomes

After completing this unit, you will be able to:

- ***E*** *comprehend Dalton's atomic theory and modern atomic theory;*
- ***E*** *understand the discovery of the electron and the nucleus;*
- ***E*** *know the terms like atomic number, mass number, atomic mass, isotope, energy level, valence electrons, and electron configuration;*
- ***E*** *understand the Dalton, Thomson, Rutherford, Bohr and the quantum mechanical atomic models;*
- ***E*** *develop skills in:*
  - *determining the number of protons, electrons, and neutrons of atoms from atomic numbers and mass numbers,*
  - *calculating the atomic masses of elements that have isotopes,*
  - *writing the ground-state electron configurations of atoms using sub-energy levels and drawing diagrammatic representations of atoms; and*
- ***E*** *demonstrate scientific inquiry skills: observing, comparing and contrasting, communicating, asking questions, and applying concepts.*

MAIN CONTENTS

1.1 Atomic theory
1.2 Discoveries of the fundamental subatomic particles and the atomic nucleus
1.3 Composition of an atom and the isotopes
1.4 The atomic models
– Unit Summary
– Review Exercise

Start-up Activity

Scientists have been able to gather information about atoms without actually seeing them. Perform the following activity to get an idea about the structure of the atom.

Take an onion. It looks solid. Peel off a layer, and you will find another layer underneath. Layer after layer surfaces as the onion is peeled off. Keep on peeling to its core.

Analysis

1. How do you compare this with the atomic model?
2. What do the layers represent in the atomic model?
3. What does the core represent in the atom?

Two thousand years ago, ancient Greek philosophers developed a theory of matter that was not based on the experimental evidences. A notable Greek philosopher, namely, Democritus, believed that all matter was composed of very tiny, indivisible particles. He called them atomos. Hence, the word 'atom' came from the Greek word atomos, which means uncuttable or indivisible.

Aristotle was part of the generation that succeeded Democritus. He did not believe in atomos. Aristotle thought that all matter was continuous. That is, if one proceeded on breaking down a substance, it would be impossible to reach to the last indivisible particle. In other words, it would continue to divide infinitely. His opinion was accepted for nearly 200 years.

The early concept of atoms was simply a result of thinking and reasoning on the part of the philosophers, instead of experimental observations. In 1803, however, John Dalton proposed a completely different theory of matter. His theory was based on scientific experimental observations and logical laws. These scientific assumptions were

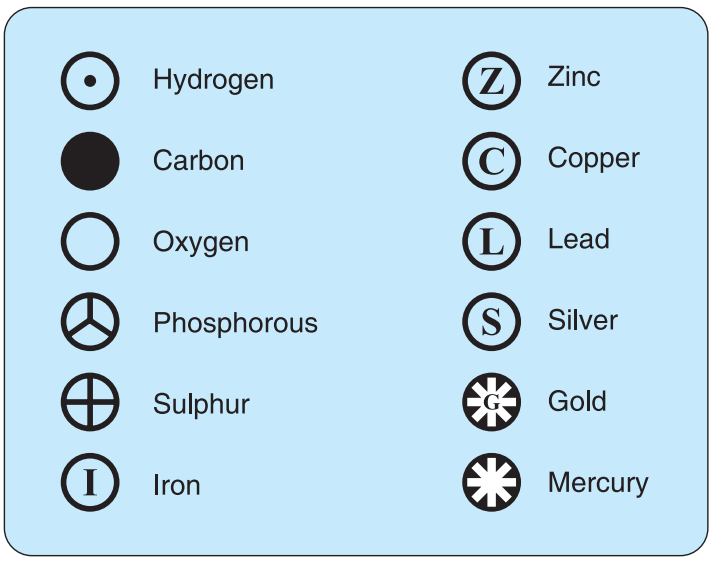


Figure 1.1 Dalton's atomic symbols.

very closely related to what is presently known about the atom. Due to this, John Dalton is often referred to as the father of modern atomic theory.

Dalton also worked on the relative masses of atoms and gave symbols to some elements as illustrated in Figure 1.1.

1.1 Atomic Theory

Competencies

By the end of this section, you will be able to:

- describe Dalton's atomic theory;
- describe modern atomic theory; and
- compare and contrast Dalton's atomic theory with modern atomic theory.

Activity 1.1



Form a group and discuss the following idea: Dalton's contribution was different from that of the ancient Greeks who postulated the existence of atoms. Point out the differences between the two ideas. Present your discussion to the class.

Even though the idea of the existence of atoms and atomic theory dates back to classical times, as discussed earlier, only Dalton's atomic theory ideas were the basis for the new era of science.

1.1.1 Dalton's Atomic Theory

Historical Note



John Dalton

British physicist and chemist John Dalton is best known for developing the atomic theory of elements and molecules, the foundation of modern physical science. While pondering the nature of the atmosphere during a meteorological study in the early 1800s, Dalton deduced the structure of carbon dioxide and proposed that an exact number of atoms constitute each molecule. He held that all atoms of a given element are identical and different from the atoms of every other element. The first to classify elements according to their atomic weights, Dalton set the stage for a revolution in scientific thought.

In 1804, John Dalton developed the first modern theory of atoms and proposed their existence. Dalton's atomic theory was based directly on the ideas of elements and compounds, and on the three laws of chemical combination. The three laws are:

i) The *law of conservation of mass* states that matter is neither created nor destroyed. This law is also called the law of indestructibility of matter. It means that the mass of the reactants is exactly equal to the mass of the products in any chemical reaction. A chemical reaction involves only the separation and union of atoms.

ii) The *law of definite proportions* states that a pure compound is always composed of the same elements combined in a definite ratio by mass. For example, water (H_2O) is composed of hydrogen and oxygen only. These elements are always in the proportion of 11.19% hydrogen to 88.81% oxygen by mass and in the proportion 2 : 1 by volume.

iii) The *law of multiple proportions* states that when two different compounds are formed from the same elements, the masses of one of the elements in the two compounds, compared to a given mass of the other element, is in a small whole-number ratio. For example, carbon and oxygen form two compounds: carbon monoxide and carbon dioxide. Carbon monoxide contains 1.3321 g of oxygen for each 1 g of carbon, whereas carbon dioxide contains 2.6642 g of oxygen for each 1g of carbon. Hence, carbon dioxide contains twice the mass of oxygen as does carbon monoxide.

Activity 1.2



Form a group and perform the following task. Present your findings to the class.

Using two chemical compounds as an example, describe the difference between the law of definite proportions and the law of multiple proportions.

Dalton proved that these laws are entirely reasonable if the elements are composed of tiny particles, which he called atoms**.** An atom is the smallest fundamental particle of an element.

The basic postulates of Dalton's Atomic Theory are summarized as follows:

1. All elements are made up of small particles called atoms.
2. Atoms are indivisible and indestructible.
3. All atoms of a given element are identical in mass and in all other properties.
4. Atoms are neither created nor destroyed in chemical reactions.
5. Compounds are formed when atoms of more than one element combine.
6. In a given compound, the relative numbers and types of atoms are constant.

Although some of these postulates, like postulate number 2 and number 3, have been shown to be incorrect by later work, Dalton's theory was a brilliant and logical explanation of many experimental discoveries and laws that were known at that time.

Activity 1.3



Form a group. Discuss the following laws and present to the class.
Use Dalton's atomic theory to explain:

a The law of conservation of mass
b The law of definite proportions
c The law of multiple proportions.

1.1.2 Modern Atomic Theory

Towards the end of the 19th century, various experimental discoveries revealed the existence of subatomic particles, isotopes, and so on. In light of these findings, Dalton's Atomic Theory was modified.

The Modern Atomic Theory can be summarized as follows:

1. Atoms are the smallest particles of all elements that can take part in a chemical reaction.
2. An atom is divisible. It can be subdivided into electrons, protons, and neutrons. An atom is also indestructible *i.e.*, atoms can neither be created nor destroyed during ordinary chemical reactions.
3. Atoms of the same element may not be identical in mass because of the existence of isotopes.
4. Atoms of the same elements have identical chemical properties.
5. Atoms of different elements have different chemical properties.
6. Atoms of two or more elements combine in simple whole-number ratios to form compounds.

Exercise 1.1

1. Name the postulates of Dalton's atomic theory.
2. Compare and contrast Dalton's atomic theory with modern atomic theory.

1.2 DISCOVERIES OF THE FUNDAMENTAL SUBATOMIC PARTICLES AND THE ATOMIC NUCLEUS

Competencies

By the end of this section, you will be able to:

- explain the discovery of the electron;
- explain the discovery of the nucleus;
- explain the discovery of the neutron.

Activity 1.4



Form a group and discuss the following ideas. Present your discussion to the rest of the class.

1. In your grade 7 chemistry lesson, you have learned that atoms consist of three fundamental subatomic particles. What are these? Where is their location in the atoms? What are their charges?
2. What were the consequences of the discoveries of subatomic particles on Dalton's atomic theory?

In the late 1880's, John Dalton thought atoms were indivisible. However, a series of investigations and astounding discoveries clearly demonstrated that atoms are made up of smaller particles, called subatomic particles. The three fundamental subatomic particles of an atom are electrons, protons, and neutrons.

1.2.1 Discovery of the Electron

Do you think that there is a similarity between cathode rays and electrons?

The electron was the first subatomic particle to be identified. In 1879, William Crooke studied electrical discharges in partially evacuated tubes called discharge tubes.

Two electrodes from a high-voltage source are sealed into a glass tube from which air has been evacuated. The negative electrode is the cathode and the positive one is the anode. When the high-voltage current is turned on, the glass tube emits a greenish light and a beam of light is seen at the anode. These rays flow from the cathode towards the anode in a straight line. They are called cathode rays. Later on, in 1897, J.J. Thomson found that the beam was deflected by both electrical and magnetic

fields. In an electric field, they bend toward the positive plate. From this he concluded that the cathode rays are made up of very small negatively charged particles, which he named electrons. He concluded that electrons are constituents of all matter, because he obtained the same results when he changed the gas and the electrode in the tube.

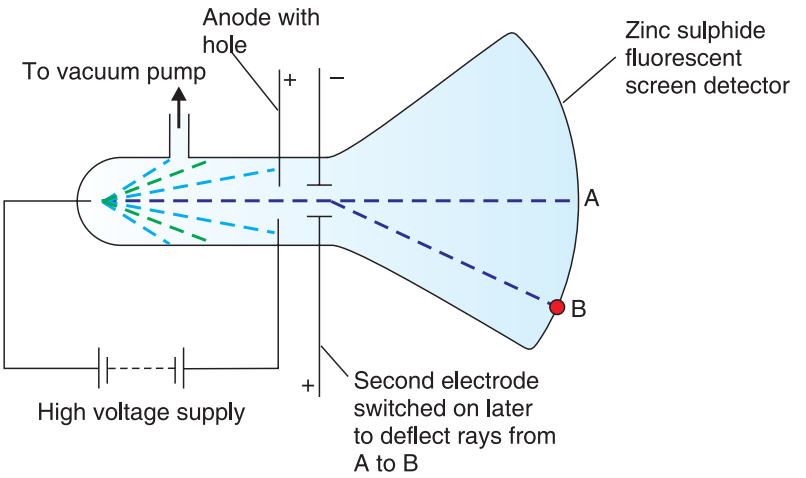


Figure 1.2 A simple cathode-ray tube.

Properties of Cathode Rays

Cathode rays possess the following properties:

1. An object placed between the cathode and the opposite end of the tube, casts a shadow on the glass. This shows that the cathode rays travel in a straight line.
2. A paddle wheel placed in the path of cathode rays rotates. This indicates that cathode rays are of particle nature-the particles strike the paddle and therefore move the wheel.

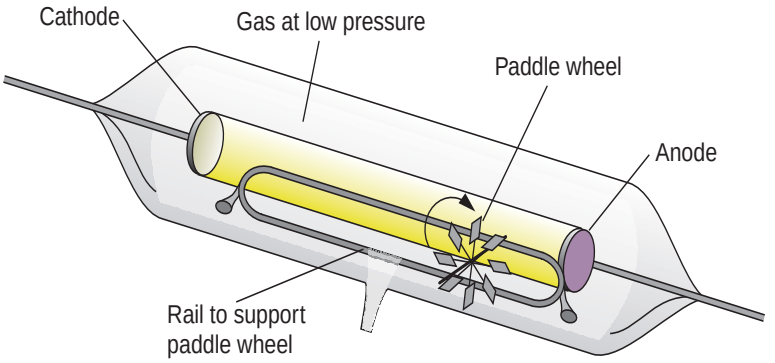


Figure 1.3 A paddle wheel placed in the path of cathode ray.

3. When an electric or magnetic field is applied in the path of cathode rays, they are deflected towards the positive plate. This shows that cathode rays are negatively charged.

4. The properties of cathode rays (like charge to mass ratio, e^-/m ratio) do not depend on the nature of the gas in the discharge tube or on the material of the cathode. This shows that electrons are present in the atoms of all elements.

Thus, cathode-ray experiments provided evidence that atoms are divisible into small particles, and that one of the atom's basic constituents is the negatively charged particle called electron.

PROJECT WORK

Produce a model of cathode ray tube from simple and locally available materials such as glass tube, copper wire and rubber stoppers. Using your model show how cathode rays are generated and explain their properties.

Charge and Mass of an Electron

J.J. Thomson studied the deflection of cathode rays under the simultaneous application of electric and magnetic fields, that are applied perpendicular to each other. His experiment led to the precise determination of the charge-to-mass ratio (e^-/m) of an electron, and this 'e^-/m' value was found to be 1.76×10^8 coulomb/g. Coulomb is a unit of electric charge.

Activity 1.5



Form a group and discuss how J.J Thomson's determination of the charge-to-mass ratio of the electron led to the conclusion that atoms were composed of sub atomic particles. Present your findings to the class.

In 1909, Robert Millikan determined the charge of an electron (e^-), using the oil drop experiment. He found the charge of an electron to be 1.60×10^{-19} coulombs.
Combination of e^-/m and e^- values are used to determine the mass of an electron, which is found to be 9.11×10^{-31} kg.
From Thomson's experiment, $e^-/m = 1.76 \times 10^8$ C/g
From Millikan's experiment, $e^- = 1.60 \times 10^{-19}$ C

$$\therefore \text{ Mass of the electron } = m = \frac{e^-}{e^-/m} = \frac{1.60 \times 10^{-19}\ \text{C}}{1.76 \times 10^{8}\ \text{C/g}}$$

$$= 9.11 \times 10^{-28}\ \text{g} = 9.11 \times 10^{-31}\ \text{kg}$$

From the above discussion, it follows that:

An electron is a fundamental particle of an atom carrying a negative charge and having a very small mass. The mass of an electron is approximately $\frac{1}{1837}$ times the mass of a hydrogen atom.

Exercise 1.2

1. Explain how the electron was discovered.
2. Give explanation on the works of J.J. Thomson and R. Millikan.

1.2.2 Discovery of the Atomic Nucleus

Radioactivity

In 1895, the German physicist, Wilhelm Röntgen noticed that cathode rays caused glass and metals to emit very unusual rays. This highly energetic radiation penetrated matter, darkened covered photographic plates, and caused a variety of substances to fluoresce. Since these rays could not be deflected by a magnet, they could not contain charged particles as cathode rays do. Röntgen called them X-rays.

Not long after Röntgen's discovery, Antoine Becquerel, a professor of physics in Paris, began to study fluorescent properties of substances. He found that exposing thickly wrapped photographic plates to a certain uranium compound caused them to darken, even without the stimulation of cathode rays. Like X-rays, the rays from the uranium compound were highly energetic and could not be deflected by a magnet. Also, these rays were generated spontaneously. One of Becquerel's students, Marie Curie, suggested the name radioactivity to describe this spontaneous emission of particles and/or radiation. Consequently, any element that spontaneously emits radiation is said to be radioactive.

Further investigation revealed that three types of rays are produced by the "decay", or breakdown of radioactive substances such as uranium, polonium, and radium. These are alpha, beta, and gamma rays.

Alpha (α) rays consist of positively charged particles called α-particles. They are deflected by positively charged plates. Beta (β) rays or β-particles are electrons of nuclear origin that are deflected by negatively charged plates. The third type of radioactive radiation consists of high-energy rays called gamma (γ) rays. Like X-rays, γ-rays have no charge and are not affected by an external electric or magnetic field.

Activity 1.6



Form a group and discuss the following ideas. Present your discussion to the class.

Scientists discovered the three subatomic particles (electrons, protons and neutrons). Compare and contrast these fundamental sub atomic particles with alpha particles, beta particles and gamma rays in terms of the nature of the particles.

More details about the structure of an atom were provided in 1911 by Ernest Rutherford and his associates, Hans Geiger and Ernest Marsden. The scientists

bombarded a thin gold foil with fast moving alpha particles. Alpha particles are positively charged particles with about four times the mass of a hydrogen atom. Geiger and Marsden assumed that mass and charge were uniformly distributed throughout the atoms of the gold foil. Therefore, they expected the alpha particles to pass through the gold foil with only a slight deflection. However, when the scientists checked for the possibility of wide-angle deflections, they were shocked to find that roughly 1 in 8000 of the alpha particles had actually been redirected back toward the source.

After thinking about the observations for two years, Rutherford finally came up with an explanation. He reasoned that the rebounded alpha particles must have experienced some powerful force within the atom. He concluded that the source of this force must occupy a very small amount of space because very few of the total number of alpha particles had been affected by it. He also concluded that the force must be caused by a very densely packed bundle of matter with a positive electric charge. Rutherford called this small, dense, positive bundle of matter the nucleus.

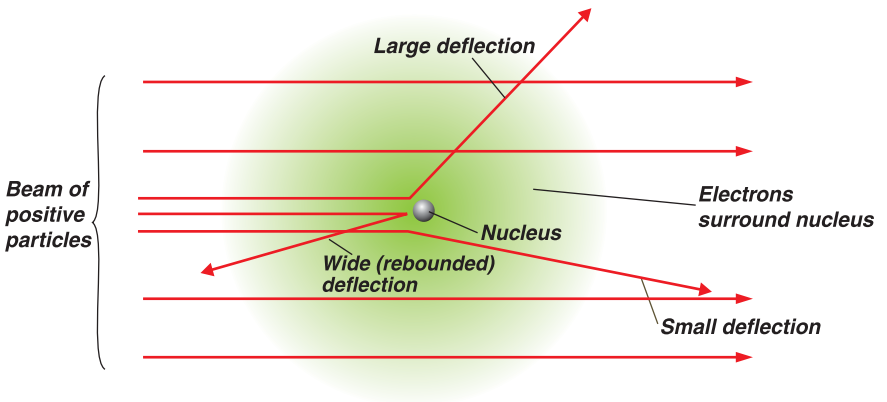


Figure 1.4 Rutherford's α – particles experiment.

Rutherford's Conclusion

1. Since most of the α-particles passed through the gold foil undeflected, most of the space in an atom is empty.
2. Some of the α-particles were deflected by small angles. This indicated the presence of a heavy positive centre in the atom, which Rutherford named the nucleus.
3. Only a few particles (1 in about a million) were either deflected by a very large angle or deflected back. This confirmed that the space occupied by the heavy positive centre must be very small.

Activity 1.7



Form a group and perform the following task. Present your findings to the class.

Predict what Rutherford might have observed if he had bombarded copper metal instead of gold metal with alpha particles.

From the angles through which alpha particles are deflected, Rutherford calculated that the nucleus of an atom has a radius of about 10^{-15} m. The radius of the whole atom is about 10^{-10} m. Therefore, the nucleus is about 1/10,000 (one ten-thousandth) of the size of the atom as a whole. Thus, if we magnified an atom to the size of a football stadium (about 100 m across), the nucleus would be represented by a pea placed at the centre of the pitch.

Rutherford suggested that the negatively charged electrons surrounding the positively charged nucleus are like planets around the sun revolving. He could not explain, however, what kept the electrons in motion around the nucleus.

1.2.3 Discovery of Neutrons

In 1932, an English scientist, James Chadwick, identified the existence of neutrons. He bombarded a thin foil of beryllium with α-particles of a radioactive substance. He then observed that highly penetrating rays, consisting of electrically neutral particles of a mass approximately equal to that of the proton, were produced. These neutral particles are called neutrons.

A neutron is a subatomic particle carrying no charge and having a mass of 1.675×10^{-24} g. This mass is almost equal to that of a proton or of a hydrogen atom.

Exercise 1.3

1. Which experimental evidence indicates that:
   i) electrons are negatively charged particles?
   ii) electrons are the constituents of all atoms?
2. In Rutherford's experiment, most of the alpha particles passed through the gold foil undeflected. What does this indicate?
3. In Rutherford's experiment, some of the α-particles were deflected by small angles. What does this indicate?
4. What was determined from Millikan's oil drop experiment?
5. Three particles are fired into a box that is positively charged on one side and negatively charged on the other side. The result is shown in Figure 1.5.

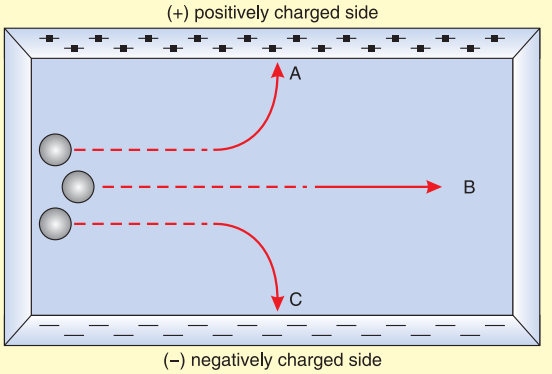


Figure 1.5 particles fired into a box.

a Which particle is a proton?

b Which particle is an electron?

c Which particle is a neutron?

Activity 1.8



Hydrogen is the most abundant element in the universe. It is the fuel for the sun and other stars. It is currently believed that there are roughly 2,000 times more hydrogen atoms than oxygen atoms. There are also 10,000 times more hydrogen atoms than carbon atoms.

Make a model of a hydrogen atom, using materials of your choice, to represent a hydrogen atom, including the proton and electron. Present the model to the class, and explain in what ways your model resembles a hydrogen atom.

1.3 COMPOSITION OF AN ATOM AND ISOTOPES

Competencies

By the end of this section, you will be able to:

- write the relative charges of an electron, a proton, and a neutron;
- tell the absolute and relative masses of an electron, a proton, and a neutron;
- tell the number of protons and electrons in an atom from the atomic number of the element;
- determine the number of neutrons from given values of atomic number and mass number;

- explain the terms atomic mass and isotope, and
- calculate the atomic masses of elements that have isotopes.

Activity 1.9



Based on your previous knowledge discuss the following concepts in groups, and share your ideas with the rest of the class.

1. Would you think that it is possible for someone to discover a new element that would fit between magnesium (atomic number 12) and aluminum (atomic number 13)?
2. You are given the following three findings:
   - a An atom of calcium has a mass of 40 a.m.u. Another atom of calcium has 44 a.m.u.
   - b An atom of calcium has a mass of 40 a.m.u. An atom of potassium has a mass of 40 a.m.u.
   - c An atom of calcium has a mass of 40 a.m.u. An atom of cobalt has a mass of 59 a.m.u.

Are these findings in agreement with Dalton's atomic theory? Discuss your conclusions in your group.

All atoms have two regions. The nucleus is a very small region located at the center of an atom. Except for the nucleus of the simplest type of hydrogen atom, all atomic nuclei are made of two kinds of particles, protons and neutrons. Surrounding the nucleus, there is a region occupied by negatively charged particles called electrons. This region is very large compared with the size of the nucleus. A proton has a positive charge, which is equal in magnitude to the negative charge of an electron. Atoms are electrically neutral because they contain equal numbers of protons and electrons.

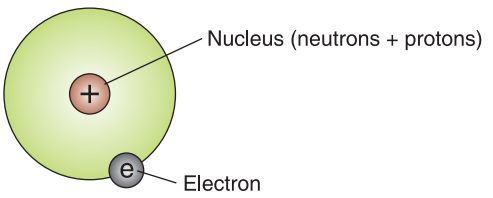


Figure 1.6 A simple representation of an atom.

A proton has a mass of 1.673×10^{-27} kg, and a neutron has a mass of 1.675×10^{-27} kg. Hence, protons and neutrons have approximately the same mass. Most of the mass of the atom is concentrated in the nucleus, assuming the mass of an electron to be negligible or almost zero.

Activity 1.10



Make a group and perform the following task: Sketch a modern atomic model of a chlorine atom and identify where each type of subatomic particles would be located. Submit the model to your teacher.

Activity 1.11



Make a group and discuss, "If atoms are primarily composed of empty space, why can't you pass your hand through a solid object"? Present your finding to the class.

Physicists have identified other subatomic particles. But particles other than electrons, protons, and neutrons have little effect on the chemical properties of matter. Table 1.1 gives a summary of the properties of electrons, protons, and neutrons. (The relative electric charge, relative mass, and actual mass are discussed in the next section.)

Table 1.1 Properties of Subatomic Particles.

Particle	Symbols	Relative electric charge	Relative mass (a.m.u)	Actual mass (kg)
Electron	e^-	–1	0.0005486 ≈ 0	9.109×10^{-31}
Proton	p^+	+1	1.007276 ≈ 1	1.673×10^{-27}
Neutron	n^o	0	1.008665 ≈ 1	1.675×10^{-27}

It is convenient to think of the region occupied by the electrons as an electron cloud, a cloud of negative charge. The radius of an atom is the distance from the center of the nucleus to the outer portion of this electron cloud. This is about 10^{-10} m = 100 pm = 0.1 nm. Because atomic radii are so small, they are expressed using a unit that is specifically convenient for the sizes of atoms. This unit can be picometer, nanometer, micrometer, etc. The abbreviation for the

picometer is pm (1 pm = 10^{-12} m = 10^{-10} cm)
nanometer is nm (1 nm = 10^{-9} m = 10^{-7} cm)
micrometer is μm (1μm = 10^{-6} m = 10^{-4} cm)

1.3.1 Atomic Number and Mass Number

Activity 1.12



Form a group and perform the following task. Present your findings to the class.

The table below gives the number of electrons, protons and neutrons in atoms or ions for a number of elements.

Atom or ion of elements	A	B	C	D	E	F	G
Number of electrons	5	7	9	12	7	6	9
Number of protons	5	7	10	10	7	5	9
Number of neutrons	5	7	10	10	8	6	10

Based on the table, answer the following questions:

i) Which of the species are neutral?

ii) Which of them are negatively charged?

iii) Give the symbolic representations for 'B' , 'D' and 'F'

All atoms are composed of the same basic particles. Yet all atoms are not the same. Atoms of different elements have different number of protons. Atoms of the same element have the same number of protons.

The Atomic Number (Z) of an element is equal to the number of protons in the nucleus of an atom. It is also equal to the number of electrons in the neutral atom.

$$Z = p^+ \quad \text{or} \quad Z = e^- \qquad \text{(in a natural atom)}$$

For different elements, the nuclei of their atoms differ in the number of protons they contain. Therefore, they differ in the amount of positive charge that their atoms possess. Thus, the number of protons in the nucleus of an atom determines the identity of an atom.

Each atom of an element is identified by its atomic number. In the periodic table, the atomic number of an element is indicated below its symbol. Notice that the elements are placed in order of increasing atomic number. Hydrogen, H, has atomic number 1, hence, atoms of the element hydrogen have one proton in the nucleus. Next is helium, He, which has two protons in its nucleus. Lithium, Li, has three protons. Beryllium, Be, has four protons, and so on.

Mass Number (A) is the sum of the number of protons and neutrons in the nucleus of an atom. Collectively, the protons and neutrons of an atom are called its nucleons.

Mass Number (A) = Number of protons + Number of neutrons

$$A = p^+ + n^\circ$$

The name of an element, its atomic number and mass number can be represented by a shorthand symbol. For example, to represent neutral atoms of magnesium with 12 protons and 12 neutrons, the symbol is $^{24}_{12}Mg$. The atomic number is written as a subscript to the left of the symbol of magnesium. The mass number is written as a superscript to the left of the symbol. This method can be represented in general by the following illustration, in which X stands for any chemical symbol.

$$\begin{matrix}\text{Mass Number}\\ \text{Atomic Number}\end{matrix}X \quad \text{or} \quad {}^{A}_{Z}X$$

$^{24}_{12}Mg$ can be read as "Magnesium-24", and "Magnesium-25" represents $^{25}_{12}Mg$. In a similar manner, $^{23}_{11}Na$ can be read as "Sodium-23".

The number of neutrons in an atom is found by subtracting the atomic number from the mass number. For example, the number of neutrons for a silver atom ($^{108}_{47}Ag$) is:

Mass number – Atomic number = Number of neutrons

$$108 - 47 = 61$$

$\therefore$ The number of neutrons is 61.

Example 1

How many protons, electrons, and neutrons are found in an atom of bromine-80 if its atomic number is 35?

Solution:

Atomic number of bromine is 35.

Mass number of bromine is already given as 80. Now we can write bromine, Br in the form of:

$$^{A}_{Z}X \quad \text{as} \quad {}^{A}_{Z}Br = {}^{80}_{35}Br$$

Mass number – Atomic number = Number of neutrons

$$80 - 35 = 45 \text{ neutrons}$$

For a neutral atom: number of protons = number of electrons = atomic number.

$\therefore$ Number of protons = Number of electrons = 35

Activity 1.13



Discuss the following idea in groups and present your discussion to the class.

Why do all atoms of a chemical element have the same atomic number although they may have different mass numbers?

Exercise 1.4

Determine the atomic number, mass number, and number of neutrons in

a P b $^{39}_{19}K$ c $^{56}_{26}Fe$

1.3.2 Isotopes and Atomic Mass

The simplest atoms are those of hydrogen. Each hydrogen atom contains one proton only. However, hydrogen atoms contain different numbers of neutrons.

Three types of hydrogen atoms are known. The most common type of hydrogen is called protium. It accounts for 99.985% of the hydrogen atoms found on Earth. The nucleus of a protium atom consists of one proton only, and it has one electron moving around it. The second form of hydrogen atom is called deuterium, which accounts for 0.015% of the Earth's hydrogen atoms. Each deuterium atom has a nucleus containing

Activity 1.14



Form a group and discuss the following idea and present your discussion to the class.

Nitrogen has two naturally occurring isotopes, N-14 and N-15. The atomic mass of nitrogen is 14.007. Which isotope is more abundant in nature? Explain.

one proton and one neutron. The third form of hydrogen atom is known as tritium, and it is radioactive. It exists in very small amounts in nature, but it can be prepared artificially. Each tritium atom contains one proton and two neutrons. Protium, deuterium and tritium are the three isotopes of hydrogen.

Isotopes are the atoms of the same element that have the same atomic number but different masses number. The isotopes of a particular element have the same number of protons and electrons but different number of neutrons. Isotopes of an element may occur naturally, or they may be made in the laboratory (artificial isotopes). Since isotopes have the same number of protons and electrons, they have the same chemical properties.

Designating Isotopes

Isotopes are usually identified by specifying their mass number. There are two methods for specifying isotopes. In the first method, the mass number is written with a hyphen after the name of the element. Tritium, for example, is written as Hydrogen-3. We will refer to this method as hyphen notation. The second method for

designating an isotope is by indicating its isotopic nuclear composition. For example, Hydrogen-3 is written as $^{3}_{1}H$. The following table illustrates this designation of isotopes for hydrogen and helium.

Table 1.2 Isotopes of Hydrogen and Helium.

Isotope	Nuclear symbol	No. of protons	No. of electrons	No. of neutrons	Z	A
Hydrogen-1 (protium)	$^{1}_{1}H$	1	1	0	1	1
Hydrogen-2 (deuterium)	$^{2}_{1}H$	1	1	1	1	2
Hydrogen-3 (tritium)	$^{3}_{1}H$	1	1	2	1	3
Helium-3	$^{3}_{2}He$	2	2	1	2	3
Helium-4	$^{4}_{2}He$	2	2	2	2	4

Exercise 1.5

1. Write the nuclear symbol for carbon-13 and carbon-12.
2. The three atoms of elements X, Y and Z are represented as: $^{23}_{11}X$, $^{27}_{13}Y$ and $^{31}_{15}Z$. Determine the number of protons, electrons, and neutrons in the neutral atoms indicated.

Relative Atomic Masses

Actual masses of atoms, measured in grams, are very small. An atom of oxygen-16, for example, has a mass of 2.657×10^{-23}g. For most chemical calculations, it is more convenient to use relative atomic masses. In order to set up a relative scale of atomic mass, one atom has been arbitrarily chosen as the standard and assigned a relative mass value. The masses of all other atoms are expressed in terms of this defined standard.

The standard used by scientists to determine units of atomic mass is the carbon-12 atom. It has been arbitrarily assigned a mass of exactly 12 atomic mass units, or 12 a.m.u. One atomic mass unit, or 1 a.m.u, is exactly 1/12 the mass of a carbon-12 atom, or 1.660×10^{-24} g. The atomic mass of any atom is determined by comparing it with the mass of the carbon-12 atom. The hydrogen-1 atom has an atomic mass of about 1/12 of that of the carbon-12 atom, or about 1 a.m.u. The precise value of the atomic mass of a hydrogen-1 atom is 1.007 a.m.u. An oxygen-16 atom has about 16/12 (or 4/3 the mass of a carbon-12 atom). Careful measurements show the atomic mass of oxygen-16 to be 15.994 a.m.u. The mass of a magnesium-24 atom is found to be slightly less than twice that of a carbon-12 atom. Its atomic mass is 23.985 a.m.u.

The masses of subatomic particles can also be expressed on the atomic-mass scale. The mass of an electron is 0.0005486 a.m.u, that of the proton is 1.007276 a.m.u, and that of the neutron is 1.008665 a.m.u. Note that the proton and neutron masses are close to 1 a.m.u but not equal to 1 a.m.u. You have learned that the mass number is the total number of protons and neutrons in the nucleus of an atom. You can now see that the mass number and relative atomic mass of a given nucleus are quite close to each other. They are not identical because the proton and neutron masses deviate slightly from 1 a.m.u.

Activity 1.15



Form a group and perform the following task, based on the given information:

Magnesium has three naturally occurring isotopes, namely Mg-12, Mg-13, and Mg-14.

Make models of three isotopes of magnesium using locally available materials such as clay, colours and sticks, and depict the nucleus, the positions of protons, electrons and neutrons.

Present your model to the class.

Average Atomic Masses of Elements

Most elements occur naturally as mixtures of isotopes, as shown in the examples in Table 1.3. The percentage of each isotope in the naturally occurring elements is nearly always the same, no matter where the element is found. The percentage at which each isotope of an element occurs in nature is taken into account when calculating the average atomic mass of the element. Average atomic mass is the weighted average of the atomic masses of the naturally occurring isotopes of an element.

Table 1.3 Atomic masses and abundances of some natural occurring isotopes.

Isotopes	Mass number	% natural abundance	Atomic mass (a.m.u)	Average atomic mass of element (a.m.u)
Hydrogen-1	1	99.985	1.6078	1.00794
Hydrogen-2	2	0.015	2.0141	
Oxygen-16	16	99.762	15.994	15.9994
Oxygen-17	17	0.038	16.999	
Oxygen-18	18	0.200	17.999	

* Since tritium is radioactive and exists in a very small amount in nature, it is not included in the calculation of atomic mass of hydrogen.

The average atomic mass of an element is calculated by summing up the products of relative atomic mass and percentage abundance of each isotope and dividing by 100.

Example 2

Chlorine has two isotopic compositions: 75.77 % of $^{35}_{17}Cl$ and 24.23 % of $^{37}_{17}Cl$. Determine the average atomic mass of chlorine.

Solution:

75.77 % of natural chlorine exists as chlorine 35, and the rest 24.23 % as chlorine 37.

The average atomic mass of chlorine is calculated as follows:

$$= \frac{(75.77 \times 35) + (24.23 \times 37)}{100} = 35.45$$

Thus, the average atomic mass of chlorine is 35.45 a.m.u

Example 3

Calculate the average atomic mass of boron, using the following data:

Isotope	Relative masses	Percent abundance
$^{10}_{5}B$	10.0134	19.70
$^{11}_{5}B$	11.0094	80.30

Solution:

The average atomic mass of boron is obtained by taking the sum of the products of the relative atomic mass and the percentage abundance of the first and the second isotopes of boron and dividing by 100.

$$\text{Average atomic mass of B} = \frac{(10.0134 \times 19.70) + (11.0094 \times 80.30)}{100}$$

$$= 10.813 \text{ a.m.u}$$

Exercise 1.6

Give appropriate answers for the following questions.

1. What are isotopes? Explain their symbolic designation using the isotopes of hydrogen.

2. The three isotopes of uranium are: $^{234}_{92}U$, $^{235}_{92}U$ and $^{238}_{92}U$. How many protons, neutrons and electrons are present in each isotope?
3. Indicate the number of fundamental particles present in an atom of $^{206}_{82}Pb$.
4. The metal thallium occurs naturally as 30% thallium-203 and 70% thallium-205. Calculate the atomic mass of thallium.

Critical Thinking

5. Why is the gravitational force in the nucleus so small?
6. Could a nucleus of more than one proton but no neutron exist? Explain.

1.4 ATOMIC MODELS

Competencies

By the end of this section, you will be able to:

- name the five atomic models,
- describe Dalton's, Thomson's and Rutherford's atomic models,
- state Bohr's postulates,
- describe Bohr's model,
- describe the quantum mechanical model,
- describe main-energy level and sub-energy level,
- define the term electronic configuration,
- write the ground state-electronic configuration of the elements,
- draw diagrams to show the electronic configuration of the first 18 elements,
- write the electronic configuration of the elements using sub–energy levels,
- write electronic configuration of elements using noble gas as a core and
- describe valence electrons.

Activity 1.16



Discuss the following concept in groups and present your conclusion to the class. Have you heard the word "*model*" in your everyday life? Car makers produce different models as time passes. What is your understanding about the atomic model?

Atoms are too small to be observed directly. Hence, it is better to develop a tentative mental picture (*model*) of the atomic concept. The ideas about atoms have changed

many times. But, at present we can reduce these ideas into five models: Dalton's Model, Thomson's Model, Rutherford's Model, Bohr's Model, and the Quantum Mechanical Model.

i) Dalton's Atomic Model

This model of atoms was accepted for 100 years without serious challenge. He thought of atoms as solid indestructible spheres. He called it the billiard ball model.

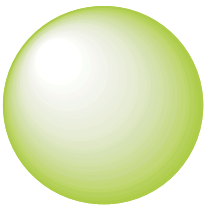


Figure 1.7 Dalton's model of an atom.

ii) Thomson Model

After learning that atoms contain electrons, Thomson proposed a new model of the atom, which is shown in Figure 1.8. It is sometimes called the plum-pudding model. In his model of an atom, Thomson proposed thought that electrons were embedded inside a positively charged sphere, just like plums in a pudding. Today we might call Thomson's model the *chocolate chip ice-cream model*.

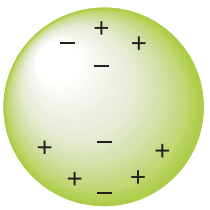


Figure 1.8 Thomson's model of an atom.

Activity 1.17



Perform the following tasks in groups and present your findings to the class.

1. Draw a sketch diagram of a carbon atom using Thomson's atomic model.
2. What was incorrect about Thomson's model of the atom?

iii) Rutherford Model

Why was Thomson's model of atom discarded and replaced by Rutherford's model?

By 1911, evidence was collected that allowed scientists to modify Thomson's model of an atom. Ernest Rutherford (1871–1937), a former student of Thomson, performed one of the classic experiments of scientific history. As indicated earlier in this unit, he bombarded different types of matter with high-energy, positively charged alpha particles. Based on the Thomson's model of an atom, Rutherford hypothesized that the alpha particles should go through very thin metal foils undeflected. However, after performing an experiment on gold leaf, he found that a small but significant number of the alpha particles were deflected through large angles by the gold atoms. These results showed Rutherford that the Thomson model of the atom was not valid. A uniform positive charge with embedded negatively charged electrons would not interact with the alpha particles in such a manner.

The Rutherford model of the atom has a small, dense, positively charged nucleus around which electrons whirl at high speeds and at relatively long distances from it. He compared the structure of an atom with the solar system, saying that the nucleus corresponds to the sun, while the electrons correspond to the planets. This picture of the atom is also called the planetary atom. In other words, Rutherford gave us the nuclear model of the atom.

Even though Rutherford showed that the atom has a nucleus, he did not know how the electrons were arranged outside the nucleus. This was the major drawback of Rutherford's model of an atom.

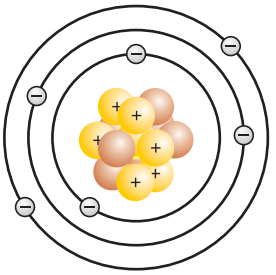
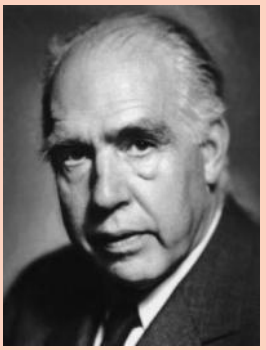


Figure 1.9 Rutherford model.

Historical Note



Niels Bohr

A Nobel Prize winner, Niels Bohr, was known not only for his own theoretical work, but also as a mentor to younger physicists who themselves made important contributions to physical theory. As the director of the Institute for Theoretical Physics at the University of Copenhagen, Bohr gathered together some of the finest minds in the physics community, including Werner Heisenberg and George Gawow. During the 1920's, the Institute was the source of many important works in quantum mechanics and in theoretical physics in general.

iv) Bohr Model

In 1913, the Danish physicist Niels Bohr (1885–1962) proposed an atomic model in which the electrons moved around the nucleus in circular paths called orbits. He assumed the electrons to be moving around the nucleus in a circular orbit as the planets move around the sun. Based on the Rutherford's atomic model, Bohr made the following modifications.

- The electrons in an atom can exist only in a restricted number of stable orbits with energy levels in which they neither absorb nor emit energy. These orbits designated by a number called the principal quantum number, n. The principal quantum number has the values of 1, 2, 3, . . . The energy levels are also called shells. They are represented by K, L, M, . . . etc. The K-shell is the first shell, the L-shell is the second shell, the M-shell is the third shell, etc.
- When an electron moves between orbits it absorbs or emits energy. When an electron jumps from lower to higher states it absorbs a fixed amount of energy. When an electron falls from a higher (excited) state to a lower (ground) state it emits a fixed amount of energy.
- The electrons move around the nucleus in energy levels.

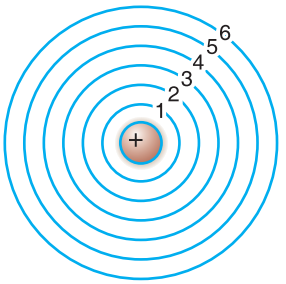


Figure 1.10 The first six energy levels in the Bohr model of an atom.

- According to Bohr, for each element, the number of energy levels or shells is fixed. Also, each different energy level or shell of an atom can only accommodate a certain number of electrons. The maximum number of electrons that the main energy level can have is given by the general formula $2n^2$, where n is equal to the main energy level.

Therefore, the maximum number of electrons in each shell in the Bohr model is:

first energy level	(n = 1) is $2(1)^2$	= 2 electrons
second energy level	(n = 2) is $2(2)^2$	= 8 electrons
third energy level	(n = 3) is $2(3)^2$	= 18 electrons
fourth energy level	(n = 4) is $2(4)^2$	= 32 electrons

It can also be summarized in the form of Table 1.4.

Table 1.4 Energy levels and electrons for the first four main energy levels.

Principal quantum number (*n*)	1	2	3	4
Shell	K	L	M	N
Maximum possible number of electrons	2	8	18	32

- When electrons fill the various energy levels in an atom, they first occupy the shell with the lowest energy level. When the lowest energy level is filled, according to the $2n^2$ rule, then the electrons enter the next higher energy level.

The outermost shell of an atom cannot accommodate more than eight electrons, even if it has the capacity to hold more electrons. This is because having more than eight electrons in the outermost shell makes the atom unstable.

For example, let us write the electron configuration of magnesium that has 12 electrons. The first 2 electrons occupy the K-shell (the first energy level). The L-shell (the second energy level) is occupied with 8 electrons. The remaining 2 electrons enter the M-shell (the third energy level). Hence, the electron configuration of magnesium atom becomes $\begin{matrix}\text{K}&\text{L}&\text{M}\\2&8&2\end{matrix}$.

Bohr's diagrammatic representation, using shells, in some atoms are given below. Note that *p* = proton, *n* = neutron, and ● = electron.

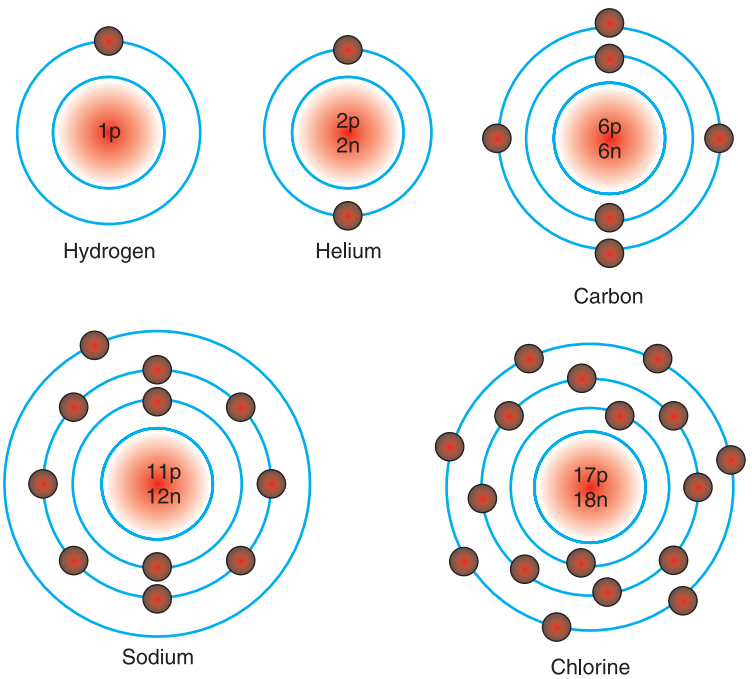


Figure 1.11 Diagrammatic representation of Bohr's model.

Activity 1.18



Form a group and try to compare Rutherford's atomic model with Bohr's atomic model? Present your discussion to the class.

Exercise 1.7

Draw Bohr's diagrammatic representations for the following atoms:

a Nitrogen (atomic number = 7)

b Sulphur (atomic number = 16)

c Potassium (atomic number = 19)

d Calcium (atomic number = 20)

v) The Quantum Mechanical Model

During the 1920's, the discovery of the wave-like properties of electrons led to the wave-mechanical model or quantum-mechanical model of atoms. In this model, the electrons are associated with definite energy levels, but their locations cannot be pinpointed. Instead, they are described in terms of the probability of being found in certain regions of space about the nucleus. These regions of space are called orbitals. An orbital is a particularly shaped volume of space where the probability of finding an electron is at a maximum.

According to this model, the electrons in an atom are in a series of energy levels or shells. Each energy level or shell is described by a number called the principal quantum number, *n*, which is related to the size of the energy level. The larger the energy value of *n*, the farther the electrons are found from the nucleus.

The quantum mechanical model introduces the concept of sublevel for each main energy level. For atoms of the known elements, there are four types of sublevels designated by the letters *s*, *p*, *d*, and *f*. The number of sublevels within each energy level or shell is equal to the numerical value of *n*.

Table 1.5 The principal quantum number specified with shell, sublevels and maximum number of electrons.

Principal quantum number (n)	Shell	Number and type of sublevels each shell ($2n^2$)	Maximum number of electrons in
1	K	1 (*s*)	2
2	L	2 (*s*, *p*)	8
3	M	3 (*s*, *p*, *d*)	18
4	N	4 (*s*, *p*, *d*, *f*)	32
5	O	5 (*s*, *p*, *d*, *f*, *g*)	50
6	P	6 (*s*, *p*, *d*, *f*, *g*, *h*)	72

Any orbital can accommodate a maximum of two electrons. Accordingly,

s-sublevel has only 1 orbital and can hold a maximum of 2 electrons,

p-sublevel has 3 orbitals and can hold a maximum of 6 electrons,

d-sublevel has 5 orbitals and can hold a maximum of 10 electrons, and

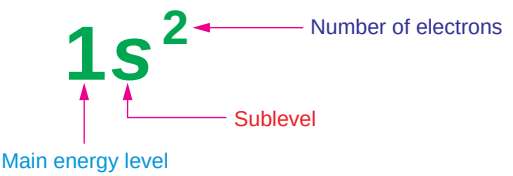
f-sublevel has 7 orbitals and can hold a maximum of 14 electrons.

Electron Configurations

The arrangement of electrons in an atom is known as the electron configuration of the atom. Because atoms of different elements have different numbers of electrons, a distinct electronic configuration exists for the atoms of each element. Like all systems in nature, electrons in atoms tend to assume arrangements that have the lowest possible energies. The lowest energy arrangement of the electrons in an atom is called the ground state electron configuration. A few simple rules, combined with the quantum number relationships discussed below, allow us to determine these ground-state electron configurations.

The quantum mechanical model is designated by the following notation: a coefficient which shows the main energy level, a letter that denotes the sublevel that an electron occupies, and a superscript that shows the number of electrons in that particular sublevel.

The designation is explained as follows



For example, the electron configuration of lithium ($^{5}_{3}$Li) is: $1s^2 2s^1$.

This indicates that there are 2 electrons in the first *s*-sublevel and 1 electron in the 2nd *s*-sublevel.

The configuration for sodium ($^{23}_{11}$Na) atom is: $1s^2 2s^2 2p^6 3s^1$. This indicate that there are 2 electrons in the first *s*-sublevel, 2 electrons in the second *s*-sublevel, 6 electrons in the second *p*-sublevel, and 1 electron in the third *s*-sublevel.

Activity 1.19



Form a group and perform the following task in groups and present your findings to the class.

The following ground state electron configurations are incorrect. Identify in each case the mistakes that have been made and write the correct electron configurations.

i) $1s^2\ 2s^2\ 2p^4\ 3s^2$

ii) $1s^2\ 2s^1\ 2p^6$

iii) $1s^2\ 2s^2\ 2p^6\ 3p^3\ 3s^2$

iv) $1s^2\ 2s^2\ 2p^6 3s^2\ 3p^3\ 4s^2\ 3d^4$

The following table lists the first 14 elements in the periodic table and shows the electron configuration of each.

Table 1.6 The ground state electron configuration of the first 14 elements.

Element	Electron Configuration
Hydrogen	$1s^1$
Helium	$1s^2$
Lithium	$1s^2\ 2s^1$
Beryllium	$1s^2\ 2s^2$
Boron	$1s^2\ 2s^2\ 2p^1$
Carbon	$1s^2\ 2s^2\ 2p^2$
Nitrogen	$1s^2\ 2s^2\ 2p^3$
Oxygen	$1s^2\ 2s^2\ 2p^4$
Fluorine	$1s^2\ 2s^2\ 2p^5$
Neon	$1s^2\ 2s^2\ 2p^6$
Sodium	$1s^2\ 2s^2\ 2p^6\ 3s^1$
Magnesium	$1s^2\ 2s^2\ 2p^6\ 3s^2$
Aluminium	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
Silicon	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

Exercise 1.8

Write the ground state electron configuration for the following elements.

a Phosphorus (atomic number = 15)
b Sulphur (atomic number = 16)
c Chlorine (atomic number = 17)
d Argon (atomic number = 18)

Rules Governing Electron Configuration

Most of the electronic configuration of an atom can be explained in terms of the building-up principle, or also known as the Aufbau principle. According to the Aufbau principle, an electron occupies the lowest energy orbital available before entering a higher energy orbital.

The atomic orbitals are filled in order of increasing energy. The orbital with the lowest energy is the 1*s* orbital. The 2*s* orbital is the next higher in energy, then the 2*p* orbitals, and so on. Beginning with the third main energy level, n = 3, the energies of the sublevels in the different main energy levels begin to overlap.

In Figure 1.12, for example, the 4*s* sublevel is lower in energy than the 3*d* sublevel. Therefore, the 4*s* orbital is filled before any electrons enter the 3*d* orbitals. Less energy is required for two electrons to pair up in the 4*s* orbital than for a single electron to occupy a 3*d* orbital. Which sublevel will be occupied after the 3*d* sublevel is fully occupied?

The electrons are arranged in sublevels according to Aufbau principle which is also known as the diagonal rule. The diagonal rule is a guide to the order of filling energy sublevels. It is particularly helpful for atoms with atomic numbers higher than 18. This is because for atoms with higher atomic numbers their sublevels are not regularly filled.

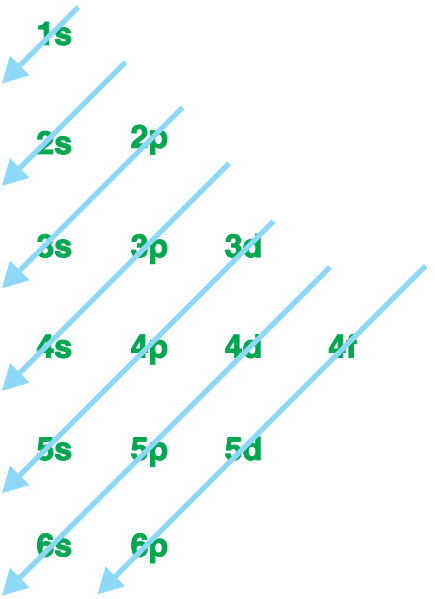


Figure 1.12 Diagonal rule for writing electron configuration.

If you list the orbitals following the direction of arrows, you can obtain the electron configuration of most atoms. Therefore, the order of filling the sublevels is given as:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6$...

The outermost occupied energy level of an atom is called the valence shell and the electrons that enter into these energy levels are referred to as valence electrons.

Example 1

What is the electron configuration of magnesium, $^{24}_{12}Mg$?

Solution:

$$^{24}_{12}Mg = \underbrace{1s^2}_{2}\ \underbrace{2s^2 2p^6}_{8}\ \underbrace{3s^2}_{2} \quad \Rightarrow \text{in sublevel}$$
$$\Rightarrow \text{in main energy level}$$

From this, we can write the configuration of elements by using a noble gas as a core and the electrons outside the core. Note that noble gases have complete outer electron shells, with 2 or 8 electrons. Therefore, the above electron configuration can be written as follows:

$$^{24}_{12}Mg = \underbrace{1s^2 2s^2 2p^6\ 3s^2}_{[Ne]}$$

$$= [Ne]\ 3s^2$$

Example 2

Write the electron configuration of copper (Z = 29).

Solution:

Following the diagonal rule, we can fill the sublevels, in order of increasing energy, until all its electrons are filled. Therefore, the configuration will be:

$$Cu = \underbrace{1s^2 2s^2 2p^6 3s^2 3p^6}_{[Ar]}\ 4s^1 3d^{10}$$

$$= [Ar]\ 4s^1 3d^{10}$$

Exercise 1.9

1. Write the electron configuration of the following elements using noble gases as a core.

   a O b Al c Cl d Ca

2. Write the electron configuration of the following elements.

a $^{84}_{36}Kr$ b $^{56}_{26}Fe$ c $^{65}_{30}Zn$

Check list

Key terms of the unit

- *Atom*
- *Atomic Model*
- *Aufbau principle*
- *Average atomic mass*
- *Bohr's Model*
- *Cathode ray*
- *Dalton's atomic theory*
- *Dalton's Model*
- *Electron*
- *Electron configuration*
- *Electron configuration*
- *Isotopes*
- *Main energy level*
- *Modern atomic theory*
- *Neutron*
- *Nucleus*
- *Proton*
- *Quantum Mechanical Model*
- *Radioactivity*
  - *α - particles*
  - *β - particles*
  - *γ - particles*
- *Rutherford's Model*
- *Sub-energy level*
- *Thomson's Model*

Unit Summary

The Greek philosopher Democritus proposed what we consider as the first atomic theory. He believed that all matter consists of very small, indivisible particles. He named the particles atoms (meaning uncuttable or indivisible).

John Dalton's work in atomic theory marked the start of modern chemistry. His concept of an atom was far more detailed and specific than Democritus'.

Dalton did not try to describe the structure or composition of atoms – he did not know what an atom really looked like. But he did realize that the different properties shown by elements such as hydrogen and oxygen can be explained by assuming that hydrogen atoms are not the same as oxygen atoms.

In our current view of atomic theory, atoms consist of smaller particles – electrons, protons, and neutrons. The structure of an atom is basically a cloud of electrons

surrounding a small, dense region, called the nucleus. An electron is a negatively charged particle with negligible mass.

The nucleus was discovered by Ernest Rutherford. Rutherford suggested that the charge and mass of an atom are concentrated in its nucleus, and that the nucleus has positive charge.

In current views, the nucleus of an atom contains protons and neutrons. Protons have positive charge, and neutrons have no charge. Each element has a unique number of protons in its atoms. This number is the atomic number of the element, and is designated by Z. The total number of protons and neutrons in the nucleus of an atom is the mass number of the element. Mass number is designated by *A*. Atomic number and mass number for an atom of an element, X, is designated by:

$$^{A}_{Z}X$$

For a neutral atom,

- the number of protons = the number of electrons = atomic number.
- the number of protons + the number of neutrons = mass number.

Over centuries, scientists have developed these main atomic models:

- Dalton's Model: an atom is a solid, indestructible sphere.
- Thomson's Model: an atom is a sphere with a uniform positive charge. Negatively charged electrons are embedded in the positive charge.
- Rutherford's Model: Rutherford discovered that atoms are mostly empty space with a dense, positive nucleus.
- Bohr's Model: in an atom, electrons follow circular orbits around the nucleus. The Bohr Model was the first quantum mechanical model of the atom.
- The Quantum Mechanical Model: in an atom, the orbits of electrons around the nucleus are elliptical. This theory is a modification of the Bohr model.

The quantum mechanical description of an atom is the electron configuration of the atom – the arrangement of electrons in the atom. The maximum number of electrons in a main energy level is $2n^2$, where n is the principal quantum number.

The quantum mechanical model is governed by this principle:

The Aufbau principle is a guide to the order of filling energy levels and sublevels. According to this principle lower energy sub levels are completely filled before electrons enter higher energy sublevels.

REVIEW EXERCISE ON UNIT 1

Part I: True/False type questions

1. Identify whether each of the following statement is correct or incorrect. Give your reasons for considering a statement to be false.
   - a Atoms of the same element have the same number of protons.
   - b Atoms of different elements can have the same number of protons.
   - c Atoms of the same element have the same number of electrons.
   - d Atoms of the same element have the same number of neutrons.
   - e The *s*- sublevel contains 2 electrons.
   - f The *d*-sublevel contains 5 orbitals.
   - g The 4*s*-sublevel comes before the 3*d* sublevel.
   - h The fundamental particle not present in a hydrogen atom is the proton.
   - i Isotopic elements contain the same number of neutrons.
   - j An alpha particle is composed of He^{2+} ions.
   - k The mass of an electron is almost equal to the mass of a proton.

Part II: Write the missing words in your exercise book

2. The mass of an atom is mainly concentrated in _____.
3. The diameter of an atom is _____ times that of the nucleus.
4. The energy level in the third principal quantum number can accommodate _____electrons.

Part III: Give the appropriate answers for the following questions

5. Draw diagrammatic representation of the following atoms by using shell or energy level.

   a $^{9}_{4}Be$ b $^{19}_{9}F$ c $^{28}_{14}Si$ d $^{39}_{19}K$

6. Write the electron configuration of the elements: (use quantum mechanical model)

   a $^{32}_{16}S$ b $^{52}_{24}Cr$ c $^{56}_{26}Fe$ d $^{64}_{29}Cu$

7. For the elements in question number 6,
   - a give the valence electrons for each.
   - b give the number of neutrons for each.

8. In a certain atom, the last electron has a $3d^2$ electron configuration. On the basis of this:
   a write the electron configuration of the element using a noble gas as a core.
   b write the electron configuration using the main energy levels.
   c what is the atomic number of the atom?
   d what are the valence shell electrons of the atom?
   e how many protons does the atom have?
9. What is the relative atomic mass of an element whose isotopic composition is 90% of ^{20}X and 10% of ^{22}X?

Part IV: Short answer type questions

10. Comment on Dalton's atomic theory and the areas in which modifications were made.
11. John Dalton made a number of statements about atoms that are now known to be incorrect. Why do you think his atomic theory is still found in science text books?
12. If scientists had tried to repeat Thomson's experiment and found that they could not, would Thomson's conclusion still have been valid? Explain your answer.

UNIT 2



Periodic Classification of the Elements

Unit Outcomes

After completing this unit, you will be able to:

- *understand the periodic classification of the elements;*
- *develop skills in correlating the electron configuration of elements with the periodicity of the elements, and in predicting the trends of periodic properties of elements in the periodic table;*
- *appreciate the importance of classification in chemistry; and*
- *demonstrate scientific inquiry skills: observing, inferring, predicting, classifying, comparing and contrasting, making models, communicating, measuring, asking questions, interpreting illustrations, drawing conclusions, applying concepts, and problem solving.*

MAIN CONTENTS

2.1 Introduction

2.2 The Modern Periodic Table

2.3 Periodic properties in the Periodic Table

2.4 Advantages of periodic classification

– Unit Summary

– Review Exercises

Start-up Activity

How can you use a table of repeating events to predict the next events? Table 2.1 shows a familiar table of repeating properties. What is it?

Of course, it is a calendar, but it is a calendar with a difference – it is missing some information. You can determine what is missing and fill in the blanks.

1. Look at the calendar again. The calendar has columns of days, Sunday through Saturday. The calendar also has horizontal rows. They are its weeks.
2. Examine the information surrounding each empty spot. Can you tell what information is needed in each empty spot?
3. Fill in the missing information. For example:

Conclude and Apply

1. One day in column 3 is marked X, and a day in column 4 is marked Y. What dates belong to these positions? Discuss your answers in your group.
2. Column 5 does not have a name. What is the correct name of this column?
3. What dates are included in the third row of the table?
4. Assuming that the previous month had 30 days, what day would the 28^{th} of that month have been? What row of this table would it appear in?
5. How do you relate this periodicity with the periodic classification of elements?

Table 2.1 Periodicity in time.

Sun	Mon	Tues	Wed		Fri	Sat
				1	2	3
4	5	6	7	8	9	10
11	12	X	Y	15	16	17
18	19	20	21	22	23	24
25	26	27	28	29	30	31

2.1 INTRODUCTION

Competencies

By the end of this section, you will be able to:

- Describe periodicity.

Activity 2.1



Form a group and discuss the following concepts. Present your discussion to the rest of the class.

What was the basis for the early attempts in classifying the elements? Outline the contributions of some scientists you have read about in classfying the elements.

Before the beginning of the 18th century, it was easy to study and remember the properties of the elements because very few were known. However, in the middle of the 19th century, many more elements were discovered. Scientists then began to investigate possibilities for classifying the known elements in a simple and useful manner. After numerous attempts, the scientists were ultimately successful. They grouped elements with similar properties together. This arrangement is known as the classification of elements.

Early Attempts in classifying the elements

What is meant by the term periodicity? Can you give some periodic events in nature?

Early attempts to classify elements were based merely on atomic mass. Then scientists began to seek relationships between atomic mass and other properties of the elements.

i) **Dobereiner's Triads:** One of the first attempts to group similar elements was made in 1817 by the German chemist J. Dobereiner. He put together similar elements in group of three or triads. According to Dobereiner, when elements in a triad are arranged in the order of increasing atomic masses, the middle element had the average atomic mass of the other two elements. For example, because the atomic mass of bromine is nearly equal to the average atomic mass of chlorine and iodine, he considered these three elements to constitute a triad when arranged in this order: chlorine, bromine, iodine see Table 2.2.

Table 2.2 Dobereiner's Triads.

Triads	Atomic Masses
Chlorine	35.5
Bromine	80
Iodine	127
Average atomic masses of chlorine and iodine = $\frac{35.5+127}{2}$ = 81.25	

Reading Check

Do you think that this classification works for all elements?

ii) **Newlands's Law of Octaves:** In 1864, John Newlands, an English chemist, reported the law of octave, which is also known as the law of eight. He stated that when elements are arranged in increasing order of their atomic masses, every eighth element had similar properties to the first element.

Newlands first two octaves of eight elements are shown below:

Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl
K	Ca					

However, the law of octaves could not be applied beyond calcium.

Reading Check

With the aid of an encyclopedia, reference books or other resources, write a report on:

a J. Dobereiner and

b J. Newlands works in organizing the elements.

In your report include the merits and demerits of their works.

Exercise 2.1

1. Decide whether the principle of Dobereiner's triad can be applied in the following groups of three elements.

   a Be, Ca, Sr b Li, Na, K

2. Newlands stated that there was a periodic similarity in properties of every eighth element in his system. However, today we see that for periods 2 and 3, the similarity occurs in every ninth element. What is the reason? Explain.

2.2 The Modern Periodic Table

Competencies

By the end of this section, you will be able to:

- state Mendeleev's and modern periodic law;
- describe period and group;
- explain the relationship between the electronic configuration and the structure of the modern periodic table;
- describe the three classes of the elements in the modern periodic table;
- explain the four blocks of the elements as related to their electronic configuration in the modern periodic table;

- tell the block of an element from its electronic configuration;
- give group names for the main group elements;
- classify the periods into short, long, and incomplete periods;
- tell the number of groups and periods in the modern periodic table;
- tell the number of elements in each period;
- predict the period and group of an element from its atomic number; and
- tell the block and group of an element from its electronic configuration.

Activity 2.2



Form a group and perform the following tasks. Share your findings with the rest of the class.

1. Make a list of the symbols for the first eighteen elements. Beside each symbol, write its electronic configuration.
2. Draw sets of vertical boxes. In order of increasing atomic number, fill into each set the symbol of all elements having the same outer electron configurations. How many sets are there? Record your answer.
3. Draw sets of horizontal boxes. In order of increasing atomic number, fill into each set the symbols of all elements having the same number of shells. How many sets are there? How many elements are there in each set? Record your answers.
4. Do you see any regular patterns that you created in Steps 2 and 3?
5. Draw one complete table which shows all elements;
   a with the same number of outermost electrons in a vertical column
   b filling the same outer electron shell in a horizontal row.

Historical Note



Dimitri Mendeleev

While attempting to group the elements according to their chemical properties and atomic weights, Dimitri Mendeleev developed the periodic table and formulated the periodic law. Because his classification revealed recurring patterns (periods) in the elements, Mendeleev was able to leave spaces in his table for elements that he correctly predicted would be discovered.

2.2.1 The Periodic Law

A. Mendeleev's Periodic Law

In 1869, the Russian chemist Dimitri Mendeleev and the German scientist Luthar Meyer independently published periodic arrangements of the elements based on increasing atomic mass.

Mendeleev observed that, when elements are arranged according to increasing atomic mass, the chemical and physical properties of the elements recur at regular intervals. This periodic variation and the recurrence of the properties of the elements led to the formulation of the Periodic Law.

Mendeleev's periodic law states that the properties of the elements are periodic functions of their atomic masses. Only 63 elements were known when Mendeleev constructed his table in 1871. Mendeleev organized his table in columns, with each column containing elements that have similar chemical properties. Accordingly, elements in the same column gave family or groups of elements.

How did Mendeleev know where to leave gaps for undiscovered elements in his periodic table?

Mendeleev left blank spaces for the undiscovered elements and also predicted masses and other properties of these unknown elements almost correctly. For example, neither gallium nor germanium were discovered when Mendeleev constructed his periodic table. But he predicted the existence and properties of these unknown elements.

Mendeleev left two blank spaces for these two elements in the table, just under aluminium and silicon. He called these unknown elements 'eka-aluminium' and 'eka-silicon.' ('eka' means 'first'). What he meant by 'eka-aluminium' is "a currently known element (galium) following aluminium".

Later on, in 1874, the element Gallium (eka-aluminium in Mendeleev's system) was discovered. In 1886, the element Germanium (eka-silicon) was discovered. The observed properties of these elements were remarkably very close to those in Mendeleev's predictions. Table 2.3 shows the properties of eka-silicon predicted by Mendeleev and compares them to the observed properties of Germanium.

Table 2.3 Comparison of Mendeleev's predictions for the properties of Eka-silicon with Germanium.

Property	Mendeleev's Predictions for eka-silicon (Es) in 1871	Observed Properties for Germanium (Ge) in 1886
Atomic mass	72	72.6
Density (g/cm^3)	5.5	5.47
Colour	Dark Gray	Light Gray
Oxide formula	EsO_2	GeO_2
Density of oxide (g/cm^3)	4.7	4.7
Chloride formula	$EsCl_4$	$GeCl_4$
Density of chloride (g/cm^3) . . .	1.9	1.887
Boiling point of chloride	< 100°C	86°C

Defects in Mendeleev's periodic table

1. Position of isotopes: The isotopes were not given separate places in Mendeleev's periodic table. Since elements are arranged in order of increasing atomic masses, the isotopes belong to different groups (because *isotopes have different masses*).
2. Wrong order of atomic masses of some elements: When certain elements are grouped on the basis of their chemical properties, some elements with higher atomic masses precede those with lower atomic masses. For example, argon, with atomic mass of 39.95, precedes potassium with atomic mass of 39.1.

B The Modern Periodic Law

What was Mosley's contribution to the modern form of the Periodic Table?

In 1913, the English physicist Henry Mosley determined the atomic number of each of the elements by analyzing their X-ray spectra. He observed that when each element was used as a target in an X-ray tube, it gave out X-rays with a characteristic wavelength. The wavelength depends on the number of protons in the nucleus of the atom and was constant for a given element. By arranging the elements in order of decreasing wavelength, Mosley was able to assign atomic number to each element. The atomic number of every element is fixed, and it clearly distinguishes one element from another. '*No two elements can have the same atomic number*.' For example, atomic number 8 identifies the element oxygen. No other element can have atomic number 8.

Therefore, the atomic number of an element is the fundamental property that determines the chemical behavior of the element. The discovery of atomic number led to the development of the modern periodic law. The modern periodic law states that: ‘‘the properties of the elements are periodic function of their atomic numbers.’’ This means that when elements are arranged according to increasing atomic number, elements with similar physical and chemical properties fall in the same group.

2.2.2 Characteristics of Groups and Periods

Many different forms of the periodic table have been published since Mendeleev's time. Today, the long form of the periodic table, which is called the modern periodic table, is commonly in use. It is based on the modern periodic law. In the modern periodic table, elements are arranged in periods and groups.

What are the basis for classifying the elements into groups and periods?

What are the similarities and differences in the electron configuration of* S *and* Cl*?

Periods: The horizontal rows of elements in the periodic table are called periods or series.

Elements in a period are arranged in increasing order of their atomic numbers from left to right.

There are 7 periods in the modern periodic table, and each period is represented by an Arabic numeral: 1, 2 . . . and 7.

- Elements in the same period have the same number of shells.
- Periods 1, 2, and 3 are called short periods while periods 4, 5, and 6 are known as long periods.
- Period 1 contains only 2 elements, hydrogen and helium. Period 2 and period 3 contain 8 elements each.
- Period 4 and period 5 contain 18 elements each. Period 6, the longest period, has 32 elements. Period 7, which is an incomplete period, contains more than 24 elements.Period 7 element is radioactive and/or an artificial element.
- Except for the first period, all periods start with an alkali metal and ends with a noble gas.

Table 2.4 The number of elements in a given period and the orbitals being filled.

Period number	Orbitals occupied	Number of elements
1	1*s*	2
2	2*s*, 2*p*	8
3	3*s*, 3*p*	8
4	4*s*, 3*d*, 4*p*	18
		and so on

The position of an element in a given period can be determined by the number of shells occupied with its electrons. Accordingly, the number of shell is equal to the number of period to which the element belongs.

Example

Electronic configuration of $^{23}_{11}Na$ = $1s^2 2s^2 2p^6 3s^1$ (2, 8, 1). Sodium has 3 main shells. Hence, sodium is found in period 3.

***What are the similarities and differences in the electron configuration of* F *and* Cl?**

Groups or families: are the vertical columns of elements in the periodic table. There are 18 columns or groups in the modern periodic table.

- Group numbers are usually designated with the Roman numerals I to VIII each followed by the letter A or B.

  These are: IA . . . . . VIII A Main groups (A groups)

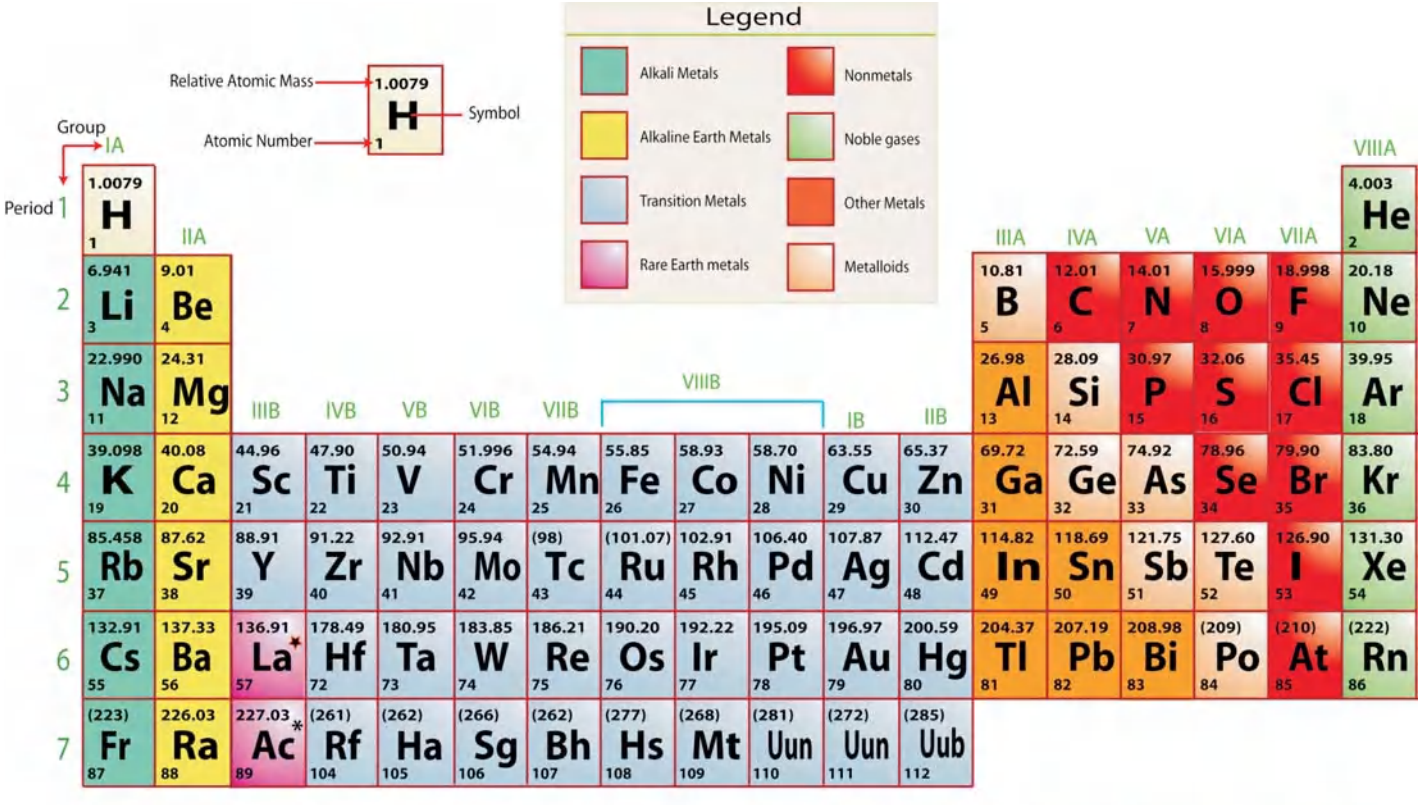
  IB . . . . . VIIIB Sub groups (B groups)

- Elements in a given group have the same number of outermost shell electrons.
- Elements in the same group have similar chemical properties.
- For the main group elements the group number equals the number of valence electrons.

Example

Electronic configuration of $^{35}_{17}Cl$ = $1s^2 2s^2 2p^6 3s^2 3p^5$ (2, 8, 7). The number of valence electrons of chlorine is 7. Hence chlorine is found in Group VIIA.

PERIODIC TABLE



Exercise 2.2

Give appropriate answers for the following questions.

1. Which group of elements was missing from Mendeleev's periodic table as compared to the modern periodic table?
2. Why does the first period contain only two elements?
3. To which group and period do the following elements belong?

   a carbon
   b neon
   c aluminium
   d potassium
   e calcium
   f sulphur

2.2.3 CLASSIFICATION OF THE ELEMENTS

Activity 2.3



Perform the following tasks in groups and present your conclusion to the class.

For the following elements, determine the valence electrons, and identify the sub–shell (*s*, *p*, *d* or *f*) in which the last electron of each element enters.

a Nitrogen (atomic number = 7)
b Sodium (atomic number = 11)
c Silicon (atomic number = 14)
d Iron (atomic number = 26)
e Zinc (atomic number = 30)
f Krypton (atomic number = 36)
g Cerium (atomic number = 58)

Elements in a periodic table can be classified into three distinct categories based on their electron configuration and the type of sub-level being filled. These are the representative elements, the transition elements, and the rare-earth elements.

1. Representative Elements: *s*-and *p*-block elements

These are elements in which valence electrons are filling the *s*- or *p*-orbitals. Representative elements are also known as Main Group Elements. They include elements in groups IA through VIIIA.

s-block elements are elements in which the last electron enters the *s*-orbital of the outermost shell. Their general valence shell configuration is ns^1 and ns^2, where *n* represents the outermost shell. They are found on the left side of the periodic table

and contain the first two groups: Group IA and Group IIA. These two groups contain very reactive metals.

p-block Elements are elements in which the last electron enters the *p*-orbital of the outermost shell. They have the general valence shell electron configuration of ns^2np^1–ns^2np^6. *p*-block elements are located at the right hand side of the periodic table and they include six groups-Group IIIA to VIIIA. Most of these elements are non-metals.

Table 2.5 Valence electron configuration, blocks and common names for the representative elements.

Group Number	Common name	General valence shell electron configuration		Number of valence electrons
IA	Alkali metals	ns^1	*s*-block	1
IIA	Alkaline earth metals	ns^2	*s*-block	2
IIIA	Boron family	ns^2np^1	*p*-block	3
IVA	Carbon family	ns^2np^2	*p*-block	4
VA	Nitrogen family	ns^2np^3	*p*-block	5
VIA	Oxygen family (Chalcogen)	ns^2np^4	*p*-block	6
VIIA	Halogens	ns^2np^5	*p*-block	7
VIIIA	Noble gases (Inert gases)	ns^2np^6	*p*-block except He	8

Why is helium classified as a noble gas?

2. Transition Elements: *d*-block elements

These elements are found in the periodic table between the s-block and p-block elements. In these elements, the valence electrons are being added to the *d*-orbital of the outermost shell. They are also known as *d*-block elements. They are designated as 'B' group, and they consist of groups IB – VIIIB. The transition elements are found in periods 4, 5, 6, and 7. They include common elements such as iron, gold, and copper. *d*-block elements are also called transition metals.

3. Rare-Earth Elements: *f*-block elements

These elements are also known as inner-transition elements. They are elements in which the last electrons are being added to the *f*-orbital of the outermost shell. The inner transition elements are in periods 6 and 7. The period 6 inner transition series fills the 4*f*-orbitals. It is known as the Lanthanide series because it occurs after the element lanthanum and its elements are similar to it. The period 7 inner transition series fills the 5*f*-orbitals and appears after the element actinium. Its elements are similar to actinium, thus it is known as the Actinide series.

Figure 2.1 presents a scheme to show the basic structure of the periodic table. On the basis of the electron configuration of the elements the table is divided into *s*, *p*, *d* and *f*-blocks.

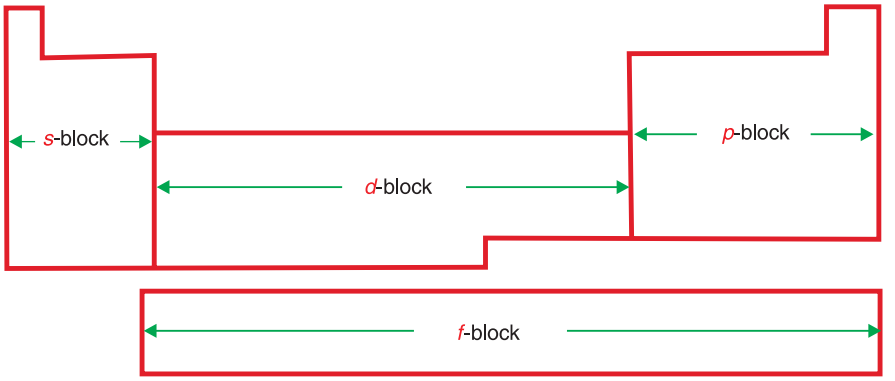


Figure 2.1 Division of the modern periodic table.

Reading Check

In the periodic table, you have seen a section called rare earth elements. Find this term in an encyclopedia, in reference books, or in other resource material. Read and analyze the information you find. Write a report on these elements.

Exercise 2.3

Give appropriate answers for the following questions:

1. For the following elements, write their electron configurations and determine the group and period number of each element.
   a Na, Ca, Al b Cl, S, Ar
2. Bromine is a Group VIIA and period-4 element. What is the valence shell configuration of bromine?
3. Deduce the group, period, and block of the elements with atomic numbers:
   a 37 b 24 c 32

Critical Thinking

4. Why is hydrogen, a non-metal, usually placed with Group I elements in the periodic table, even though it does not show a metallic property like the alkali metals?

2.3 PERIODIC PROPERTIES IN THE PERIODIC TABLE

Competencies

By the end of this section, you will be able to:

- explain the general trends in properties of the elements as they move down a group of the periodic table;

- explain the general trends in properties of elements across a period;
- deduce the properties of an element from its position in the periodic table; and
- make a chart to show the trends in properties of elements in the periodic table.

Activity 2.4



Refer any chemistry book and look up values for atomic size, ionization energy, electronegativity and other properties of elements. Enter the values in the periodic table, and see if there any obvious trends in the properties as you go across a period or down a column of the table. Report your findings to the class.

In the periodic table the properties of the elements such as atomic size, ionization energy, electron affinity, and electronegativity show a regular variation with in a group or across a period.

2.3.1 Periodic Properties within a Group

Elements in the same group have the same number of valence electrons and also exhibit similar chemical properties. For example, group IA elements have 1 valence electron; those in Group IIA have 2 valence electrons, and so on. Generally, it is possible to conclude that the number of valence electrons determines the group number of an element.

Table 2.6 Electron configuration and number of valence electrons of Group IA elements.

Element	Electron Configuration	Number of Valence electron	Group Number
Li	2, 1	1	
Na	2, 8, 1	1	
K	2, 8, 8, 1	1	IA
Rb	2, 8, 18, 8, 1	1	
Cs	2, 8, 18, 18, 8, 1	1	

The periodic properties of the elements can be explained on the basis of nuclear charge and effective nuclear charge.

Nuclear charge (Z): is the total positive charge in the nucleus of an atom.

Effective Nuclear charge (Z_{eff}): In an atom, the outermost shell electrons (or valence electrons) are attracted to the nucleus and simultaneously repelled by the inner shell electrons. The attraction of the nucleus for the valence electrons is also reduced because inner electrons shield (or screen) the valence electrons. As a result, these

inner electrons reduce the attraction of the nuclear charge. The resulting net-positive nuclear charge attracting the valence electrons is called effective nuclear charge, Z_{eff}.

Effective nuclear charge relates the nuclear charge to the number of shells (size) of the atom.

The effective nuclear charge is the difference between the nuclear charge (Z) and the inner electrons (S) that shield the valence electrons. The effective nuclear charge is always less than the actual nuclear charge.

$$Z_{eff} = Z - S$$

Let us consider the sodium atom with its 11 electrons: $1s^2 2s^2 2p^6 3s^1$ (2, 8, 1). The 10 inner electrons of sodium completely cancel the 10 units of nuclear charge on the nucleus. In this way, the inner electrons shield the valence electrons from the full attractive force of the nucleus and leave an effective nuclear charge of + 1 (Z_{eff} = + 1).

The fact that the inner electrons shield or screen the outer electrons from the full charge of the nucleus is known as shielding or screening effect.

Activity 2.5



Form a group and perform the following task:

Draw Bohr's model for the following elements and indicate the shielding shells, shielding electrons and effective nuclear charge:

a beryllium, magnesium and calcium;

b lithium, carbon and fluorine.

Present your findings to the class.

The following properties of elements vary in a regular periodic manner.

1. Nuclear Charge

On moving down a given group of the periodic table, nuclear charge progressively increases, but effective nuclear charge remains nearly constant.

2. Atomic Size/Atomic Radius

Where in a group do you find atoms with the largest atomic radius? Why?

It is difficult to measure the size of an atom directly. The electron cloud enveloping the nucleus does not have a clear boundary because its electrons do not have fixed distances from the nucleus. Therefore, atoms do not have definite outer boundaries.

The size of an atom is defined in terms of its atomic radius. For metals, atomic radius is defined as one-half the distance between the nuclei of the two adjacent atoms. For

elements that exist as diatomic molecules (such as chlorine), atomic radius is equal to one-half of the distance between the nuclei of the atoms in the molecule. Figure 2.2 illustrates the atomic radius of chlorine in a chlorine molecule.

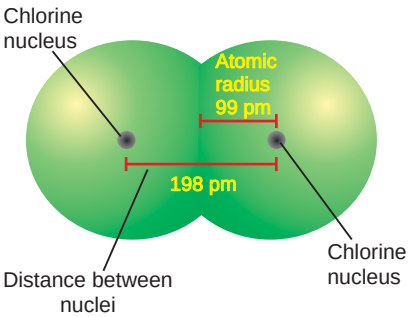


Figure 2.2 The representation of atomic radius in a chlorine molecule.

In moving down a group, atomic radius of the elements mainly depends on the number of shells.

Activity 2.6



Form a group and perform the following task; present your findings to the class.

The following values are given for atomic radii (in Å) of group IA elements:

1.54, 1.34, 2.35, 2.16 and 1.96

Based on the given information;

1. fill the following table with the number of shells and the appropriate values for atomic radii corresponding to the symbols of the elements.
2. explain the reason for the observed trend.

Elements	Number of Shells	Atomic Radius (Å)
Li Na K Rb Cs		

3. Ionization Energy

In which region of the periodic table do you find elements with the:

a lowest tendency to lose electrons, and

b highest tendency to lose electrons?

Ionization energy is the minimum energy required to remove the outermost shell electron from an isolated gaseous atom or ion.

Ionization energy is represented by the following equation (*where* M *denotes any metal*).

$$M\ (g) + \text{energy} \longrightarrow M^{+}\ (g) + e^{-}$$

The electrons in an atom can be successively removed, one after another. Thus, the first ionization energy is the energy needed to remove the first valence electron, the second ionization energy is the energy needed to remove the second valence electron, and so on.

$M\ (g) + \text{energy} \longrightarrow M^{+}\ (g) + e^{-}$ $\Rightarrow$ *First ionization energy*

$M^{+}\ (g) + \text{energy} \longrightarrow M^{2+}\ (g) + e^{-}$ $\Rightarrow$ *Second ionization energy*

For a given element, the second ionization energy is higher than the first one.

Ionization energy is always a positive value (and therefore is an endothermic process) because energy is required to remove an electron from an atom. Ionization energy is measured in electron volts (eV) or kiloJoules per mole (kJ/mol).

Ionization energy is a measure of the tendency of an atom to lose an electron. Metals easily lose electrons and thus have low ionization energy. Non-metals have high ionization energy because they do not easily lose electrons.

Activity 2.7



Form a group and perform the following task. Present your findings to the class.

Rank each set of the following elements in order of decreasing ionization energy and explain the trend in ionization energy of the elements down a group.

i) Ca, Sr, Mg, Be

ii) K, Li, Rb, Na

iii) Cl, F, I, Br

Generally, ionization energy is affected by the following factors:

i) Atomic size: As atomic size increases, the valence electrons are less tightly held by the nucleus. Thus, less energy is required to remove these electrons. For example, the energy needed to remove an electron from a cesium atom is lower than from a lithium atom.

ii) Effective nuclear charge: The smaller the effective nuclear charge of an atom, the lower is the energy needed to remove an electron from the atom.

iii) Types of electrons: The closer an electron is to the nucleus, the more difficult it is to remove the electron. In a given energy level, *s*-electrons are closer to

the nucleus than *p*-electrons. Similarly, *p*-electrons are closer than *d*-electrons, and *d*-electrons are closer than *f*-electrons. Hence, ionization energy decreases in the order of: $s > p > d > f$.

iv) Screening effect by the inner electrons: As described earlier, inner shell electrons shield the valence electrons from the nuclear charge. The more inner electrons there are, the higher the screening effect, and therefore the easier it is to remove the valence electrons. Screening decreases ionization energy.

v) Electron configuration (stability): It is easier to remove electrons from unstable sublevels than from stable ones. Half-filled (p^3, d^5, f^7) and completely-filled (d^{10}, p^6, f^{14}) sublevels are more stable. For example, more energy is required to remove a p^3 electron than a p^4 electron. As a result of it, the first ionization energy of nitrogen is higher than that of oxygen.

Generally, with the increasing atomic number, the first ionization energy decreases down the same group.

Activity 2.8



After studying the given information, identify the elements represented by X, M, and Y; discuss the findings in your group, and then share your conclusion with the other groups.

Element X

- has a relatively high ionization energy.
- generally forms an ion with a – 2 charge.
- has an outermost electron configuration of $3s^2 3p^4$.

Element M

- reacts with oxygen to form M_2O.
- has a very low ionization energy.
- is in the fourth period of the periodic table.

Element Y

- is a transition element.
- is used in Ethiopian coinage.
- has 10 electrons in the 3d orbital

4. Electron Affinity (E_A)

To which group it is easier to add electrons; alkali metals or halogens?

Non-metals gain electrons and therefore form negative ions. The tendency of an atom to form a negative ion is expressed in terms of electron affinity (E_A).

Electron affinity is defined as the energy released in kilojoules/mole, when an electron is added to an isolated gaseous atom to form a gaseous ion. It is a measure of the attraction or 'affinity' of the atom for the extra added electron.

$$X(g) + e^- \longrightarrow X^- (g) + \text{energy}$$

Since energy is liberated during the process, electron affinity is expressed as a negative value. For example, when an electron is added to a fluorine (F) atom, 328 kJ/mol of energy is released to produce a fluoride ion (F^-), and E_A is – 328 kJ/mol.

$$F(g) + e^- \longrightarrow F^- (g) + E_A = -328 \text{ kJ/mol}$$

Electron affinity is a measure of the strength of an atom to attract an additional electron. The smaller is the atomic size of an element, the stronger is the tendency to form negative ions, and consequently the higher the electron affinity. Generally, electron affinity depends on atomic size and effective nuclear charge of the elements.

Activity 2.9



Form a group and discuss the following concepts.

1. Why does the electron affinity of Cl is higher than that of F?
2. Explain why noble gases have extremely low (almost zero) electron affinities?
3. Explain why halogens have the highest electron affinities?

Present your findings to the class.

5. Electronegativity

Where do you find the most electronegative element in the periodic table?

Electronegativity is the ability of an atom in a molecule to attract the shared electrons in the chemical bond. The American chemist Linus Pauling (1901-1994) developed the most widely used scale of electronegativity values based on bond strength. The Pauling scale ranges from 0.7 to 4.0. Fluorine, the most electronegative element, is assigned a value of 4.0, and the least electronegative element, cesium, has an electronegativity value of 0.7. The electronegativity values for all the rest elements lie between these extremes.

The electronegativity of an atom is related to its ionization energy and electron affinity.

An atom with high ionization energy and high electron affinity also tends to have a high electronegativity value because of its strong attraction for electrons in a chemical bond.

Activity 2.10



Form a group and perform the following task:

Arrange each set of the given elements in order of decreasing electronegativity and explain the observed trend.

a Ba, Mg, Be, Ca

b C, Pb, Ge, Si

c Cl, F, I , Br

Present your findings to the class.

6. Metallic Character

In which region of the periodic table do you find metals and non-metals?

Metals have the tendency to lose electrons and form positive ions. As a result, metals are called electropositive elements.

In moving down a group, atomic size increases progressively, and it becomes easier for elements to lose their valence electrons and form positive ions. Therefore, metallic character increases down a group.

In the periodic table, metals and non-metals are separated by a stair step diagonal line, and elements near this border line are called metalloids. Metals are found on the left side of the line and nonmetals on its right side.

Activity 2.11



Form a group, perform the following task and present your findings to the class:

1. The members of group IVA of the periodic table are: Ge, Sn, C, Pb and Si.

   Classify these elements into metals, non-metals and metalloids.

2. Explain the differences between silicon and lead in terms of:

   a Atomic size

   b Ionization energy

   c Electron affinity

   d Electronegativity

Exercise 2.4

Part I: *Choose the correct answer from the given choices*

1. Which of the following properties of the elements remain unchanged down a group?

   a Ionization energy c Electron affinity

   b Nuclear charge d Valence electrons

2. Which of the following elements has the largest atomic size?

   a Be c Ca

   b Ba d Mg

3. Which of the following elements has the lowest electronegativty?

   a F b Br c I d Cl

Part II: *Give short answers for the following*

4. What is screening effect? How does it relate to effective nuclear charge?
5. What is the relationship between first ionization energy and the metallic properties of elements?

2.3.2 Periodic Properties within a Period

Activity 2.12



Form a group and draw a rough sketch of the periodic table (*no details are required*). Based on this, perform the following activities and present your findings to the class.

1. Indicate the regions where metals, non-metals, and metalloids are located in the periodic table. (*Use different colours*).
2. Where do you find the most active metal and the most active non-metal? .

As we move from left to right across a period, the number of valence electrons of the elements increases. But the number of energy levels or main shells remains the same in a given period and electrons are filling the same energy level until stable noble gas configuration is achieved. The additional electrons from Li (2, 1) to Ne (2, 8) are added to the second shell. In fact, the period number equals the number of energy level being filled.

Table 2.7 Electron configuration and number of shells for period-2 elements.

Elements	Li	Be	B	C	N	O	F	Ne
Atomic number	3	4	5	6	7	8	9	10
Electron configuration	2, 1	2, 2	2, 3	2, 4	2, 5	2, 6	2, 7	2, 8
Number of shells (period number)	2	2	2	2	2	2	2	2

Let us now consider ***periodic properties of elements across a period.***

1. **Atomic Size:** From left to right in a given period, nuclear charge or atomic number progressively increases by one for every succeeding element. However, increasing number of valence electrons is being added to the same shell. This results in an increase in effective nuclear charge.

Activity 2.13



Form a group and perform the following task. Present your findings to the class.

The following values are given for atomic radii (in Å) of period 2 elements:

0.77, 1.34, 0.69, 0.75, 0.73, 0.82, 0.90, and 0.71.

Based on the information given;

1. fill the table given below with the appropriate values for atomic radii of the elements.
2. explain the reason for the observed trend.

Element	Li	Be	B	C	N	O	F	Ne
Atomic radii (Å)								

2. **Ionization Energy:** Two factors account for the general trend in ionization energy across a given period. These are nuclear charge and atomic size.

Activity 2.14



Form a group and perform the following task, and present your findings to the class.

The following table lists ionization energy values of period 3 elements

Elements	Na	Mg	Al	Si	P	S	Cl	Ar
Ionization energy (kJ/mol)	496	738	578	787	1012	1000	1251	1521

1. Explain why there is a general increase in the first ionization energy across the period from Na to Ar.
2. Explain why the first ionization energy of:
   i) aluminium is lower than that of magnesium.
   ii) sulphur is lower than that of phosphorus.
   (Hint: Use sublevel configurations for the elements).

3. **Electron Affinity:** The variation in electron affinity of elements in the same period is due to changes in nuclear charge and atomic size of the elements.

   A cross a period from left to right, electron affinity increases due to an increase in effective nuclear charge. The elements show a greater attraction for an extra added electron.

4. **Electronegativity:** Across a period, a gradual change in nuclear charge and atomic size determine the trends in the electronegativity of the elements.

Activity 2.15



Form a group and perform the following task. Present your findings to the class.

The following values are given for electronegativity of period 3 elements:

2.1, 0.9, 1.5, 3.0, 1.8, 2.5 and 1.2

Based on the information given;

1. Draw a table of period 3 elements and fill with the appropriate electronegativity values corresponding to the symbols of the elements.
2. Explain the reason for the observed trend.

5. **Metallic character:** From left to right in a period, metallic character of the elements decreases. Elements on the left end of a period have a higher tendency to form positive ions. Those at the right end have a greater tendency to form negative ions. In any period, elements on the left side are metals and those on the right side are nonmetals.

   Consider the following period-3 elements and observe how the elements become more non-metallic on moving from left to right in the periodic table:

Elements	Na	Mg	Al	Si	P	S	Cl	Ar
Character	Metals			Metalloid	Non-metals			

Metallic character decreases

Project work

Model of a periodic table

The following table shows a section of a periodic table. It is incomplete. According to the data given in table 2.8,

a Construct a model of this section of the periodic table using locally available materials to show the trends in atomic radii of the elements.

b Complete the section of the periodic table with the elements in the appropriate position.

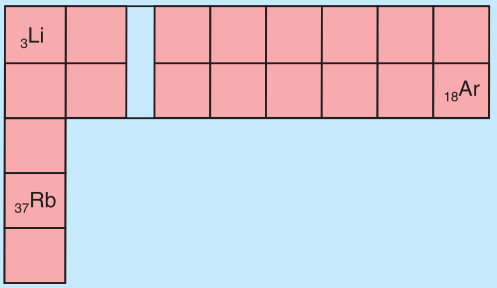


Table 2.8 Atomic number and atomic radii of some selected elements.

Element	Atomic number	Atomic radii (nm)	Element	Atomic number	Atomic radii (nm)
Li	3	0.123	Al	13	0.125
Be	4	0.089	Si	14	0.117
B	5	0.080	P	15	0.110
C	6	0.077	S	16	0.104
N	7	0.074	Cl	17	0.099
O	8	0.074	Ar	18	0.099
F	9	0.072	K	19	0.203
Ne	10	-	Rb	37	0.216
Na	11	0.157	Cs	55	0.235
Mg	12	0.136			

(*Hint: Be innovative. Use cheap local material available to you. Your teacher will help you whenever necessary.*)

Questions:

Give reasons for the increase or decrease in atomic radii that you observe in your model:

i from lithium to fluorine and sodium to chlorine

ii from lithium to cesium

Present your findings to your class teacher.

Exercise 2.5

Give appropriate answers for the following questions.

1. The element with the electron configuration of $1s^2\ 2s^2 2p^6\ 3s^2 3p^6$ is a:
   a metal b non-metal c metalloid d noble gas
2. Which of the main groups in the periodic table has the elements with the most negative value of electron affinity?
   a halogens b noble gases c Alkali metals d none
3. Which of the following is the correct increasing order of atomic size for the elements: Al, Ar, Na, Si?
   a Si, Na, Al, Ar c Ar, Si, Al, Na
   b Na, Al, Si, Ar d Al, Si, Ar, Na

Critical Thinking

4. Explain why the first ionization energy of nitrogen is the highest as compared to that of carbon and oxygen.
5. Compare the elements fluorine and chlorine with respect to their electronegativity and electron affinity values.

2.4 ADVANTAGES OF PERIODIC CLASSIFICATION

Competencies

By the end of this section, you will be able to:

- describe the advantages of periodic classification in studying chemistry.

Activity 2.16



Form a group and discuss the importance of the periodic table for predicting the atomic size and ionization energy of the elements. Share your idea with the class.

Why do we need the classification of elements?

The main advantages of using the periodic table are:

1. The periodic table is useful for predicting the formulas of compounds. The elements in a given group form compounds with the same atomic ratio

because of their similar electron configuration. For example, if the chemical formula of sodium oxide is Na_2O, then we can predict the formulas of the other oxides of alkali metals. These are Li_2O, K_2O, Rb_2O, and Cs_2O.

2. The periodic table is useful for predicting the physical and chemical properties of elements. For example, radium is a rare and radioactive element and therefore difficult to handle in many experiments. Since its properties can be predicted from the general trends of group IIA elements, sometimes we do not need to analyze it directly.
3. The periodic table is also useful for predicting the behaviour of many compounds. For example, oxides of the elements become more acidic across a period and more basic in character down a group. The trends in the oxides of period-3 elements vary from strongly basic oxides to amphoteric and then acidic oxides as we move across a period.

Table 2.9 Oxides of period 3 elements.

Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7
Basic oxide		Amphoteric oxide	Acidic oxide			

For a given element, the important information indicated below, in i, ii and iii can be read, deduced or stated from the periodic table.

(*i*) Read the
- Name
- Symbol
- Atomic number
- Atomic mass

(*ii*) Deduce the
- Number of protons and electrons
- Electron configuration (number of shells and of valence electrons)
- Character (behavior) as metal, non metal, or metalloid
- Nature (property) of the oxides to form acids, bases ...

Based on the information in i and ii, we can:

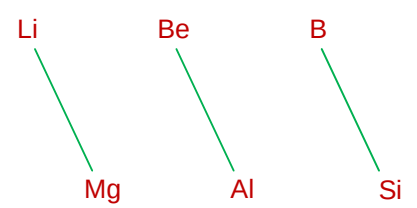
(*iii*) State the
- Period number
- Group number and
- Block type (*s*, *p*, *d* or *f*)

Diagonal Relationship

How do you account for the unexpected resemblance of the properties of the following sets of elements: Li and Mg, Be and Al, and B and Si?

In addition to the group and period relationships, the elements of s and *p* block also exhibit diagonal relationship. On moving diagonally across the periodic table, the elements show certain similarities. Though clear upon examination, these diagonal relationships are far less pronounced than the similarities within a group.

Diagonal relationship is particularly noticeable in the elements of second and third periods of the periodic table. The following illustrations show the diagonal relationship between Li and Mg, Be and Al, and B and Si.



Check List

Key terms of the unit

- *Actinide series*
- *Atomic size*
- *Block of elements*
- *Dobereiner's traids*
- *Effective nuclear charge*
- *Electron affinity*
- *Electron configuration*
- *Electronegativity*
- *Groups*
- *Inner transition elements*
- *Ionization energy*
- *Lanthanide series*
- *Law of octave*
- *Mendeleev's law*
- *Metalloids*
- *Metals*
- *Non-metals*
- *Nuclear charge*
- *Periodic law*
- *Periodic table*
- *Periods*
- *Representative elements*
- *Shielding (screening) effect*
- *Transition elements*

Unit Summary

- Early attempts of classification of the elements were made by Dobereiner and Newlands.
- Mendeleev's law states that properties of the elements are periodic functions of their atomic masses.
- Mendeleev arranged the elements based on increasing atomic mass.
- Modern periodic law states that the properties of the elements are periodic function of their atomic numbers.
- Periods are horizontal rows, and groups are vertical columns of the elements in the periodic table.
- Elements in the same group show similar chemical properties.
- Elements are classified as representative, transition, and rare-earth elements. This classification is based on the type of sub-level (*s*, *p*, *d*, or *f*) being filled.
- Ionization energy is the energy required to remove the outermost shell electron from an isolated gaseous atom.
- Electron affinity of an element is the energy released when an electron is added to an isolated gaseous atom to form a gaseous ion.
- Metallic character is the tendency to lose electrons and form positively charged ions.
- Electronegativity of an element is its ability to attract electrons.
- Trends in atomic size determine the trends in ionization energy, electron affinity, electronegativity, and metallic character of the elements in the periodic table.
- Atomic size itself is determined by the number of energy levels, nuclear charge, and effective nuclear charge.
- Periodic properties of elements, such as ionization energy, electron affinity, electronegativity, etc. show regular variation within a group or period.

REVIEW EXERCISE ON UNIT 2

Part I: Identify whether each of the following statements is true or false. Give your reasons when you consider a statement to be false.

1. Metallic properties of the elements increase from left to right within a period.
2. Elements in a group have consecutive atomic numbers.
3. All elements with high ionization energy also have high electron affinity.

4. All the elements that belong to s and p-blocks are metals.
5. The modern periodic law was proposed by Mosley.
6. As the atomic number of elements increases in the periodic table, their atomic radius also increases.
7. Transition metals are found in four periods. Each corresponds to the filling of valence electrons in the 3*d*, 4*d*, 5*d*, and 6*d* orbitals.

Part II: Multiple Choices Type Questions

8. The element with atomic number 35 belongs to:
   a s-block c d-block
   b p-block d f-block
9. An element with the valence electron configuration $3s^2 3p^2$ belongs to group:
   a IIIA c VIA
   b IVA d VIIA
10. An element with the valence electron configuration $3s^2 3p^6$ belongs to period:
    a 2 c 8
    b 6 d 3
11. An element with the valence electron configuration $2s^2 2p^4$ belongs to:
    a alkali metals c halogens
    b oxygen group d noble gases
12. The property of the element with atomic number 18 resembles that of the element with atomic number:
    a 8 b 36 c 19 d 40
13. An element with high ionization energy and high electron affinity also tends to have a high value of:
    a electronegativity c atomic size
    b electropositivity d metallic character

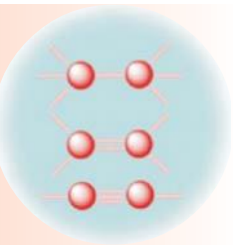
Part III: To which group, period, and sublevel block do the following elements belong?

		Group	Period	Block
14.	Calcium	_____	_____	_____
15.	Sulphur	_____	_____	_____
16.	Iron	_____	_____	_____
17.	Zinc	_____	_____	_____
18.	Xenon	_____	_____	_____

Part IV: Give short answers

19. Explain, with examples, the meaning of diagonal relationship.
20. From the given elements : Na, P, Ca, and Br; which one has the:
    a highest first ionization energy?
    b smallest atomic size?
    c most metallic character?
21. On the basis of electron configuration, explain why nitrogen is less electronegative than oxygen?
22. On the basis of electron configuration, explain why atomic size of sodium is larger than that of chlorine?

UNIT 3



Chemical Bonding and Intermolecular Forces

Unit Outcomes

After completing this unit, you will be able to:

- *discuss the formation of ionic, covalent, and metallic bonds;*
- *know the general properties of substances containing ionic, covalent, and metallic bonds;*
- *develop the skills of drawing the electron-dot (Lewis structures) for simple ionic and covalent compounds;*
- *understand the origin of polarity within molecules;*
- *appreciate the importance of intermolecular forces in plant and animal life;*
- *understand the formation and nature of intermolecular forces; and*
- *demonstrate scientific inquiry skills: observing, predicting, making model, communicating, asking questions, measuring, applying concepts, comparing and contrasting, relating cause and effects.*

MAIN CONTENTS

3.1 Chemical bonding
3.2 Ionic bonding
3.3 Covalent bonding
3.4 Metallic bonding
3.5 Inter-molecular forces
– Unit Summary
– Review Exercises

Start-up Activity

Objective of the Activity

Scientists have identified different types of attractive forces between atoms in forming bonds. The strength of the forces relies on the types of bonds. For instance, in covalent bonds, the strength of the bonds depends on whether the bonds are single, double or triple bonds.

In this activity, you will use bundles of sticks to develop your ideas about strength of bonds.

1. Collect the following materials and bring to school:

   – Six sticks of wood of the same length and of the same thickness.

2. Place your sticks on the table in the classroom,

   Your teacher will place your sticks in three groups.

   – A single stick

   – Pairs of sticks

   – Sets of three sticks

   Your teacher will assign three students and will give for each student a single stick, a pair of sticks, and a set of three sticks.



Figure 3.1 A bundle of sticks.



Figure 3.2 Students trying to break the sticks.

Analysis

1. Which of your bundles of sticks was the strongest? What is the reason for the different strengths of the bundle of sticks?
2. Draw your conclusions and present to the rest of the class.

3.1 CHEMICAL BONDING

Competencies

By the end of this section, you will be able to:

- define chemical bonding;
- explain why atoms form chemical bonds.

Activity 3.1



Most of the elements are not found free in nature. Why do the elements exists in combined form ? Discuss in group and present your findings to the class.

A chemical bond is the attractive force that binds atoms together in a molecule, or a crystal lattice.

After the periodic table and the concept of electron configuration were developed, scientists began to develop ideas about molecules and compounds. In 1916, G.N Lewis concluded that atoms combine in order to achieve a more stable electron configuration resulting in molecules or compounds.

Historical Note



G.N. Lewis

Gilbert Newton Lewis (1875 - 1946), introduced the theory of the shared-electron-pair chemical-bond formation in a paper published in the journal of the American chemical society. In honor of this contribution, we sometimes refer to “electron-dot” structures as “Lewis structures”.

As independent particles, atoms are at relatively high potential energy. Nature, favors arrangements in which potential energy is minimized. Most individual atoms exist in a less stable state than in their combined form.

When atoms form bonds with each other, they attain lower potential energy states. This decrease in atomic energy generally results in a more stable arrangement of matter. When atoms interact to form a chemical bond, only their outer regions are in contact. In the process of the interaction, atoms achieve stable outermost shell configuration. For this reason, when we study chemical bonding, we are concerned primarily with the valence electrons. As described in Unit 2, valence electrons are electrons that exist in the outermost shell of an atom.

Activity 3.2



Form a group and discuss the following:

The following table shows elements of group VIIIA and their atomic numbers. In the space provided fill the valence shell electron configuration and number of valence electrons for the elements. In your discussion include why helium is placed in group VIIIA even though it has only 2 valence electrons.

Elements	Atomic Number	Valence-Shell Electron Configuration	Number of Valence Electrons
Helium, He	2		
Neon, Ne	10		
Argon, Ar	18		
Krypton, Kr	36		
Xenon, Xe	54		
Radon, Rn	86		

Share your discussion with the rest of the class.

Noble gas atoms with eight electrons (except for, He) in the outermost shell are stable. Thus, the ns^2np^6 electronic valence structure has maximum stability. Atoms containing less than eight electrons in their outermost shell are unstable. To attain stability, these atoms tend to have eight electrons in their valence shells. This leads to the explanation of the octet rule. The rule states that atoms tend to gain, lose or share electrons until there are eight electrons in their valence shell.

The type or characteristic of the resulting arrangement depends largely on the type of chemical bonding that exists between the atoms. These are ionic, covalent, and metallic bonds.

Exercise 3.1

1 Why do atoms combine to form compounds?
2 How is a chemical bond formed to make a compound or molecule?
3 Which electron (s) of an atom take (s) part in bond formation?
4 How does the chemical reactivity of halogens compare with that of the noble-gas family?

3.2 IONIC BONDING

Competencies

By the end of this section, you will be able to:

- explain the term ion;
- illustrate the formation of ions by giving examples;
- define ionic bonding;
- describe the formation of an ionic bond;
- give examples of simple ionic compounds;
- draw Lewis structures or electron-dot formulas of simple ionic compounds;
- explain the general properties of ionic compounds; and
- investigate the properties of given samples of ionic compounds.

Activity 3.3

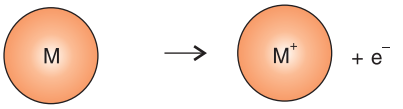


Form a group and discuss each of the following concepts. Share your ideas with the class.

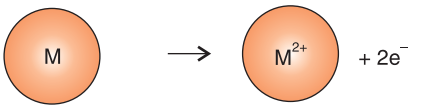
1. Why do some atoms easily lose electrons and others do not?
2. Sodium chloride, NaCl is a good conductor in the form of liquid state, but a non conductor in the form of solid state.

When an atom either loses or gains electrons, it becomes an ion. An ion is an electrically charged particles. Two different types of ions exist. These are the positive ions called cations and the negative ions called anions.

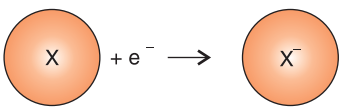
The chemical properties of metals differ from those of non-metals. A metal has 1, 2, or 3 electrons in its outermost shell. Metals tend to lose these electrons and become positively charged ions. For example, if a metal (M) loses one electron, it becomes an ion with a charge of +1.



However, if it loses two electrons it becomes an ion with a charge of +2:



A non-metal may have 4, 5, 6, or 7 electrons in its outermost shell. Non-metals tend to gain electrons to form negatively charged ions. For example, if a non-metal (X) gains one electron, it becomes an ion with a charge of –1.



Note that hydrogen can form both a cation, H^+ (*hydrogen ion*) as in HCl, or an anion H^- (*hydride ion*) as in NaH.

Metals in Group IA, the alkali metals, tend to lose one electron when they combine with other elements, producing cations of +1 charge. For example, Na and K each lose one electron to form ions of +1 charge.

$$\underset{\text{sodium atom}}{Na} \rightarrow \underset{\text{sodium ion}}{Na^+} + e^-$$

$$\underset{\text{potassium atom}}{K} \rightarrow \underset{\text{potassium ion}}{K^+} + e^-$$

On the other hand, Group VIIA elements, the halogens, usually gain one electron and produce an ion with –1 charge. For example, each Cl and Br atom accepts one electron to produce an ion with –1 charge.

$$\underset{\text{chlorine atom}}{Cl} + e^- \rightarrow \underset{\text{chloride ion}}{Cl^-}$$

$$\underset{\text{bromine atom}}{Br} + e^- \rightarrow \underset{\text{bromide ion}}{Br^-}$$

Activity 3.4



Form a group and perform the following tasks:

Consider the elements:

1. Calcium (atomic number = 20)
2. Barium (atomic number = 56)
3. Oxygen (atomic number = 8)
4. Sulphur (atomic number = 16)

   a determine whether each of the elements gain or lose electrons in chemical bond formation.

   b write the type of ions they form; and

   c indicate the charges on the ions formed.

Present your findings to the class.

The following table relates the position of some elements in the periodic table to the ions they normally produce.Note that the charge is the same for each ion in a given group or column.

Table 3.1 Selected ions in the periodic table.

Period	IA	IIA	IIIA	IVA	VA	VIA	VIIA
1	H^+						
2	Li^+	Be^{2+}			N^{3-}	O^{2-}	F^-
3	Na^+	Mg^{2+}	Al^{3+}		P^{3-}	S^{2-}	Cl^-
4	K^+	Ca^{2+}					Br^-
5	Rb^+	Sr^{2+}					I^-
6	Cs^+	Ba^{2+}					
7							

Ionic Bond formation

When two atoms combine, one of the atoms gains electrons and becomes an anion and the other loses electrons to form a cation. When a cation and an anion are brought close to one another, an electrostatic force of attraction is set up between them. This force of attraction between oppositely charged ions is called an ionic bond. It is also called an electrovalent bond. The bond is produced when electrons are transferred from the outermost shell of a metal atom to the outermost shell of a nonmetal atom.

To illustrate ionic bonding, let us consider the formation of the bond between sodium and chlorine. A sodium atom has 1 valence electron. In order to attain the electron configuration of the nearest noble gas (Ne), it has to lose its valence electron and form a sodium ion (Na^+). Chlorine has 7 valence electrons. By gaining 1 electron, chlorine attains the electron configuration of argon (Ar) and form a chloride ion (Cl^-).

In general, an ionic bond is formed by the transfer of electron from a metal to a nonmetal- for example, sodium and chlorine. Atoms that are bound together by an ionic bond form ionic compounds. For example, Na^+ and Cl^- ions form sodium chloride, NaCl. The transfer of an electron from sodium to chlorine and the formation of the ionic bond in sodium chloride is shown with the following shell diagrams.

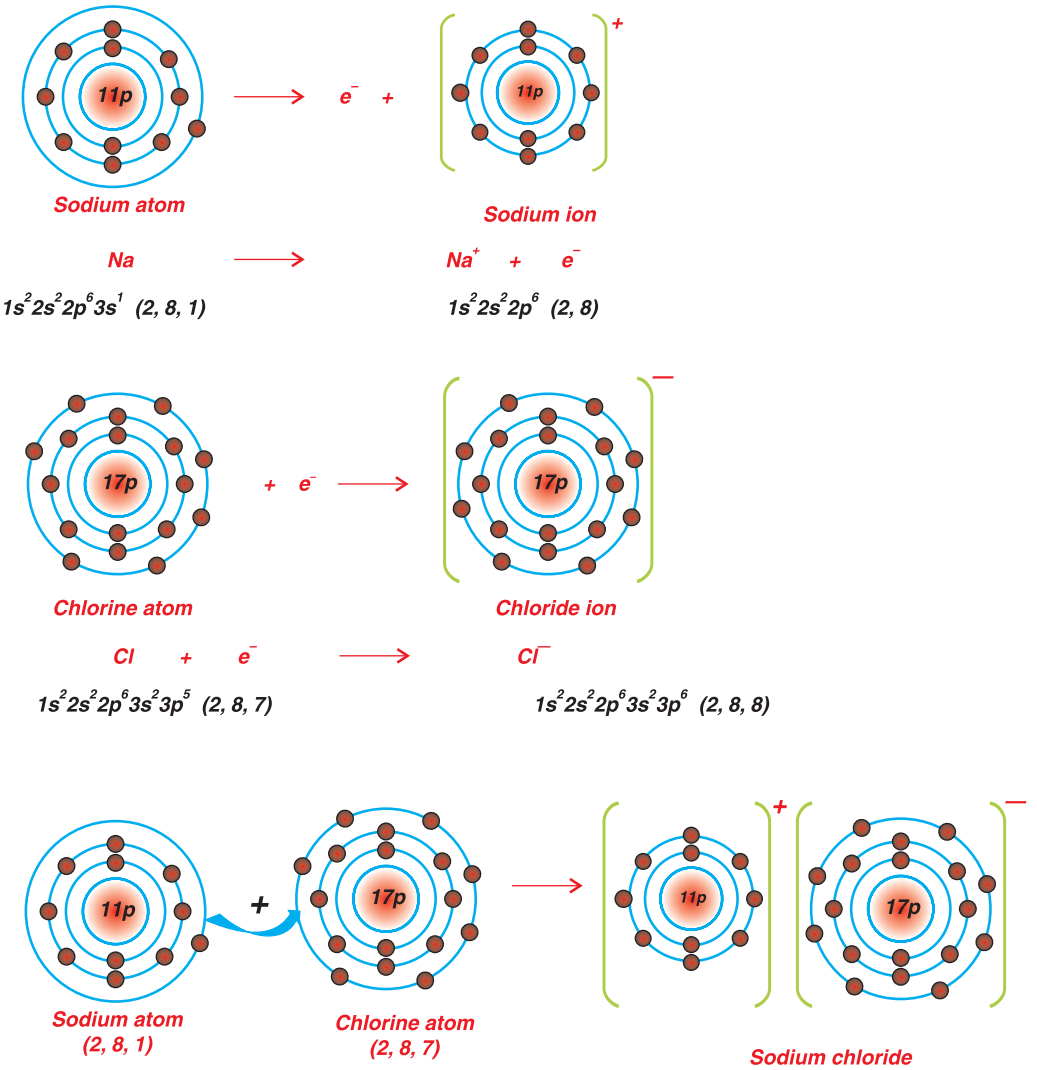


Figure 3.3 Formation of sodium chloride.

Electron-dot notation is often used to represent the outermost shell electron configurations of the elements. These formulas, also called Lewis formulas, consist of the symbol of the element plus dots equal to the number of valence electrons in the atom or ion. Since valence shells contain a maximum of eight electrons, electron-dot symbols contain a maximum of eight dots. Electron-dot formulas of sodium and chlorine

are shown below. Sodium is an alkali metal with one valence electron:

Na: $1s^2 2s^2 2p^6 3s^1$

The Lewis symbol for sodium is Na·.

Chlorine is a halogen with seven valence electrons:

Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$

The Lewis symbol for chlorine is ·Cl·.

The formation of ionic bond can also be represented by using electron-dot formulas. Therefore, the Lewis structure for the ionic compound sodium chloride will be

$$Na\cdot + \cdot\ddot{Cl}: \rightarrow Na^+ [:\ddot{Cl}:]^-$$

Activity 3.5



Form a group and perform each of the following task:

Draw Bohr's diagrammatic representation and write the Lewis formula for the following ionic compounds:

a Potassium chloride
b Magnesium oxide
c Calcium chloride
d Potassium sulphide
e Aluminium oxide

Share your ideas with rest of the class.

The following table illustrates the formation of ionic bonds between representative metals and non-metals. Careful observation indicates that, in each case, both the metal and the nonmetal acquire a noble-gas configuration. The compounds formed in each case are electrically neutral as the sum of positive charges equals the sum of negative charges.

Table 3.2 Summary of formula of ionic-compounds.

Metal group	Non-metal group	Formula of Ionic Compound	Examples
IA	VIIA	$MX\ (M^+X^-)$	NaCl, KBr
IA	VIA	$M_2X\ (2M^+X^{2-})$	Li_2O, K_2O
IA	VA	$M_3X\ (3M^+X^{3-})$,	Na_3N, K_3P
IIA	VIIA	$MX_2\ (M^{2+}2X^-)$,	$MgCl_2$, CaI_2
IIA	VIA	$MX\ (M^{2+}\ X^{2-})$,	BaS, MgO, MgS
IIA	VA	$M_3X_2\ (3M^{2+}\ 2X^{3-})$	Ca_3N_2, Mg_3P_2
IIIA	VIIA	$MX_3\ (M^{3+}\ 3X^-)$,	AlF_3
IIIA	VIA	$M_2X_3\ (2M^{3+}3X^{2-})$	Al_2O_3
IIIA	VA	$MX\ (M^{3+}\ X^{3-})$	AlN

Note: M = metal; X = non-metal.

Exercise 3.2

1. Draw the Lewis structure for the nitrogen atom, nitrogen molecule and ammonia.
2. Show that the following species have the same number of electrons.
   Na^+, Mg^{2+}, O^{2-}, and Ne

How do you name ionic compounds, for example NaCl?

To name an ionic compound that is formed from a metal and a non-metal, follow the given procedure, by considering the example of NaCl:

1. Write the name of the metal (sodium).
2. Modify the last characters of the name of the non-metal (chlorine) to end it with ide. (Chloride).

Therefore, the name of NaCl is sodium chloride. Similarly, MgO is named as magnesium oxide, Na_3N is named as sodium nitride.

General Properties of Ionic Compounds

Activity 3.6

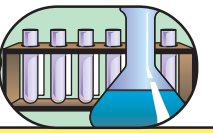


Form a group and perform the following task: Collect samples of ionic compounds from your school laboratory and investigate whether the samples are:

a hard or soft b brittle or strong c liquids or solids

What is your generalization about the physical properties of ionic compounds? Share your ideas with the rest of the class.

Experiment 3.1



Investigating the Physical Properties of Ionic Compounds

I Melting Point and Solubility

Objective: To investigate the melting point and solubility of some ionic compounds.

Apparatus: Test -tube, Bunsen burner.

Chemicals: NaCl, $CuCl_2$, ethanol, hexane and benzene.

Procedure:

Perform the following three experiments:

1. Put a few crystals of dry sodium chloride (NaCl) and copper (II) chloride ($CuCl_2$) into separate glass test tubes. Heat strongly on a Bunsen burner. What do you observe?

2. Place about 1 g each of sodium chloride (NaCl) and copper (II) chloride ($CuCl_2$) in separate test tubes. Add about 5 mL of water (polar solvent) and shake well.
3. Repeat experiment 2 using the following solvents instead of water. Ethanol (polar solvent), hexane and benzene (non-polar solvents). These solvents are highly flammable and should be kept away from flames.

Observations and analysis

a In experiment 1, do the crystals melt? Do they have high or low melting points?

b In experiment 2 and 3, does each of the solids dissolve in the given solvents? Why? Do they have low or high solubility in the given solvents?

Prepare a table as shown below and fill in the results of the solubility tests.

Substances	Water	Ethanol	Hexane	Benzene
NaCl (s)				
$CuCl_2$ (s)				

II Conductivity

Objective: To investigate the conductivity of some ionic compounds.

Apparatus: Beaker, conductivity cell, graphite or iron rods, bulb, wire, battery.

Chemicals: Sodium chloride, copper (II) chloride, benzene and charcoal.

Procedure:

1. Put some amount of sodium chloride crystals in a beaker.
2. Place two electrodes that are made of iron or graphite in the beaker.
3. Connect the two electrodes to a bulb and a 6-volt battery as shown in Figure 3.5.

   Record your observation.
4. Now add distilled water to the beaker and stir dissolve the salt. Observe the changes.
5. Repeat the experiment using aqueous copper (II) chloride, benzene and charcoal.

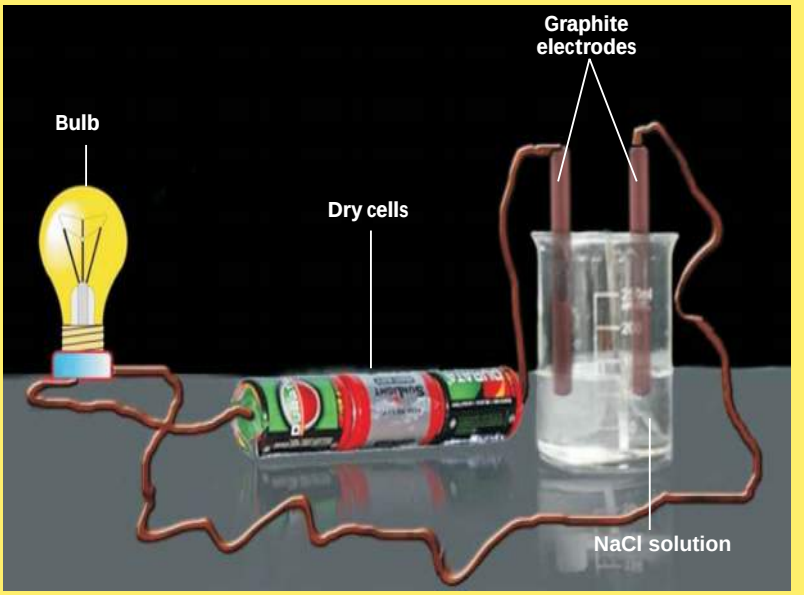


Figure 3.5 Conductivity of sodium chloride solution.

Observations and analysis

a Does sodium chloride conduct electricity in its solid state?

b What do you observe in the bulb when the sodium chloride is dissolved? What does this indicate?

c What can you conclude from this experiment?

d What are your observations for aqueous copper (II) chloride, benzene and charcoal?

Conclusion:

What are your conclusions about:

a melting point,

b solubility in polar and non-polar solvents, and

c electrical conductivity of ionic compounds?

A summary of the general properties of ionic compounds:

1. Ionic compounds do not contain molecules. They are aggregates of positive ions and negative ions. In the solid state, each ion is surrounded by ions of the opposite charge, producing an orderly array of ions called crystal.

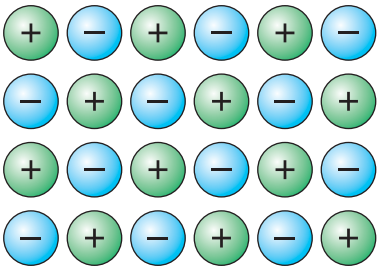


Figure 3.4 Arrangement of ions.

2. At room temperature ionic compounds are hard and rigid crystalline solids. This is due to the existence of strong electrostatic forces of attraction between the ions.
3. Ionic compounds have relatively high melting and boiling points. This is due to the presence of strong electrostatic forces between the ions. These forces can be overcome only by applying very large amounts of energy.
4. Ionic compounds can conduct electric currents when molten or in aqueous solution. This is due to the presence of mobile ions in molten state or in solution. However, ionic compounds do not conduct electricity in the solid state.
5. Ionic compounds are soluble in polar solvents such as water. They are insoluble in non-polar solvents such as benzene.

Exercise 3.3

1. KCl is soluble in water but insoluble in benzene. Explain.
2. Which of the following substances conduct electricity? Give reasons for your answer in each case:

   a NaCl (aq) b NaCl (l) c NaCl (s)
3. Name the ionic compounds formed from the following pairs of elements:

   a calcium and sulphur b sodium and iodine c silver and bromine
4. List the properties of ionic compounds.

3.3 COVALENT BONDING

Competencies

By the end of this section, you will be able to:

- define covalent bonding,
- describe formation of a covalent bond,

- draw Lewis structures or electron-dot formulas of simple covalent molecules,
- give examples of different types of covalent molecules,
- make models of covalent molecules to show single, double and triple bonds using sticks and balls or locally available materials,
- explain polarity in covalent molecules,
- distinguish between polar and non-polar covalent molecules,
- define coordinate (dative) covalent bond,
- illustrate the formation of coordinate covalent bond using appropriate examples,
- explain the general properties of covalent compounds, and
- investigate the properties of given samples of covalent compounds.

Activity 3.7



Form a group and discuss each of the following concepts.

1. What is the difference between the bond when two chlorine atoms combine to form a chlorine molecule (Cl_2) and that formed when sodium combines with chlorine to form sodium chloride (NaCl)?
2. Carbon tetrachloride (CCl_4) is a covalent compound. Would you expect it to be :
   i) a conductor of electricity
   ii) soluble in water.

Share your ideas with the class.

Many molecules are formed when outermost shell or valence electrons are shared between two atoms. This sharing of electrons creates a covalent bond.

Covalent bond formation can be illustrated by the sharing of electrons between two hydrogen atoms to form a molecule of hydrogen.

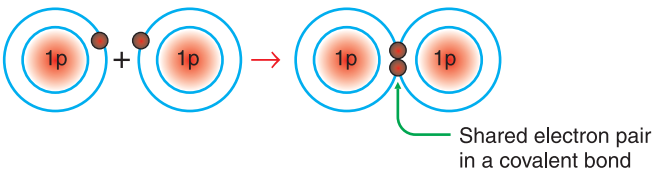


Figure 3.6 Sharing of electrons between hydrogen atoms in H_2 molecule.

In the hydrogen molecule, each hydrogen atom attains the stable electron configuration of helium.

In a covalent bond, each electron in a shared pair is attracted to the nuclei of both atoms as shown in Figure 3.6. The shared electrons spend most of their time between the two nuclei. The electrostatic attraction between the two positively charged nuclei and the two negatively charged electrons hold the atoms in the molecule together. This

attractive force between positively charged nuclei of atoms and the shared electrons in a molecule is known as covalent bond.

A molecule of hydrogen chloride is also formed by a pair of electrons shared between the two atoms. Each atom in the molecule attains a stable electron configuration.

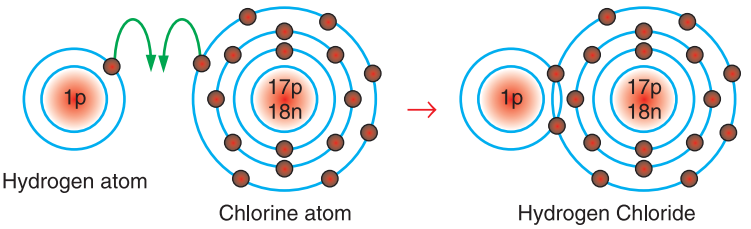


Figure 3.7 Hydrogen and chlorine share a pair of electrons in HCl.

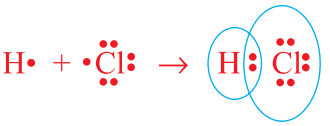
The concept of Lewis formula representation can also be extended to covalent bonds. The Lewis structure for a covalent compound shows the arrangement of atoms in a molecule and all the valence electrons for the atoms involved in the compound. It is conventional to represent the non-bonding (lone pair) electrons by dots and the pair of electrons that are shared between atoms by a dash. For example, consider the hydrogen molecule:

The electron-dot formula of the hydrogen molecule is:

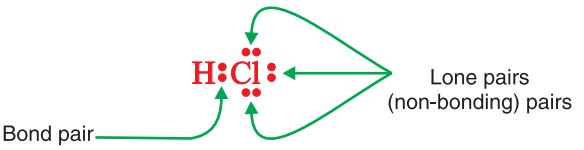
$$H\bullet + \bullet H \rightarrow H:H$$

The covalent bond in hydrogen molecule is also written as H – H.

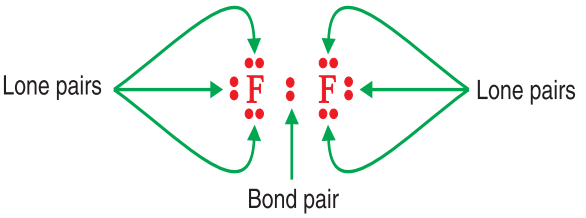
The formation of the covalent bond in hydrogen chloride is shown by the following electron-dot formula. This formula must satisfy the octet rule (for chlorine) and the doublet rule (for hydrogen). As shown in the illustration, these requirements are satisfied. The shared pair belongs to both of the atoms (hydrogen and chlorine) in the hydrogen chloride molecule. The resulting valence electron configuration provides two valence electrons to hydrogen and eight to chlorine.



The chlorine atom in the molecule has three pairs of electrons, which are not used for bonding. Pairs of electrons that is not used for bonding are called lone-pair electrons. Pairs that are used for bonding are called bonding-pair electrons.

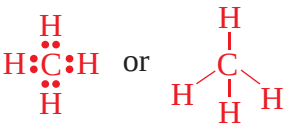


Consider the fluorine molecule, F_2. The electron configuration of fluorine is 2, 7. Thus each fluorine atom has seven valence electrons. Accordingly, there is only one unpaired electron on fluorine. Therefore, the formation of the fluorine molecule is represented as

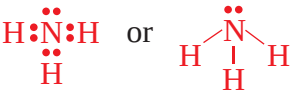


Note that only two valence electrons participate in the formation of fluorine molecule. The others are non-bonding electron (lone pairs). Thus each fluorine atom in fluorine molecule has three lone-pairs of electrons. The resulting molecule is a diatomic molecule. A diatomic molecule consists of two atoms. All the other members of the halogen family form diatomic molecules in the same way as fluorine does.

The maximum number of covalent bonds that an atom can form can be predicted from the number of electrons needed to fill its valence shell. For example, each member of Group IVA elements has four electrons in its valence shell, and it needs four more electrons to achieve stable noble-gas electron configuration. Thus, it forms four covalent bonds for carbon in methane, CH_4 as shown below:



Elements of Group VA need three additional electrons to achieve noble gas configuration and they form three covalent bonds as shown below for nitrogen in ammonia NH_3.



Similarly, elements of group VIA form two covalent bonds and Group VIIA elements form single covalent bonds.



Types of Covalent Bonds

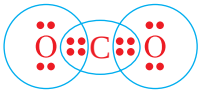
How do you compare the nature and strength of the bonds in H_2, O_2 and N_2?

Atoms can form different types of covalent bonds. These are single bonds, double bonds and triple bonds.

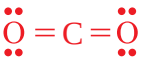
In a single bond two atoms are held together by one electron pair.

How are the covalent bonds in H_2, Cl_2 and HCl formed?

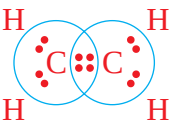
Many covalent compounds are held together by multiple bonds. Multiple bonds are formed when two or three electron pairs are shared by two atoms. If two atoms share two pairs of electrons, the covalent bond is called a double bond. For example, double bonds are found in molecules of carbon dioxide (CO_2) and ethene (C_2H_4).



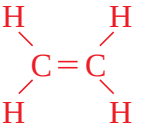
or



Carbon dioxide



or



Ethene

A triple bond is formed when two atoms share three pairs of electrons, as in the nitrogen molecule (N_2).



or



Nitrogen molecule

The ethyne (acetylene) molecule (C_2H_2) also contains a triple bond. In this case the bond is between two carbon atoms.



or



Ethyne

Activity 3.8



Form a group and perform the following tasks:

1. Construct a molecular model for each of the following species using locally available materials such as toothpicks to represent bonds, and styrofoam spheres to represent atoms.

   a N_2 b H_2 c O_2 d C_2H_4

2. Based on your models, identify which species contain:

   i) only single bonds

   ii) double bonds

   iii) triple bonds

Present your models and findings to the rest of the class.

Exercise 3.4

1. How many bonding pair and lone pair electrons are found in each of the following molecules?
   a CO_2 b C_2H_4 c N_2 d C_2H_2
2. Consider molecules of carbon disulfide, CS_2, and hydrogen cyanide, HCN.
   a What types of bonds do they contain?
   b Draw their electron-dot formulas.
   c Are there any lone-pair electrons in these molecules?
3. Why is the melting point of ionic compounds higher than that of covalent compounds?

Reading Check

Does the hydrogen atom form covalent as well as ionic bonds? How?

3.3.1 Polarity in Covalent Molecules

Activity 3.9



Form a group and discuss the following idea:

The covalent bonds are formed by sharing of electrons. Compare covalent bonds formed between atoms of the same elements and those formed between atoms of different elements. (Example: H_2 and HCl).

Present your conclusion to the class.

A covalent bond is formed when electron pairs are shared between two atoms. In molecules like H_2, in which the atoms are identical, the electrons are shared equally between the atoms. A covalent bond in which the electrons are shared equally between the two atoms is called a non-polar covalent bond.

H – H

In other words, a non-polar bond is a covalent bond in which bonding electrons are shared equally between identical atoms, resulting in a balanced distribution of electrical charge.

In contrast, in the covalently bonded HCl molecule, the H and Cl atoms are of different elements; therefore, they do not share the bonding electrons equally.



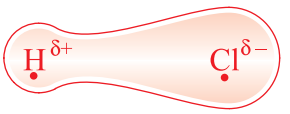
A chemical bond in which shared electrons spend more time in the vicinity of one atom than the other is called a polar covalent bond, or simply a polar bond.

Polarity of bonds is caused by differences in the electronegativity of the two atoms forming the bonds. Electronegativity is the ability of an atom to attract the shared electrons in a chemical bond toward itself.

Elements with high electronegativity have a higher tendency to attract electrons than elements with low electronegativity. For example, in the case of HCl, the electronegativity of the chlorine atom is higher than that of the hydrogen atom. The shared pair of electrons is more strongly attracted to the nucleus of the chlorine atom. As a result, the chlorine atom acquires a partial negative charge (δ^-) whereas the hydrogen atom acquires a partial positive charge (δ^+). The delta is read as "partial" or "slightly."

If a molecule has a positive end and a negative end, it is said to be polar and posses a dipole. Dipole means 'two poles'.

Experimental evidence indicates that, in the HCl molecule, the electrons spend more time near the chlorine atom. We can think of this unequal sharing of electrons as a partial electron transfer or a shift in electron density as shown below:



This unequal sharing of the bonding electron pair results in a relatively higher electron density near the chlorine atom and a correspondingly lower electron density near hydrogen.

Exercise 3.5

1. How many electrons are shared in a:
   a single bond, b double bond, and c triple covalent bond?
2. Draw Lewis structures for:
   a H_2 b Cl_2 c C_3H_6
3. Draw Lewis structures for each of the following molecules:
   a HBr b CO_2 c H_2O
   Also indicate the partial charges using δ^+ and δ^-.

3.3.2 COORDINATE COVALENT BOND

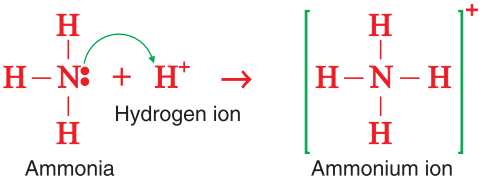
A covalent bond in which one atom donates both electrons of the bond is called a coordinate covalent bond. It is also called a dative bond. Such a bond is hypothetically represented as:



Once formed, a coordinate covalent bond has the same properties as any other covalent bond. The atom that contributes both electrons for the bond is the donor atom, and the atom that shares the electron pair is the acceptor atom.

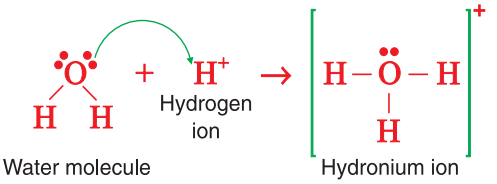
For an atom to act as a donor, it must contain lone pair of electrons in its valence shell and the acceptor atom must have at least one vacant orbital.

For example, the ammonium ion, NH_4^+, is formed by a coordinate covalent bond in which the two non-bonding electrons on NH_3 bond with a hydrogen ion, H^+, which has no electrons to contribute.



In the resulting ion, NH_4^+, the four N – H bonds are identical.

Similarly, a coordinate covalent bond can be formed between a hydrogen ion and a molecule of water, which has two lone pairs of electrons.



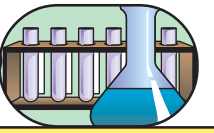
Carbon monoxide, CO, also has a coordinate covalent bond. In order for both carbon and oxygen atoms to attain noble-gas electron arrangements, oxygen donates a pair of electrons to the carbon atom. In the process a coordinate covalent bond is formed between the two atoms.



General properties of covalent compounds

1. Covalent compounds are generally liquids or gases at ordinary temperature. For example : water and ethyl alcohol are liquids. Hydrogen chloride, methane and cabon dioxide are gases. Same covalent compounds are solids (*e.g.* sugar)
2. As compared to ionic compounds, covalent compounds have relatively lower melting points and boiling points.
3. They do not conduct electric current when molten or in aqueous solution, because they consist of molecules rather than of ions.
4. Covalent compounds are insoluble in polar solvents such as water. They are soluble in non-polar solvents such as benzene and carbon tetrachloride.

Experiment 3.2



Investigating the Physical Properties of Covalent Compounds

I. Melting point

Objective: To investigate the effect of heat on covalent compounds.

Apparatus: Test tube, Bunsen burner.

Chemicals: Naphthalene, graphite, iodine.

1. Put a small amount of naphthalene into a dry test tube. Heat it strongly. Naphthalene is toxic. Do not inhale it or get it on your skin.
   a What do you observe? Does the solid melt or vaporize?
   b Is the melting point high or low?
2. Repeat the procedure with graphite, and iodine separately. Iodine vapor is toxic. Do not inhale it.
   Observe and record your observation in a tabular form as shown below.

Substance	Melted, vaporized or nothing happened	High or low melting point
Naphthalene Graphite Iodine		

II. Solubility

Objective: To investigate the solubility of covalent compounds.

Apparatus: Test-tubes, test tube rack.

Chemicals: Naphthalene, graphite, iodine, ethanol, hexane and benzene.

3. Arrange 12 test tubes in three sets (A, B, C) of 4 test tube each. To each test tube of set A, add 1 g of naphthalene. To each test tube of set B add 1 g of graphite and to each test tube of set C add 1 g of iodine.
4. Add about 10 mL of each the following solvents to the four test tubes of each set separately and shake well.
   - Water
   - Ethanol
   - Hexane
   - Benzene

Caution: Ethanol, hexane and benzene are all highly flammable.

Observe and record whether the solids are very soluble, slightly soluble or insoluble.

Substance	Solubility			
	Water	Ethanol	Hexane	Benzene
Naphthalene				
Graphite				
Iodine				

Observations and analysis

Draw general conclusions on the

a melting points,

b solubility in polar and non-polar solvents of the covalent compounds given.

Exercise 3.6

1. Which of the following molecules contain a covalent bond?

   a CaO d SO_2 g MgO
   b HCl e Na_2O h NaH
   c CO_2 f PCl_3 i CH_4

2. Which of the following contain a dative bond?

   a H_3O^+ b NH_3 c NH_2^- d CaO

3. Which of the following molecules are polar ?

   a SO_2 c H_2O e BCl_3 g CH_3Cl
   b CO_2 d CS_2 f CH_4

4. Which of the following are non-polar covalent compounds?

   a O_2 c CH_4 e H_2O g Br_2
   b HCl d O_3 f Cl_2

Activity 3.10



Form a group and perform the following task:

Collect samples of covalent compounds from your school laboratory and investigate whether the samples are:

a liquids or solids b hard or soft c brittle or strong

What is your generalization about the physical properties of covalent compounds? Share your ideas with the class.

Reading Check

What is the difference between a coordinate covalent bond and covalent bond?

3.4 METALLIC BONDING

Competencies

By the end of this section, you will be able to:

- explain the formation of metallic bond;
- explain the electrical and thermal conductivity of metals in relation to metallic bonding; and

Activity 3.11



Form a group and discuss the following concepts and present your discussion to the class.

a Metals are solids. They contain large number of atoms in their crystals. What kind of force do you think holds these metal atoms together?

b How do you account for the properties of metals, such as conductivity, malleability, and ductility in terms of the bonds in metals?

The highest energy orbitals of most metals are occupied by very few electrons. In *s*-block metals, for example, one or two valence electrons occupy *s* orbitals in the outermost levels (for example Na and Mg). Furthermore, the *p* orbitals of the outer most level are also occupied partially in *p*-block metals (example Tl, Pb and Bi).

The *d*-block metals contain partially filled $(n–1)d$ levels in their atomic states or principal oxidation states. The bonding in metals is different from that in other types of crystals. The valence electrons of metals are not held by individual atoms. Rather, they are delocalized and mobile (free to move throughout the structure).

The valence electrons form a sea of electrons around the metal ions and these metal ions are organized as a crystal. Metallic bonding results from the attraction between the metal ions and the surrounding sea of electrons.

For example, as illustrated in Figure 3.8, a sodium metal crystal is a lattice-like array of Na^+ ions surrounded by a sea of mobile bonding valence electrons.

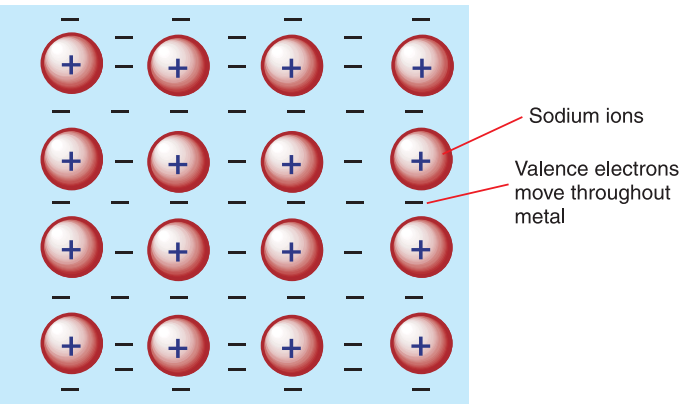


Figure 3.8 Metallic bonding in sodium metal.

The bonding valence electrons move freely throughout the entire crystal. This freedom of movement is responsible for the electrical conductivity of metals.

Properties of Metallic Bonding

Have you ever visited a goldsmith workshop? Why are metals easily shaped into thin sheets and drawn into wires?

The freedom of movement of bonding valence electrons is responsible for the high electrical and thermal conductivity that characterizes the metals. Other properties of metallic bonding contribute to unique properties of metals. For example, most metals are easy to shape due to their malleability and ductility.

Malleability allows a substance such as a metal to be reshaped. By hammering and bending some metals, you can create thin sheets. Ductility allows a substance to be drawn or pulled out into long thin pieces, such as wires.

Metals are malleable and ductile because metallic bonding is the same in all directions throughout the solid.

When we apply a force to metal, its cations swim freely within the sea of electrons without breaking the crystal structure. For example, when you hammer, bend, or pull on a metal to reshape it, you shift its cations around. The force you apply moves the atoms around, for example, around corners in the lattice. This is the basis for malleability and ductility of metals, which allows you to change its shape.

Project work

Model of a Metallic Crystal

Put about one hundred balls (for example, marble balls or balls made from other locally available materials) into a rectangular glass trough. Shake the trough. Allow the balls to settle.

a Draw a two-dimensional diagram to show how the marble balls are now arranged in the trough.

b If the balls represent atoms in a metallic lattice, which species are occupying the 'empty' space between and around them?

Present your model and findings to the class.

Exercise 3.7

1. Describe how a metallic bond is different from those of an ionic bond and a covalent bond.
2. Explain thermal and electrical conductivity in metals.
3. Is metallic bonding responsible to form compounds?

3.5 INTERMOLECULAR FORCES

Competencies

By the end of this section, you will be able to:

- define inter-molecular forces;
- explain hydrogen bonding;
- explain the effects of hydrogen bonding on the properties of substances;
- describe Van der Waal's forces;
- explain dipole-dipole forces;
- give examples of molecules with dipole-dipole forces;
- explain dispersion forces;
- give examples of molecules in which dispersion force is important; and
- compare and contrast the three types of intermolecular forces.

Activity 3.12



Form a group and discuss the following phenomenon: Why do covalent compounds usually exist as gases and liquids?

Share your views with the rest of the class.

Inter-molecular forces are relatively weak forces of attraction that occur between molecules. Inter-molecular forces vary in strength but are generally weaker than the bonds that join atoms in molecules, ions in ionic compounds, and metal atoms in solid metals. Inter-molecular forces acting between molecules include: *dipole-dipole forces*, *London dispersion forces* and *hydrogen bonding*. Dipole-dipole attractions and London forces are collectively called Van der Waal's forces.

A Dipole-Dipole Forces

Dipole-dipole forces are strong inter-molecular forces between polar molecules. A dipole is created by equal but opposite charges separated by a short distance. A polar molecule acts as a tiny dipole because of its uneven charge distribution.

A dipole is represented by an arrow with a head pointing toward the negative pole and crossed tail situated at the positive pole. The dipole created by a hydrogen chloride molecule, which has its negative end at the more electronegative chlorine atom, is as shown below:

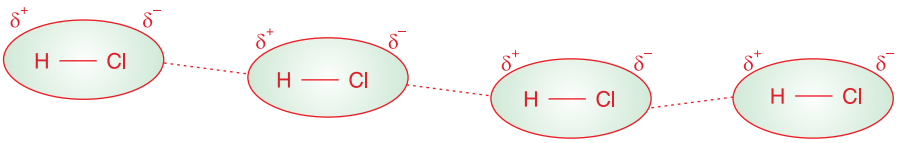


Figure 3.9 Dipole-dipole interactions in HCl molecules.

The negative end in one polar molecule attracts the positive end in an adjacent molecule in a liquid or solid. Dipole-dipole forces occur in molecules such as ethyl alcohol and water.

B London Dispersion Forces

All molecules, including those without dipole moments, exert forces on each other. We know this because all substances, even the noble gases, change from liquid to solid state under different conditions.

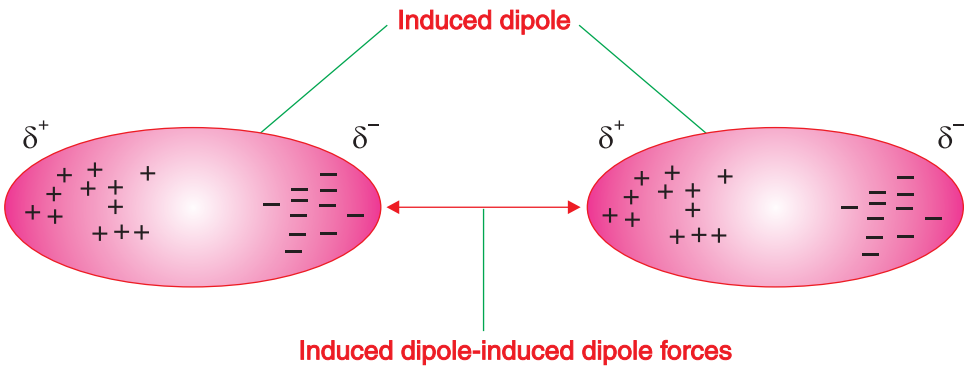


Figure 3.10 Induced dipole-induced dipole forces between non-polar molecules.

London dispersion forces act between all atoms and molecules. They are the only forces that exist between noble gas atoms and non-polar molecules. This fact is reflected in the low boiling points of noble gases and non-polar molecules. Because dispersion forces result from temporary redistribution of the electrons causing induced dipole-dipole interactions, their strength increases with the number of electrons in the interacting atoms or molecules. Hence, dispersion forces increase with atomic number or molar mass. This trend can be seen by comparing the boiling points of gases (helium, He, and argon, Ar), (hydrogen, H_2, and oxygen, O_2), and (chlorine, Cl_2, and bromine, Br_2).

As an illustration, the boiling points of the noble gases are presented in Table 3.4.

Table 3.3 Boiling points of noble gases.

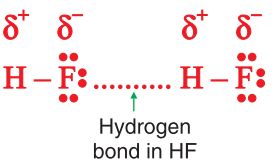
Noble gas	Boiling Point (°C)	Number of electrons
He	-269	2
Ne	-246	10
Ar	-186	18
Kr	-152	36
Xe	-107	54
Rn	-62	86

As you look down the column of noble gases, you note that boiling point increases. This is because the induced dipole-dipole interaction increases.

C Hydrogen Bonding

Hydrogen bonding is a particular type of intermolecular force arising when a hydrogen atom is bonded to highly electronegative elements, fluorine, oxygen and nitrogen. Hydrogen bonding is a particular type of dipole-dipole interactions between polar compounds. In such compounds, large electronegativity differences between the hydrogen and the fluorine, oxygen, or nitrogen atoms make the bonds connecting them highly polar. This polarity gives the hydrogen atom a positive charge. Moreover, the small size of the hydrogen atom allows the atom to come very close to an unshared pair of electrons on an adjacent molecule. Hydrogen bonding is responsible for the unusual high boiling points of some compounds such as hydrogen fluoride (HF), water (H_2O) and ammonia (NH_3).

Hydrogen bonds are usually represented by dotted lines connecting the hydrogen atom to the unshared electron pair of the electronegative atom to which it is attracted. For example, the hydrogen bond in hydrogen fluoride, HF, results when the highly electronegative F atom attracts the H atoms of an adjacent molecule.



Do you think that the intermolecular forces between molecules containing C-H, N-H, and O-H bonds are as strong as the intermolecular forces containing F-H bonds?

Exercise 3.8

1. Which of the following exists predominately in the water (H_2O) molecule?
   - a Van der Waal's force
   - b hydrogen bond
   - c coordinate covalent bond
   - d none of these
2. Which of the following has the highest induced dipole interactions in its molecule?
   - a He
   - b Ar
   - c Ne
   - d Kr

Critical Thinking

Oxygen ($^{16}_{8}O$) and Sulphur ($^{32}_{16}S$) are in the same group in the periodic table. They form compounds with hydrogen, H_2O and H_2S. However, H_2O is a liquid, whereas H_2S is a gas at room temperature. Give explanation?

Check List

Key terms of the unit

- *Anions*
- *Bonding-pair electrons*
- *Cations*
- *Chemical bond*
- *Conductivity*
- *Coordinate/dative bond*
- *Covalent bond*
- *Delocalized electrons*
- *Dipole*
- *Dipole-dipole interaction*
- *Double-bond*
- *Ductility*
- *Electronegativity*
- *Electrovalent bond*
- *Hydrogen bonding*
- *Inter-molecular forces*
- *Ionic bonding*
- *Ionization energy*
- *London forces*
- *Lone-pair electrons*
- *Malleability*
- *Metallic bond*
- *Mobile electrons*
- *Noble gases*
- *Non-bonding electrons*
- *Octet rule*
- *Polar bond*
- *Polar covalent bond*
- *Polarity*
- *Sea of electrons*
- *Single-bond*
- *Triple-bond*
- *Valence electrons*
- *Van der Waal's forces*

Unit Summary

- A chemical bond is the attractive force that binds atoms together to form a molecule (or a crystal lattice), an ionic or metallic crystal lattice.
- An ionic bond is the electrostatic attraction between oppositely charged ions (cations and anions).
- A covalent bond is formed by a shared pair of electrons.
- A covalent bond in which one pair of electrons is shared is known as a single bond; for example, H_2 written as H – H.
- Atoms can also share more than one pair of electrons to form a multiple bond.
- The sharing of three pairs of electrons forms a triple bond - for example, N_2, written as N≡N.
- A dative or coordinate covalent bond is a bond in which one of the atoms supplies both of the shared electrons to the covalent bond.
- A metallic bond is the electrostatic attraction between positive metal ions and delocalized electrons.

- Inter-molecular forces are forces of attraction between covalently bonded molecules. These include: London forces, dipole-dipole forces, and hydrogen bonding.
- London forces are forces of attractions between nonpolar molecules.
- Dipole-dipole forces are the attractions between dipoles of polar molecules.
- Hydrogen bonding is the attraction of covalently bonded hydrogen to lone pairs on N, O, or F atoms in other molecules or in the same molecule (if the molecule is large enough).

REVIEW EXERCISE ON UNIT 3

Part I: Multiple Choice Type Questions

1. Which of the following contain localized valence electrons?
   a Cu c C_6H_6
   b C_2H_6 d none
2. Which of the following is a compound?
   a Cl_2 c Na
   b HCl d liquid oxygen
3. In the formation of ionic bonding, valence electrons are:
   a shared c transferred
   b delocalized d not affected
4. What force is responsible for the formation of ice during the freezing of water?
   a ionic c dipole-dipole interaction
   b covalent d dative bond
5. Which of the following has the highest electrical conductivity in the solid state?
   a water c sodium
   b common salt d rubber

Part II: Write the missing words in your exercise book

6. Ionic bonding is formed between the atoms of _____ and _____.
7. Covalent bonding is formed between the atoms of _____.
8. Metallic bonding is formed between the atoms of _____.
9. The forces that hold atoms together in molecules of compounds are called _____.

Part III: Short Answer Type Questions

10. Give a simple explanation of the following, using an example:
    a a covalent bond
    b an ionic bond
    c a dative bond
    d hydrogen bonding
    e metallic bond

11. Classify the bonds that can be formed between the following pairs of atoms as principally ionic or covalent:
    a calcium and chlorine
    b boron and carbon
    c sodium and bromine
    d magnesium and nitrogen
    e sodium and hydrogen
    f aluminium and oxygen
    g chlorine and oxygen
    h iodine and chlorine

12. Give the Lewis structures for the following:
    a H_2S
    b CaS
    c Al_2O_3
    d HF
    e N_2
    f C_2H_4
    g NH_3
    h CH_4
    i CF_4
    j NO
    k $CaCl_2$

13. Which of the following molecules can form hydrogen bonding?
    a H_2S
    b CO_2
    c SO_2
    d CH_4
    e NH_3
    f HF
    g CH_3OH

14. Which of the following substances contain hydrogen bonding?
    a hydrogen chloride
    b water
    c ammonia
    d methane

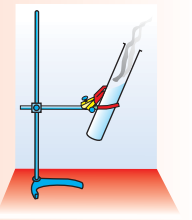
15. Classify the following molecules as polar or non-polar:
    a CH_4
    b CH_3Cl
    c C_2H_2
    d CO_2
    e H_2O_2
    f BCl_3
    g H_2S
    h HBr

16. What is meant by a polar molecule?

17. Explain the fact that HCl is polar, whereas Cl_2 is a non-polar molecule.

18. What is the difference between intermolecular and intra-molecular attractive forces?

UNIT 4



Chemical Reactions and Stoichiometry

Unit Outcomes

After completing this unit, you will be able to:

- *understand the fundamental laws of chemical reactions and how they are applied*;
- *develop skills in writing and balancing chemical equations*;
- *understand energy changes in chemical reactions*;
- *know types of chemical reactions*;
- *develop skills in solving problems based on chemical equations* (*mass-mass, volume-volume and mass-volume problems);*
- *develop skills in determining limiting reactant, theoretical yield, actual yield, and percentage yield*;
- *understand oxidation-reduction reactions and analyze redox reactions by specifying the oxidizing agents and reducing agents, the substance reduced or oxidized;*
- *understand the rate of chemical reaction, the state of a chemical equilibrium and factors affecting them; and*
- *demonstrate scientific enquiry skills: observing, inferring, predicting, classifying, comparing and contrasting, communicating, measuring, asking questions, designing experiments, interpreting data, drawing conclusions, applying concepts, relating cause and effect, and problem-solving.*

MAIN CONTENTS

4.1 Introduction
4.2 Fundamental Laws of Chemical Reactions
4.3 Chemical Equations
4.4 Energy Changes in Chemical Reactions
4.5 Types of Chemical Reactions
4.6 Stoichiometry
4.7 Oxidation-Reduction Reactions
4.8 Rate of Chemical Reaction and Chemical Equilibrium
– Unit Summary
– Review Exercises

Start-up Activity

A chemical reaction enables a space shuttle to be launched, which is powered by a chemical reaction between pure liquid hydrogen (serving as a fuel) and oxygen. Assume that the fuel tank contains 32,000 litres of H_2 and the oxidizer tank contains 40,000 litres of O_2;

Analysis

1. What type of reaction takes place?
2. Write the balanced chemical equation for the reaction.
3. What volume of product is formed in the reaction?
4. What mass of product is formed in the reaction?
   (Assume that the pressure remains constant in this process).
   Submit your findings to the teacher.

4.1 INTRODUCTION

Competencies

By the end of this unit, you will be able to:

- define chemical reaction; and
- give some examples of chemical reactions.

Activity 4.1



Form a group and discuss the following phenomenon:

1. When a space shuttle leaves the ground on its way into orbit, what does the brightness and warmth of the flame indicate?
2. What are the notations that indicate a chemical change might be taking place?

Present your conclusion to the class.

Chemical reactions are the basis of chemistry. Chemical reactions occur around us all the time. For example, the burning of fuel, the souring of milk, metabolic processes of our body and the decay of plants are some familiar chemical reactions in daily life.

A chemical reaction is the process in which reacting substances, called reactants, are converted to new substances, called products. The characteristics of the products are completely different from those of the reactants. The conversion process is a chemical change.

Reactants $\rightarrow$ Products

For example, if you burn magnesium with oxygen, the magnesium and oxygen are completely converted to magnesium oxide. Magnesium oxide is a soft, white, crumbling powder. These characteristics of magnesium oxide are completely different from the characteristics of the original substances, magnesium and oxygen. Magnesium and oxygen are no longer present in the elemental form.

In summary, a chemical reaction has occurred in which the reactants, magnesium and oxygen, underwent a complete chemical change, giving the product magnesium oxide.

All chemical reactions include three types of changes in the original substances. These are changes in composition, properties and energy.

Activity 4.2



Form a group and perform the following task.

List some chemical processes that occur in your daily life. Identify the reactants and products in each of these chemical processes.

Present your findings to the class.

Note that, in daily life, we use different terms for the same process of chemical change. For example "the souring of milk" occurs due to the process of fermentation. In scientific discussion we generally have a single term for each process.

4.2 FUNDAMENTAL LAWS OF CHEMICAL REACTIONS

Competencies

***By the end of this unit, you will be able to*:**

- state the law of conservation of mass and illustrate the law, using examples;
- demonstrate the law of conservation of mass, using simple experiments;
- state the law of definite proportion and illustrate it, using examples;
- demonstrate the law of definite proportion, using a simple experiment; and
- State the law of multiple proportion and illustrate it, using examples.

Activity 4.3



Form a group and discuss the following phenomenon. When wood burns, the ash weighs much less than the original wood. Where did the "lost mass" go? How can you estimate the mass of the wood that is no longer present?

Present your conclusion to the class.

While investigating the quantitative relations between substances in chemical reactions, scientists formulated the three basic laws of chemical combination. These are:

i) The law of conservation of mass

ii) The law of definite proportions

iii) The law of multiple proportions

i) *The Law of Conservation of Mass*

In 1774, the French chemist Antoine Lavoisier performed an experiment in which he heated a sealed glass container that held a sample of tin in air. He measured the mass of the substances before and after heating and found them to be the same. This and other similar experimental observations became the basis of the law of conservation of mass.

Historical Note



Antoine Lavoisier

French chemist Antoine Lavoisier is considered the founder of modern chemistry. He found that the amount of matter before a chemical reaction is equal to the amount of matter afterwards, even though the matter may change in its form. Lavoisier also experimented with the role of oxygen in combustion and respiration in both plants and animals.

The law of conservation of mass states that matter is neither created nor destroyed in a chemical reaction. In other words, the mass of the reactants is exactly equal to the mass of the products, within the limits of experimental error. This law is also known as the law of indestructibility of matter.

Mass of reactants = Mass of products

There is no loss or gain of substances during a chemical reaction, and mass is conserved.

For example, consider the decomposition of mercury (II) oxide. When 100 g of mercury (II) oxide decomposes by heat, 92.6 g of mercury and 7.4 g of oxygen are formed. Note that the total mass of mercury and oxygen after decomposition is 100 g:

$$\underset{\underbrace{100\text{g}}_{100\text{g}}}{2\text{HgO}} \rightarrow \underbrace{2\text{Hg} + \underset{92.6\text{g}}{\text{O}_2} \quad \underset{7.4\text{g}}{}}_{100\text{g}}$$

Reactant Products

Experiment 4.1



Investigation of the Law of Conservation of Mass

Objective: To determine the mass of substances before and after a reaction.

Apparatus: Flask, test tube, thread, rubber stopper, balance.

Chemicals: Sodium chloride, silver nitrate.

Procedure:

1. Take 50 mL of silver nitrate solution in a conical flask.
2. Tie a thread around the top of a test tube. Fill the test tube with a saturated solution of sodium chloride. Suspend the test tube in the flask by means of a thread held by a rubber stopper, as shown in Figure 4.1.
3. Weigh the flask (*and its contents*). Record the result as m_1.
4. Mix the liquids by tilting the conical flask so that the sodium chloride pours into the silver nitrate solution.
5. Reweigh the conical flask and contents and record as m_2. Compare m_1 and m_2.

Observations and analysis

1. What was the colour of the solution after the reaction?
2. Is there any difference in mass between m_1 and m_2?
3. What is your conclusion from this experiment?
4. Write the balanced chemical equation for the reaction.

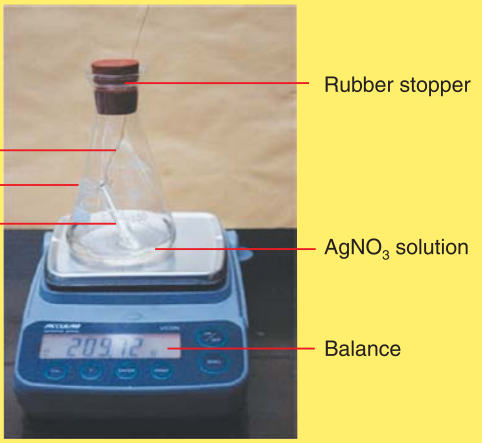


Figure 4.1 Conservation of mass.

ii) The Law of Definite Proportions

The law of *definite proportions states* that a compound always contains the same elements in the same proportion by mass. This means that all pure samples of a compound have the same composition regardless of the source of sample. This law is also known as the law of constant composition. For example, sample of water could be obtained from different sources, such as from a river, the ground, or the ocean. But whatever the original source, all forms of pure water contains 11.2% hydrogen and 88.8% oxygen by mass. These percentages represent a ratio of 1.0 to 8.0 (1:8), by mass, of hydrogen to oxygen.

This ratio is constant (fixed) for water. In other words, a compound with a different ratio of hydrogen and oxygen is not water.

Similarly, in forming the compound ZnO, 65.0 g of zinc combines with 16.0 g of oxygen. This is 80.2% zinc and 19.8% oxygen, by mass. As is the case for water, the composition of ZnO is constant. In forming ZnO, zinc combines with oxygen in a definite proportion.

Experiment 4.2



Investigation of the Law of definite proportions

Objective: To determine the mass of copper from copper (II) oxide.

Apparatus: Burner, stand, combustion tube, two glass test tubes, two watch glasses.

Chemicals: Copper powder, copper (II) carbonate, hydrogen gas

Procedure:

1. Prepare samples of copper (II) oxide using the following two methods:
   - *i*) Make copper (II) oxide by heating copper powder in one of the test tubes.
   - *ii*) Make copper (II) oxide by heating copper (II) carbonate in the second test tube.

     (In this case, the heating process produces a chemical change through thermal decomposition)
2. Take 1 g from each of the samples of copper (II) oxide (*from i and ii*). Place each of these samples in a watch glass.
3. Reduce each of these samples: use the combustion tube to heat the samples in a stream of hydrogen as shown in Figure 4.2.
4. Weigh the copper metal that remains in each case. Compare the measurements.

Observations and analysis

1. What is the mass of copper produced in each case?
2. Why is copper metal produced in each case?
3. What can you conclude from the experiment? Write a short report on your observations.

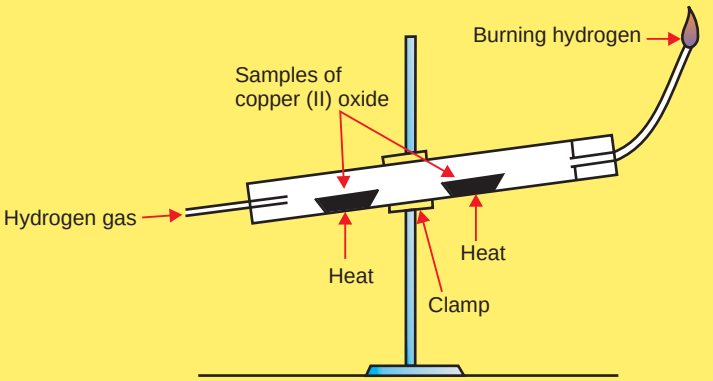


Figure 4.2 Reduction of copper (II) oxide by hydrogen.

iii) The Law of Multiple Proportions

The law of multiple proportions states that when two elements combine to form more than one compound, the masses of one element combined with a fixed mass of the second element are in the ratio of small whole numbers. This law can be illustrated by the two oxides of carbon. The two oxides of carbon are carbon monoxide (CO) and carbon dioxide (CO_2). In CO_2, 1.0 g of carbon is combined with 2.67 g of oxygen; whereas in CO, 1.0 g of carbon is combined with 1.33 g of oxygen. By comparing 2.67 g of oxygen with 1.33 g of oxygen, it is found that the masses of oxygen in the two compounds that combine with the same mass of carbon are in the simple whole number ratio, 2:1.

$$\frac{2.67\text{ g of oxygen in } CO_2}{1.33\text{ g of oxygen in CO}} = \frac{2}{1} = 2:1$$

Activity 4.4



Form a group and perform the following task:

The following table illustrates the law of multiple proportions using five oxides of nitrogen. In the table, fill the mass ratio of nitrogen to oxygen and determine the mass of oxygen in each compound that combine with a fixed mass (1g) of nitrogen.

Present your conclusion to the class.

Compound	Molecular formula	Mass ratio of N to O (N:O)	Ratio of Oxygen in the oxides
Dinitrogen monoxide	N_2O		
Nitrogen monoxide	NO		
Dinitrogen trioxide	N_2O_3		
Nitrogen dioxide	NO_2		
Dinitrogen pentoxide	N_2O_5		

Exercise 4.1

Give appropriate answers for the following questions.

1. Classify the following as chemical or physical changes:

   a the souring of tella b freezing ice cream

   c plant growth d boiling of an egg

   e heating sugar f fermentation

   g the magnetization of iron h the fading of dye in cloth

2. Iron and chlorine form two compounds, A and B. Compound A contains 1.27 g of chlorine for each 1 g of iron whereas compound B contains 1.9 g of chlorine for each 1 g of iron. Show that the masses of chlorine are in the ratio 2:3. Do they obey the law of multiple proportions? Explain.
3. Consider the following two chemical changes:

   i) When a material made of iron rusts, its mass increases.

   ii) When a match stick burns, its mass decreases.

   Do you think that these two observations violate the law of conservation of mass? Explain.

Critical Thinking

4. Discuss how the law of conservation of matter is explained by Dalton's atomic theory.

4.3 CHEMICAL EQUATIONS

Competencies

By the end of this unit, you will be able to:

- describe the conventions used to write chemical equations;
- balance chemical equations, using the inspection method;

- balance chemical equations, using the Least-Common-Multiple (LCM) method.

Activity 4.5



Form a group and discuss each of the following:

1. What is the difference between a chemical equation and a chemical reaction?
2. Which law is satisfied when a chemical equation is balanced? Take a simple chemical reaction to illustrate this law.

Present your conclusion to the class.

A chemical equation is a shorthand representation of a chemical reaction in terms of chemical symbols and formulas. In a chemical equation the starting substances are called reactants; and the new substances produced are known as products.

Reactants are written on the left side and products on the right side of the equation. An arrow (→) is placed between the two sides to indicate transformation of reactants into products.

Reactants → products

4.3.1 Writing Chemical Equation

In writing chemical equation, instead of using words, chemical symbols and formulas are used to represent the reaction.

Steps to Write a Chemical Equation

1. Write a word equation: A word equation is stated in words. For example, the word equation for the reaction between sodium and chlorine to produce sodium chloride is written as:

   Sodium + Chlorine → Sodium chloride (*word equation*)

   Note that we read the ' + ' sign as 'reacts with' and the arrow can be read as 'to produce', 'to form', 'to give' or 'to yield'.

2. Write the symbols and formulas for the reactants and products in the word equation.

   $Na + Cl_2 \rightarrow NaCl$ (*Chemical equation*)

3. Balance the equation.

   $2Na + Cl_2 \rightarrow 2NaCl$

Generally, any chemical equation must fulfil the following conditions:

i) The equation must represent a true and possible chemical reaction.

ii) The symbols and formulas must be written correctly. The elements– hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine and iodine exist as diatomic molecules. These elements should be written as molecules in the equation.

iii) The equation must be balanced.

A chemical equation has both *qualitative* and *quantitative meanings*.

Qualitatively, a chemical equation indicates the types of the reactants and products in the reaction.

Quantitatively, a chemical equation expresses the relative number (*amount)* of moles, molecules or masses of the reactants and products.

4.3.2 Balancing Chemical Equation

Which should be adjusted in balancing a chemical equation, the subscripts or the coefficients?

According to the law of conservation of mass, atoms are neither created nor destroyed during a chemical reaction. As a result, the number of atoms of each element should remain the same before and after the reaction. Therefore, the main reason why all chemical equations must be balanced is just to obey the law of conservation of mass.

To balance a chemical equation means to equalize the number of atoms on both sides of the equation by putting appropriate coefficients in front of the formulas.

Only two methods of balancing chemical equations will be discussed under this topic. These are the inspection and the Least Common Multiple (LCM) method.

1. *The Inspection Method*

Most simple chemical equations can be balanced using this method. Balancing an equation by inspection means to adjust coefficients by trial and error until the equation is balanced. Follow the following four steps to balance the chemical equation.

Step 1: Write the word equation.

Step 2: Write the correct symbols or formulas for the reactants and products.

Step 3: Place the smallest whole number coefficients in front of the symbols or formulas until the number of atoms of each element is the same on both sides of the equation.

Step 4: Checking: By counting the number of atoms on both sides of the equation, make sure that the atoms of all elements are balanced and also the coefficients are expressed as the smallest whole number ratio.

Note:

When you balance an equation, do not change any symbol or formula of any compound. If you change a symbol or formula, it no longer represents the element or compound required by the equation.

Example 1

Balance the equation for the reaction between magnesium and oxygen to produces magnesium oxide.

Solution:

Step 1: Magnesium + Oxygen $\rightarrow$ Magnesium oxide

Step 2: $Mg + O_2 \rightarrow MgO$ (*unbalanced*)

Step 3: Put coefficients to balance the equation

- Oxygen is not balanced. There are two oxygen atoms on the left side and one on the right side. Hence, place the coefficient 2 in front of MgO.

$$Mg + O_2 \rightarrow 2MgO \text{ (}unbalanced\text{)}$$

- Now Mg is not balanced. There is one Mg on the left side and two on the right side. Thus, place the coefficient 2 in front of Mg.

$$2Mg + O_2 \rightarrow 2MgO \text{ (}balanced\text{)}$$

Step 4: **Checking:** There are two Mg and two O atoms on each side of the equation. Therefore, the equation is correctly balanced.

$$2Mg + O_2 \rightarrow 2MgO$$

Exercise 4.2

Balance the following chemical equation, using the inspection method:

1. $Na + H_2O \rightarrow NaOH + H_2$
2. $CaCO_3 \rightarrow CaO + CO_2$
3. $H_2O_2 \rightarrow H_2O + O_2$
4. $Al + H_3PO_4 \rightarrow AlPO_4 + H_2$
5. $HNO_3 + H_2S \rightarrow NO + S + H_2O$

2. *The LCM Method*

In the LCM method, the coefficients for the balanced chemical equation are obtained by taking the LCM of the total valency of reactants and products and then dividing it by total valency of reactants and products. All the necessary steps to balance a chemical equation by the LCM method, are shown by the following examples.

Example 2

When aluminium reacts with oxygen, aluminium oxide is formed. Write the balanced chemical equation for the reaction.

Solution:

Step 1: Represent the reaction by a word equation.

$$\text{Aluminium} + \text{Oxygen} \rightarrow \text{Aluminium oxide}$$

Step 2: Change the words to symbols and formulas for the reactants and products.

$$Al + O_2 \rightarrow Al_2O_3$$

Step 3: Place the total valency of each atom above it.

$$\overset{3}{Al} + \overset{4}{O_2} \rightarrow \overset{6}{Al_2}\ \overset{6}{O_3}$$

Now the equation shows

- The valency of aluminium as 3.
- The total valency of oxygen is $2 \times 2 = 4$.
- The total valency of aluminium in Al_2O_3 is $3 \times 2 = 6$.
- The total valency of oxygen in Al_2O_3 is $2 \times 3 = 6$.

Step 4: Find the LCM of each total valency and place it above the arrow.

$$\overset{3}{Al} + \overset{4}{O_2} \xrightarrow[\text{L.C.M}]{12} \overset{6}{Al_2}\ \overset{6}{O_3}$$

Step 5: Divide the LCM by each total valency number to obtain the coefficients for each of the reactants and products. Place the obtained coefficients in front of the respective formulas.

$$4Al + 3O_2 \rightarrow 2Al_2O_3 \text{ (}\textit{balanced}\text{)}$$

Checking: There are 4 aluminium and 6 oxygen atoms on both sides of the equation. Hence, the chemical equation is correctly balanced.

Example 3

When iron reacts with water, iron (III) oxide and hydrogen are produced. Write the balanced equation.

Solution:

Step 1: Iron + water $\rightarrow$ Iron (III) oxide + hydrogen.

Step 2: $Fe + H_2O \rightarrow Fe_2O_3 + H_2$

Step 3: $\overset{3}{Fe} + \overset{2}{H_2}\overset{2}{O} \rightarrow \overset{6}{Fe_2}\overset{6}{O_3} + \overset{2}{H_2}$

Step 4: $\overset{3}{Fe} + \overset{2}{H_2}\overset{2}{O} \xrightarrow[\text{L.C.M}]{6} \overset{6}{Fe_2}\overset{6}{O_3} + \overset{2}{H_2}$

Step 5: $2Fe + 3H_2O \rightarrow Fe_2O_3 + 3H_2$ (*balanced*)

Checking: There are 2 iron, 6 hydrogen, and 3 oxygen atoms on each side of the equation. Thus, the equation is balanced.

Example 4

The reaction of ammonium sulphate with aluminium nitrate would form aluminium sulphate and ammonium nitrate.

Solution:

Step 1: Ammonium sulphate + Aluminium nitrate $\rightarrow$ Aluminium sulphate + Ammonium nitrate

Step 2: $(NH_4)_2SO_4 + Al(NO_3)_3 \longrightarrow Al_2(SO_4)_3 + NH_4NO_3$

Step 3: $\overset{2}{(N}H_4)_2\overset{2}{SO_4} + \overset{3}{Al}(\overset{3}{NO_3})_3 \longrightarrow \overset{6}{Al_2}(\overset{6}{SO_4})_3 + \overset{1}{NH_4}\overset{1}{NO_3}$

Step 4: $\overset{2}{(NH_4)_2}\overset{2}{SO_4} + \overset{3}{Al}(\overset{3}{NO_3})_3 \xrightarrow[\text{L.C.M}]{6} \overset{6}{Al_2}(\overset{6}{SO_4})_3 + \overset{1}{NH_4}\overset{1}{NO_3}$

Step 5: $3(NH_4)_2SO_4 + 2Al(NO_3)_3 \longrightarrow Al_2(SO_4)_3 + 6NH_4NO_3$ (balanced)

Checking: There are 12 nitrogen, 24 hydrogen, 3 sulphur, 30 oxygen and 2 aluminium atoms on both sides of the equation. Thus, the equation is correctly balanced.

Exercise 4.3

1. Write the balanced chemical equation to represent the following reactions.
   - a Sulphur dioxide reacts with oxygen to produce sulphur trioxide.
   - b Potassium chlorate when heated produces potassium chloride and oxygen.

c Sodium carbonate reacts with hydrochloric acid to form water, carbon dioxide and sodium chloride.

d Silver oxide decomposes to silver and oxygen gas.

2. Balance the following equations by the LCM method.

a $PCl_5 + H_2O \longrightarrow H_3PO_4 + HCl$

b $Mg + H_2O \longrightarrow Mg(OH)_2 + H_2$

c $Zn(NO_3)_2 \longrightarrow ZnO + NO_2 + O_2$

d $H_2SO_4 + NaOH \longrightarrow Na_2SO_4 + H_2O$

e $NH_3 + O_2 \longrightarrow NO + H_2O$

4.4 ENERGY CHANGES IN CHEMICAL REACTIONS

Competencies

***By the end of this section, you will be able to*:**

- explain energy changes in chemical reactions;
- define endothermic and exothermic reactions;
- describe endothermic and exothermic reactions;
- illustrate endothermic and exothermic reactions using diagrams;
- conduct simple experiment to demonstrate exothermic and endothermic reactions;
- describe the importance of chemical changes in production of new substances and energy.

Activity 4.6



Form a group and discuss each of the following phenomena:

When the bread baked, does the bread absorb or release heat energy? Justify your answer.

Present your conclusion to the class.

Almost all chemical reactions are accompanied by energy changes. These energy changes could be in the form of heat energy, light energy, electrical energy, and so on.

On the basis of energy changes, chemical reactions can be divided into exothermic and endothermic reactions.

4.4.1 Exothermic and Endothermic Reactions

Can heat energy be considered as a reactant or product?

Exothermic Reaction

A chemical reaction that releases heat energy to the surroundings is known as an exothermic reaction. During an exothermic process, heat is given out from the system to its surroundings and this heat energy is written on the right side of the equation as shown below.

$$\text{Reactants} \rightarrow \text{Products} + \text{Heat}$$

For example, the burning of carbon with oxygen produces carbon dioxide and heat is released during the reaction. Thus, the reaction is exothermic and written as:

$$C + O_2 \rightarrow CO_2 + \text{Heat}$$

Endothermic Reaction

A chemical reaction which absorbs heat energy from the surroundings is known as an endothermic reaction. During an endothermic process, heat flows into the system from its surroundings and the heat is written on the left side of the equation.

$$\text{Reactants} + \text{Heat} \rightarrow \text{products}$$

For example, the reaction between carbon and sulphur to form carbon disulphide is an endothermic reaction because heat is absorbed in the reaction.

$$C + 2S + \text{Heat} \rightarrow CS_2$$

The amount of heat energy liberated or absorbed by a chemical reaction is called heat of reaction or change in enthalpy for the reaction. It is symbolized as ΔH. Its unit is expressed in kilojoules per mol $\left(\frac{kJ}{mol}\right)$. The change in enthalpy (ΔH) is the difference between the energy of the products and the energy of the reactants.

$\Delta H = H_p - H_r$; where H_p is the heat content (*energy*) of the product, H_r is the heat content (*energy*) of the reactant.

4.4.2 Energy Diagrams

For endothermic reactions, ΔH is positive because the energy of the product is higher than the energy of the reactant. As a result, the enthalpy of the system increases as shown in Figure 4.3.

since, $H_p > H_r$, ΔH = positive ($\Delta H > 0$)

For example, when nitrogen reacts with oxygen to form nitrogen dioxide, 66.4 kJ of heat energy is absorbed (*i.e.*, ΔH = + 66.4 kJ/mol) and thus the reaction is endothermic.

$$N_2\,(g) + 2O_2\,(g) \rightarrow 2NO_2\,(g);\ \Delta H = +\,66.4\ kJ/mol$$

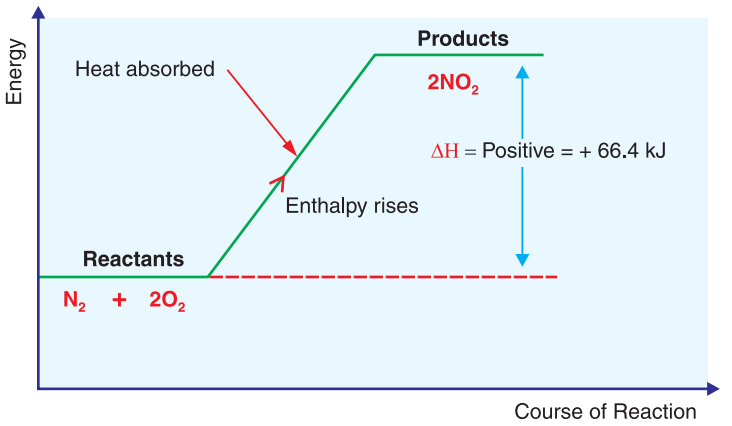


Figure 4.3 Energy diagram for an endothermic reaction.

For exothermic reactions, ΔH is negative because the energy of the reactants is greater than the energy of the products. Thus, the enthalpy of the system decreases, as shown in Figure 4.4.

$$H_p < H_r \Rightarrow \Delta H = \text{negative}\ (\Delta H < 0)$$

For example, when carbon burns in oxygen to produce carbon dioxide, 393.5 kJ of heat energy is liberated and hence the reaction is exothermic (ΔH = –393.5 kJ/mol).

$$C(s) + O_2(g) \rightarrow CO_2(g):\ \Delta H = -393.5\ kJ/mol$$

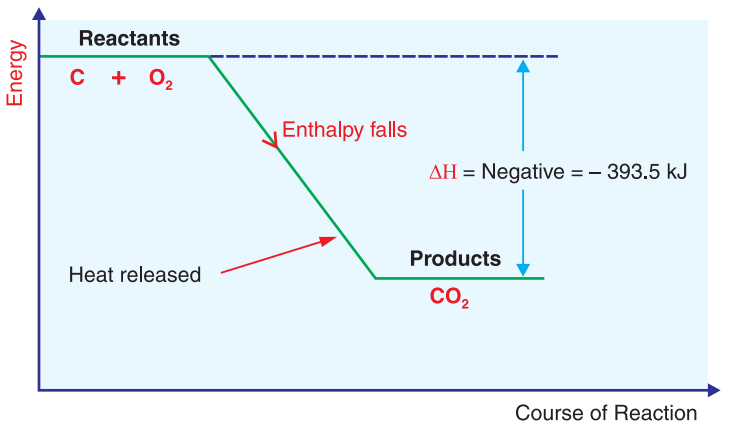


Figure 4.4 Energy diagram for an exothermic reaction.

Experiment 4.3



Investigating the Heat Involved in a Chemical Reaction

Objective: To determine the exothermic/endothermic nature of the reaction between sulphuric acid and sugar.

Apparatus: beaker and reagent bottle.

Chemicals: Concentrated H_2SO_4 and sugar.

Procedure:

1. Take small amount of sugar in a beaker.
2. Add a little concentrated sulphuric acid to the sugar.
3. Touch the outer surface of the beaker and record your observation.

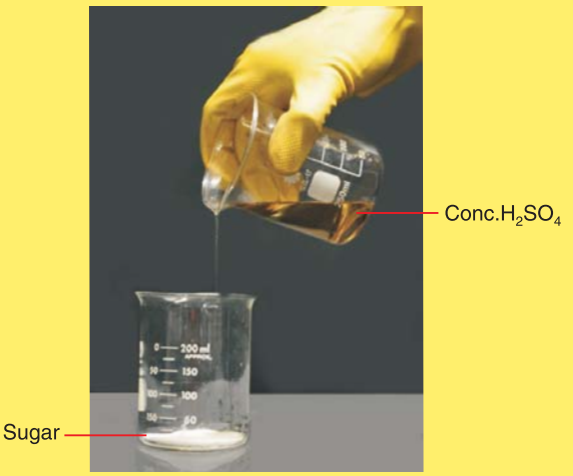


Figure 4.5 Reaction between sulphuric acid and sugar.

Observations and analysis

1. Does the beaker feel hot or cold when you touch it?
2. Did you see any steam in the beaker?
3. What is the colour of the product formed?
4. Write a balanced chemical equation.
5. What can you conclude from the experiment?

[Caution-When mixing concentrated acid and water, always add the acid to the water; never add water to concentrated acid.]

Experiment 4.4



Investigating the Heat Involved in a Reaction

Objective: To investigate the exothermic/endothermic nature of the process when ammonium nitrate is dissolved in water.

Apparatus: Beaker, thermometer, stirrer.

Chemicals: Ammonium nitrate and water.

***Procedure*:**

1. Take 100 mL of water in a beaker and record its temperature.
2. Dissolve 15 g of solid ammonium nitrate (NH_4NO_3) in the 100 mL of water.
3. Touch the outer surface of the beaker and record the temperature of the solution with the help of a thermometer.

Observations and analysis:

1. Does the beaker feel hot or cold when you touch it?
2. Is the temperature increased or decreased after the addition of NH_4NO_3?
3. What do you conclude from this experiment?

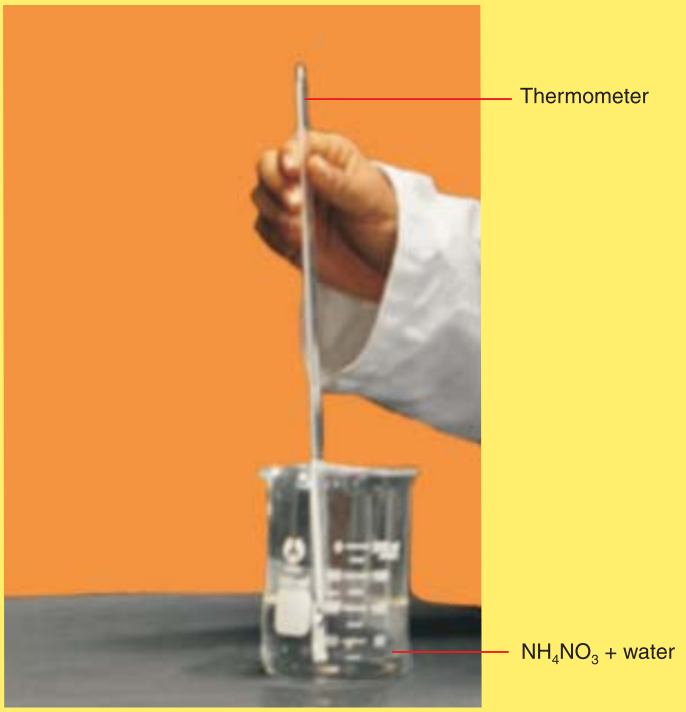


Figure 4.6 Dissolution of ammonium nitrate in water.

Activity 4.7



Form a group and perform the following task. In your daily life you encounter with many chemical changes involving energy. List some of such changes and discuss their importance.

Share your findings with the rest of the class.

Chemical reactions bring about chemical changes. All chemical changes are accompanied by energy changes. This energy is usually in the form of heat, light, or electricity.

Energy changes produced by chemical reactions have many practical applications (*uses*). For example, energy lifts rockets, runs cars, and extracts metal from compounds.

Many applications involve the energy produced by fuel combustion, which liberates large amounts of heat. The energy can be converted from one form to another. For example, the energy that fuel combustion produces can convert water to steam. The steam can run a turbine that creates electricity.

Respiration (breathing) creates energy for our bodies. Breathing releases the energy our living cells produce by oxidizing glucose. This energy helps to maintain our body temperature and body exercises.

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O + \text{Energy}$$

Exercise 4.4

In each of the following cases, determine the sign of ΔH. State whether the reaction is exothermic or endothermic, and draw an enthalpy diagram.

a $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) + 285.8\ kJ$

b $H_2O(I) + 40.7\ kJ \rightarrow H_2O(g)$

4.5 TYPES OF CHEMICAL REACTIONS

Competencies

By the end of this section, you will be able to:

- list the four types of chemical reactions;
- define combination reaction and give examples;
- conduct some experiments on combination reactions in groups;

- define decomposition reaction and give examples;
- conduct some experiments on decomposition reactions in group;
- define single displacement reactions and give examples;
- conduct some experiments on simple displacement reactions in groups;
- define double displacement reaction and give examples; and
- conduct some experiments on double displacement reactions in groups.

Activity 4.8



Form a group and discuss the following chemical reactions that occur during the:

a digestion of food in our body.

b fermentation of 'tej'.

c burning of kerosene in a stove.

Share your discussion with the rest of the class.

Chemical reactions are classified into four categories. These are combination, decomposition, single displacement and double displacement reactions.

i) Combination Reactions

Experiment 4.5



Investigation of Combination Reaction

Objective: To investigate the reaction between sulphur and iron.

Apparatus: Test tube, stand, burner, watch glass

Chemicals: Sulphur powder, iron filings

Procedure:

1. Mix about 3 g of iron filings and 2 g of powdered sulphur in a watch glass.
2. Transfer the mixture in a glass test tube.
3. Mount the test tube in a sloping position on a stand as shown in Figure 4.7.
4. Heat the test tube until the mixture in the glass glows red hot.
5. Remove the test tube from the flame and observe the result.

Observations and analysis:

1. What were the colours of iron filings and sulphur before the reaction?
2. What was the colour of the resulting compound after the reaction?
3. Write a balanced chemical equation for the reaction.
4. Identify the type of reaction.

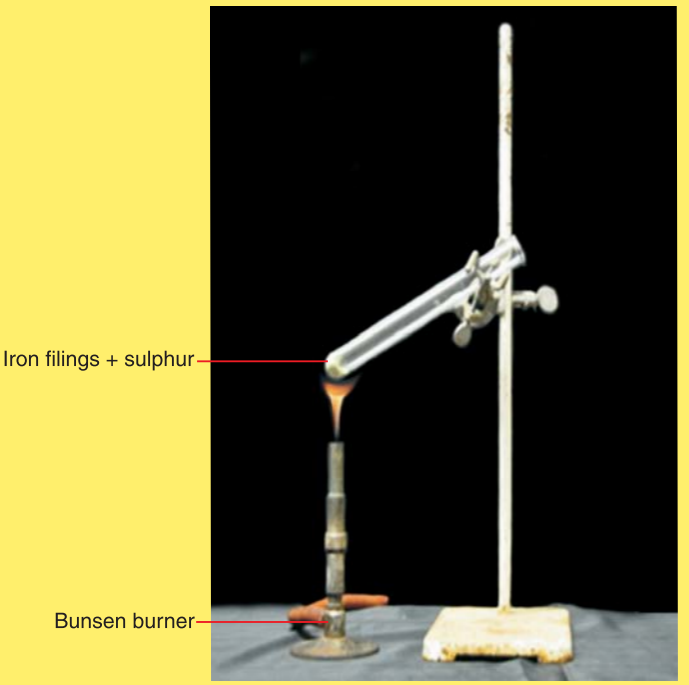


Figure 4.7 The reaction between iron and sulphur.

A reaction in which two or more substances combine to form a single substance is called a combination reaction. In a combination reaction, two elements, two compounds, or an element and a compound react to form a single compound. Combination reactions can be represented by the following general form of equation.

$$A + B \rightarrow AB;$$

where the reactants A and B are elements or compounds, the product AB is a compound.

Such type of reaction is also known as synthesis or composition reaction.

Examples

- Magnesium burns in oxygen to form magnesium oxide.

$$\underset{\text{element}}{2Mg} + \underset{\text{element}}{O_2} \rightarrow 2MgO$$

- Water and carbon dioxide combine to form carbonic acid.

$$H_2O + CO_2 \rightarrow H_2CO_3$$

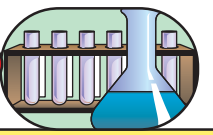
Exercise 4.5

Complete and balance the following combination reactions.

1. $CO + O_2 \rightarrow$
2. $H_2O + SO_2 \rightarrow$
3. $Na_2O + H_2O \rightarrow$
4. $CaO + CO_2 \rightarrow$

ii) Decomposition Reactions

Experiment 4.6



Investigation of Decomposition Reaction

Objective: To investigate the decomposition of copper (II) carbonate.

Apparatus: Test tube, stand, burner, cork, delivery tube.

Chemicals: Copper (II) carbonate and lime water.

Procedure:

Put copper (II) carbonate powder in a glass test-tube. Mount the test tube in a sloping position on a stand as shown in Figure 4.8. Fit a cork and a delivery tube to the test tube. Put another test tube containing lime water at the end of the delivery tube. Heat the copper (II) carbonate with a burner.

Observations and analysis:

1. What was the colour of copper (II) carbonate before heating?
2. What was the colour during heating and after cooling?
3. What change did you observe in the lime water?
4. Write a balanced chemical equation for the reaction.

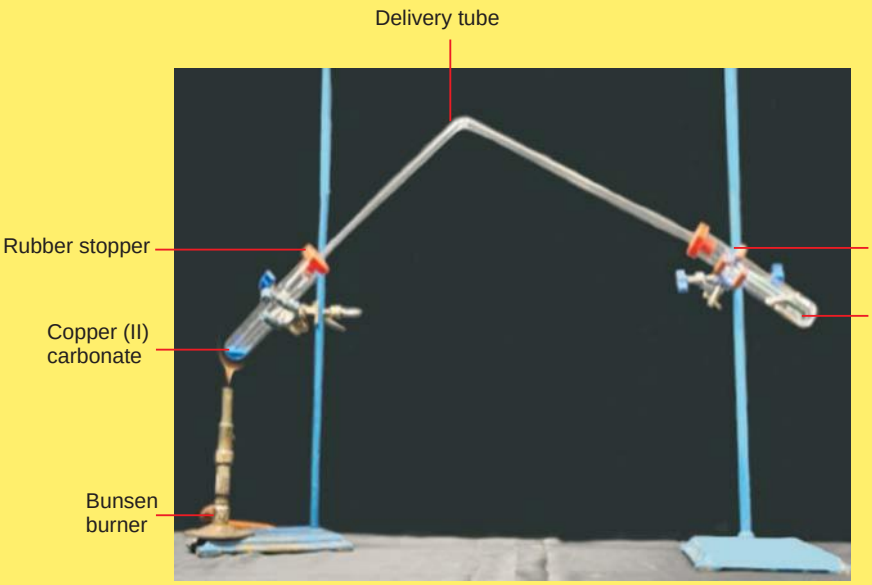


Figure 4.8 Decomposition of copper (II) carbonate.

A decomposition reaction is a reaction that involves the breaking down of a single compound into two or more elements or simpler compounds. A decomposition reaction can be carried out using heat, light, electricity or a catalyst. But most decomposition reactions are carried out when heat is supplied and this heat energy is indicated by a 'delta' (Δ) symbol above the arrow. The general form of equation for a decomposition reaction is:

$$AB \rightarrow A + B$$

where the reactant AB must be a compound and the products A and B could be elements or compounds.

Examples

- Water is decomposed to hydrogen and oxygen gases when electricity is passed through it.

$$2H_2O \xrightarrow[\text{current}]{\text{electric}} 2H_2 + O_2$$

- When sodium bicarbonate is heated, it decomposes to give sodium carbonate, carbon dioxide, and water.

$$2NaHCO_3 \xrightarrow{\Delta} Na_2CO_3 + CO_2 + H_2O$$

Let us consider the decompositions of nitrates and carbonates:

1. *Decomposition of Metallic Nitrates*

a) Decomposition of group IA nitrates produces nitrites and oxygen.

Examples

$$2NaNO_3 \xrightarrow{\Delta} 2NaNO_2 + O_2$$

$$2KNO_3 \xrightarrow{\Delta} 2KNO_2 + O_2$$

b) Decomposition of all metal nitrates, except group IA metals, gives nitrogen dioxide, metal oxide and oxygen gas.

Examples

$$2Ca(NO_3)_2 \xrightarrow{\Delta} 2CaO + 4NO_2 + O_2$$

$$2Zn(NO_3)_2 \xrightarrow{\Delta} 2ZnO + 4NO_2 + O_2$$

$$2Pb(NO_3)_2 \xrightarrow{\Delta} 2PbO + 4NO_2 + O_2$$

2. *Decomposition of Metallic Carbonates*

All metal carbonates, except sodium and potassium, decompose when heated to form the metal oxide and carbon dioxide.

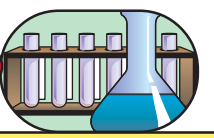
Examples

$$ZnCO_3 \xrightarrow{\Delta} ZnO + CO_2$$

$$CuCO_3 \xrightarrow{\Delta} CuO + CO_2$$

iii) *Single Displacement Reactions*

Experiment 4.7



Investigation of Single Displacement Reaction

Objective: To investigate the displacement reaction between iron and copper (II) sulphate.

Apparatus: Iron rod and beaker.

Chemicals: Copper (II) sulphate.

***Procedure*:**

1. Clean a piece of iron rod or iron knife with emery paper to remove any rust.
2. Take copper sulphate solution in a beaker.
3. Dip the iron rod into the copper (II) sulphate solution as shown in Figure 4.9 and wait for a few minutes. What did you observe on the iron rod?
4. Allow the reactants to stand for one day and observe any change on the iron rod.

Observations and analysis:

1. What did you observe on the iron rod after one day?
2. Write a balanced chemical equation for the reaction.
3. Write the conclusion for the experiment.

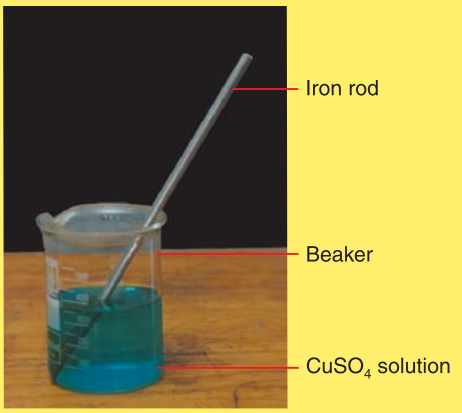


Figure 4.9 Reaction between iron and copper (II) sulphate.

A reaction in which one element displaces another element from its compound is known as single displacement or replacement reaction. Such a reaction is represented by the following two general forms.

$$A + BC \rightarrow B + AC$$

If A is a metal, it will displace B to form AC, provided A is a more active metal than B.

$$A + BC \rightarrow BA + C$$

If A is a non-metal, it will displace C to form BA, provided A is a more active non-metal than C.

In general, a more reactive element displaces a less reactive element from a compound.

Examples of single-displacement reactions

a Active metals displace hydrogen from acids.

Reactive metals such as potassium, calcium, sodium, and zinc displace hydrogen gas from dilute acids.

For example, zinc is an active metal, and it displaces hydrogen from hydrochloric acid; but copper metal cannot do so.

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$
$$Au + HCl \rightarrow \textit{No reaction}$$

b Reactive metals, such as potassium, calcium, and sodium react vigorously with water to displace hydrogen:

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$
$$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$$

Exercise 4.6

Complete and balance the following single displacement reactions:

1. $Zn + CuSO_4 \longrightarrow$
2. $Cu + Zn(NO_3)_2 \longrightarrow$
3. $F_2 + CaCl_2 \longrightarrow$
4. $Br_2 + NaCl \longrightarrow$

iv) Double Displacement Reactions

Experiment 4.8



Investigation of Double Displacement Reaction

Objective: To observe the displacement reaction between Na_2SO_4 and $Ba(NO_3)_2$.

Apparatus: Beaker, stirrer, filter paper, filter funnel.

Chemicals: Na_2SO_4 and $Ba(NO_3)_2$.

Procedure:

1. Take solution of $Ba(NO_3)_2$ into a beaker and add dropwise Na_2SO_4 solution. Then stir it continuously.
2. Filter the precipitate using a filter paper and funnel. Collect the filtrate or the solution in a clean beaker.

Observations and analysis:

1. Write the names of the compounds that are formed as a precipitate and as solution at the end of the reaction.
2. What was the colour of the precipitate.
3. Write the balanced chemical equation for the reaction.

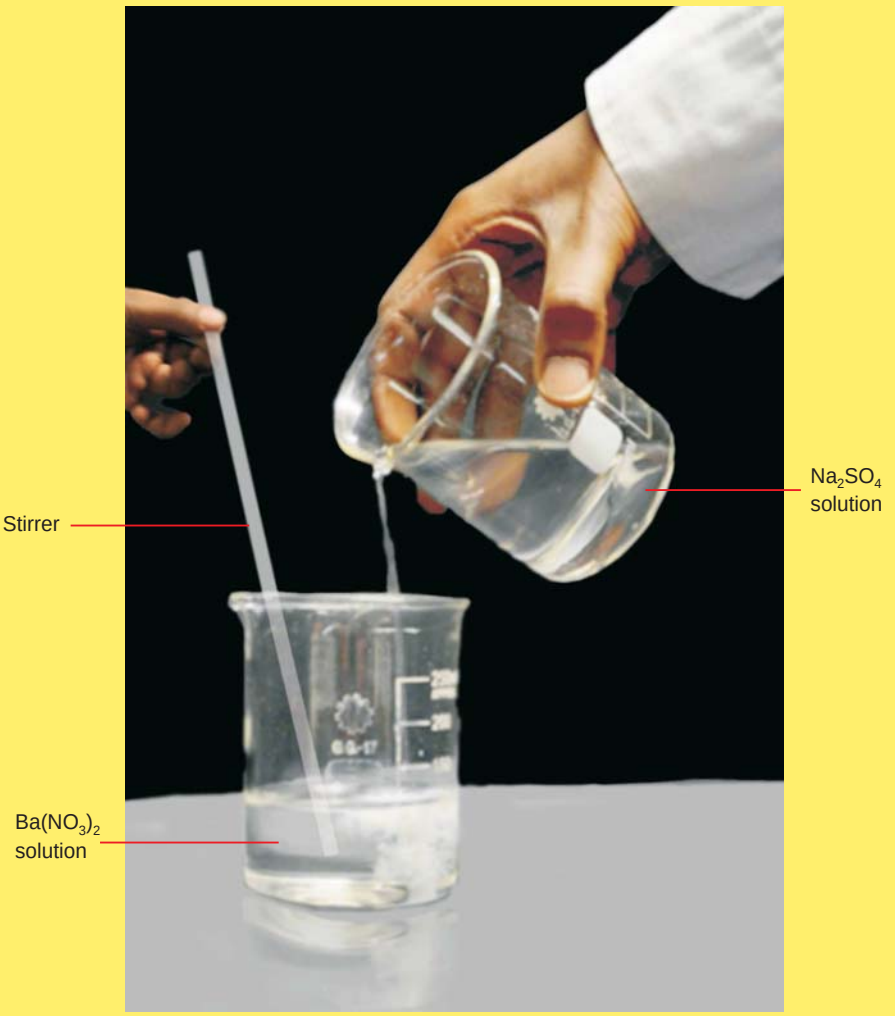


Figure 4.10 The double displacement reaction between Na_2SO_4 and $Ba(NO_3)_2$.

Double displacement reaction is a reaction in which two compounds react together to form two new compounds by exchange of the positive and negative ions of each reactant. Such a reaction is also known as double replacement reaction or metathesis.

This type of reaction can be written in the following general form of equation.

$$AB + CD \rightarrow AD + CB$$

Examples

- The two soluble compounds $AgNO_3$ and NaCl react to produce an insoluble precipitate of AgCl and a soluble $NaNO_3$ solution.

$$AgNO_3 + NaCl \rightarrow \underset{\textit{Insoluble}}{AgCl\downarrow} + \underset{\textit{Soluble}}{NaNO_3}$$

- When aqueous solutions of $BaCl_2$ and Na_2SO_4 react, a precipitate of $BaSO_4$ is formed.

$$BaCl_2 + Na_2SO_4 \rightarrow \underset{\textit{Insoluble}}{BaSO_4\downarrow} + \underset{\textit{Soluble}}{2NaCl}$$

Exercise 4.7

Give appropriate answers for the following questions.

1. What type of reaction does usually take place in each of the following reactions?
   - a a metal reacting with water.
   - b a metal reacting with a non-metal.
   - c an acid reacting with a metal hydroxide.
   - d heating of a metal hydrogen carbonate.
2. Classify the following reactions as combination, decomposition, single or double displacement reactions.
   - a $FeO + C \rightarrow Fe + CO$
   - b $2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$
   - c $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$
   - d $2Cu(NO_3)_2 \rightarrow 2CuO + 4NO_2 + O_2$
   - e $2Na_3PO_4 + 3Ca(OH)_2 \rightarrow Ca_3(PO_4)_2 + 6NaOH$
   - f $CuSO_4.5H_2O \rightarrow CuSO_4 + 5H_2O$

3. Complete and balance the following equations. If the reaction does not take place, write "No Reaction"

   a $Mg + N_2 \xrightarrow{\Delta}$

   b $Na_2CO_3 \xrightarrow{\Delta}$

   c $BaCO_3 + HNO_3 \rightarrow$

   d $Zn + H_2SO_4 \rightarrow$

   e $FeCO_3 \xrightarrow{\Delta}$

   f $H_2CO_3 + NaOH \rightarrow$

4.6 STOICHIOMETRY

Competencies

By the end of this section, you will be able to:

- deduce mole ratios from balanced chemical equations;
- solve mass-mass problems based on the given chemical equation;
- define molar volume;
- state Avogadro's principle;
- solve volume-volume problems based on the given chemical equation;
- solve mass-volume problems based on the given chemical equation;
- define limiting and excess reactants;
- determine limiting and excess reactants of a given chemical reaction;
- show that the amount of product formed in a chemical reaction is based on the limiting reactant;
- define the term theoretical yield, actual yield and percentage yield; and
- calculate the percentage yield of a chemical reaction from given information.

Activity 4.9



Form a group and discuss the following concepts:

a A bicycle mechanic has 10 frames (body parts) and 16 wheels in the shop. How many complete bicycles can he assemble using these parts? Which parts of the bicycle are left over?

b Based on your conculsion in (a), do you think that the masses of reactants are always completely converted to products in a chemical reaction?

Present your conclusion to the class.

The quantitative relationship between reactants and products in a balanced chemical equation is known as stoichiometry. In other words, stoichiometry is the study of the amount or ratio of moles, mass, energy and volumes (for gases) of reactants and products. Stoichiometric calculations are based on the following two major principles.

i) The composition of any substance in the chemical equation should be expressed by a definite formula.

ii) The law of conservation of mass must be obeyed (the mass of reactants equals the mass of products).

4.6.1 Molar Ratios in Balanced Chemical Equation

From a balanced chemical equation, it is possible to determine the:

- number of moles of each reactant and product; and
- relative mass of each of the reactants and products

For example, in the reaction of hydrogen with oxygen to produce water, 2 moles of H_2 combines with 1 mole of O_2 to yield 2 moles of H_2O. The equation also tells us 4 g of hydrogen reacts with 32 g of oxygen to produce 36 g of water. This can be further interpreted as follows:

	$2H_2(g)$	+	$O_2(g)$	→	$2H_2O(l)$
Mole →	2 mole		1 mole		2 mole
Molecule →	2 molecule		1 molecule		2 molecule
Mass →	4 g		32 g		2 × 18 g
	36 g reactant				36 g product

Calculations based on chemical equations (*stoichiometric problems*) are classified into mass-mass problem, volume-volume problems and mass-volume problems.

4.6.2 Mass–Mass Relationships

In mass-mass problems, the mass of one substance is given, and the mass of the second substance is determined from the same reaction. There are two methods for solving such types of problems:

i) Mass-ratio method

ii) Mole-ratio method

Let us see each method by using the necessary steps.

i) The mass-ratio method

In this type of stoichiometric calculation, the mass of one substance is determined from the given mass of the other substance using the following steps.

Step 1: Write the balanced chemical equation.

Step 2: Place the given mass above the corresponding formula, and x above the formula of the substance whose mass is to be determined.

Step 3: Write the total molar mass of the substances below the formula of each substance. (*Total molar mass is the molar mass of the substance multiplied by its coefficient*).

Step 4: Set up the proportion.

Step 5: Solve for the unknown mass, x.

Example 1

How many grams of calcium chloride are formed when 15 g of calcium metal reacts with hydrochloric acid?

Solution:

Step 1: $Ca + 2HCl \rightarrow CaCl_2 + H_2$

Step 2: $\overset{15\ g}{Ca} + 2HCl \rightarrow \overset{x}{CaCl_2} + H_2$

Step 3: $\underset{40\ g}{\overset{15\ g}{Ca}} + 2HCl \rightarrow \underset{111g}{\overset{x}{CaCl_2}} + H_2$

Step 4: $\dfrac{15\ g}{40\ g} = \dfrac{x}{111\ g}$

Step 5: $x = 41.63$ g

Therefore, 41.63 g of $CaCl_2$ is produced.

Example 2

How many grams of oxygen are produced by the decomposition of 145 grams of potassium chlorate?

Solution:

Step 1: $2KClO_3 \rightarrow 2KCl + 3O_2$

Step 2: $\overset{145\,g}{2\,KClO_3} \rightarrow 2KCl + 3\overset{x}{O_2}$

Step 3: $\underset{(2\times122.5)\,g}{\overset{145\,g}{2\,KClO_3}} \rightarrow 2KCl + \underset{(3\times32)\,g}{\overset{x}{3\,O_2}}$

Step 4: $\frac{145\,g}{245\,g} = \frac{x}{96\,g}$

Step 5: $x = 56.8\ g$

ii) The mole-ratio method

The mole ratio is the ratio between the numbers of moles of any two substances in a given reaction. In this method, the given mass is converted into moles, and the number of moles for the required substance is calculated. If needed, convert the obtained moles back to mass.

Follow the steps given below to solve problems of mass-mass relationships by the mole ratio method:

Step 1: Write the balanced chemical equation.

Step 2: Convert the given mass to moles and write the obtained moles and the required quantity, *x*, above the formulas of the respective substances.

Step 3: Place the coefficients as the number of moles under the formula of each substance involved.

Step 4: Set up the proportion.

Step 5: Solve for the unknown value, *x*; and convert the moles obtained into mass.

Example 3

How many grams of sodium metal are needed to react with 10.0 g of water?

Solution:

Step 1: $2Na + 2H_2O \rightarrow 2NaOH + H_2$

Step 2: moles of $H_2O = \frac{\text{given mass}}{\text{molar mass}} = \frac{10.0\,g}{18\,g/mol} = 0.56\ mol$

$\overset{x}{2\,Na} + \overset{0.56\,mol}{2H_2O} \rightarrow 2NaOH + H_2$

Step 3: $\overset{x}{\underset{2\,mol}{2Na}} + \overset{0.56\,mol}{\underset{2\,mol}{2H_2O}} \rightarrow 2NaOH + H_2$

Step 4: $\frac{x}{2\,mol} = \frac{0.56\,mol}{2\,mol}$

Step 5: x = 0.56 mol of Na

Now, convert 0.56 mole of Na to grams

mass of Na = mole × molar mass

= 0.56 mol × 23 g/mol

= 12.88 g

Therefore, 12.88 g of sodium metal is needed to react with 10 g of water.

Example 4

What mass of nitrogen dioxide is produced by the decomposition of 182 g of magnesium nitrate?

Solution:

Step 1: $2Mg(NO_3)_2 \rightarrow 2MgO + 4NO_2 + O_2$

Step 2: moles of $Mg(NO_3)_2 = \frac{182\,g}{148\,g/mol} = 1.23$ mol

$$\overset{1.23\,mol}{2\,Mg(NO_3)_2} \rightarrow 2MgO + 4\overset{x}{NO_2} + O_2$$

Step 3: $\overset{1.23\,mol}{\underset{2\,mol}{2\,Mg(NO_3)_2}} \rightarrow 2MgO + 4\overset{x}{\underset{4\,mol}{NO_2}} + O_2$

Step 4: $\frac{1.23\,mol}{2\,mol} = \frac{x}{4\,mol}$

Step 5: x = 2.46 moles of NO_2

Mass of NO_2 = 2.46 mol × 46 g/mol = 113.2 g

Therefore, 113.2 g of NO_2 is produced.

Exercise 4.8

1. How many grams of $CaCO_3$ are needed to react with 15.2 g of HCl in according to the following equation?

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$$

2. How many grams of NaOH are needed to neutralize 50 grams of H_2SO_4?
3. Calculate the mass of $CaCl_2$ formed when 5 moles of chlorine reacts with calcium metal.
4. How many moles of H_2O are required to produce 4.5 moles of HNO_3 according to the following reaction:

$$3NO_2 + H_2O \rightarrow 2HNO_3 + NO$$

5. In the decomposition of $KClO_3$, how many moles of KCl are formed in the reaction that produces 0.05 moles of O_2?
6. How many moles of CaO are needed to react with excess water to produce 370 g of calcium hydroxide?

4.6.3 Volume-Volume Relationships

In reactions involving gases, the volume of gases can be determined on the principle that 1 mole of any gas occupies a volume of 22.4 litres at STP (*standard temperature* and *pressure*, STP, the *temperature is* 0°C and *the pressure is 1 atm*). It is also known that 22.4 L of any gas weighs exactly its molecular mass at STP. This volume, 22.4 litres, of a gas is known as the molar volume.

At STP, 1 mole of any gas = 22.4 L = gram volume mass of the gas

The relationship between the volume of a gas and its number of molecules was explained by Avogadro. Avogadro's law states that equal volumes of different gases, under the same conditions of temperature and pressure, contain equal number of molecules. This law can also be stated as the volume of a gas is proportional to the number of molecules (moles) of the gas at STP.

Mathematically, $V \propto n$; where V is the volume and n is the number of moles.

In volume-volume problems, the volume of one substance is given and the volume of the other substance is calculated. All the steps to solve volume-volume problems are shown by the following example.

Example 5

What volume of oxygen will react with carbon monoxide to produce 20 litres of carbon dioxide at STP?

Solution:

Step 1: Write the balanced chemical equation

$$2CO + O_2 \rightarrow 2CO_2$$

Step 2: Place the given volume and the required volume, x, above the corresponding formulas.

$$2CO + \overset{x}{O_2} \rightarrow \overset{20\,L}{2CO_2}$$

Step 3: Write the total molar volume (22.4 L multiplied by any coefficient) below the formulas.

$$N_2(g) + 2H_2O(g) + \text{heat} \rightleftharpoons 2NO(g) + 2H_2(g)$$

$$2CO + \underset{22.4\,L}{\overset{x}{O_2}} \rightarrow \underset{2(22.4\,L)}{\overset{20\,L}{2CO_2}}$$

Step 4: Set up the proportion.

$$\frac{x}{22.4\,L} = \frac{20\,L}{44.8\,L}$$

Step 5: Solve for the unknown volume, x.

x = 10 L of O_2 are needed.

Exercise 4.9

1. What volume of nitrogen reacts with 33.6 litres of oxygen to produce nitrogen dioxide?
2. How many litres of sulphur trioxide are formed when 4800 cm^3 of sulphur dioxide is burned in air?
3. How many litres of ammonia are required to react with 145 litres of oxygen according to the following reaction?

$$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$$

4. Calculate the volume of oxygen produced in the decomposition of 5 moles of $KClO_3$ at STP?
5. How many moles of water vapour are formed when 10 litres of butane gas, C_4H_{10} is burned in oxygen at STP?

4.6.4 Mass–Volume Relationships

In mass-volume problems, either the mass of one substance is given and the volume of the other is required or the volume of one substance is given and the mass of the other one is required. The steps to solve such type of problems are the same as the previous steps except putting the masses on one side and the volumes on the other side of the equality sign.

Example 6

How many grams of calcium carbonate are decomposed to produce 11.2 L of carbon dioxide at STP?

Solution:

Step 1: $CaCO_3 \rightarrow CaO + CO_2$

Step 2: $\overset{x}{CaCO_3} \rightarrow CaO + \overset{11.2\,L}{CO_2}$

Step 3: $\underset{100\,g}{\overset{x}{CaCO_3}} \rightarrow CaO + \underset{22.4\,L}{\overset{11.2\,L}{CO_2}}$

Step 4: $\dfrac{x}{100\,g} = \dfrac{11.2\,L}{22.4\,L}$

Step 5: $x = 50$ g of $CaCO_3$ is decomposed.

Example 7

How many litres of oxygen at STP react with 72 g of aluminum to produce aluminum oxide?

***Solution*:**

Step 1: $4Al + 3O_2 \rightarrow 2Al_2O_3$

Step 2: $\overset{72\,g}{4Al} + \overset{x}{3O_2} \rightarrow 2Al_2O_3$

Step 3: $\underset{(4\times27)\,g}{\overset{72\,g}{4Al}} + \underset{(3\times22.4)\,L}{\overset{x}{3O_2}} \rightarrow 2Al_2O_3$

Step 4: $\dfrac{72\,g}{108\,g} = \dfrac{x}{67.2\,L}$

Step 5: $x = 44.8$ L of O_2

Hence, 44.8 litres of oxygen is required at STP to react with 72 g of aluminium.

Exercise 4.10

1. How many litres of oxygen are required to react with 23 g of methane according to the following equation?

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

2. What mass of aluminium would be completely oxidized by 44.8 L of oxygen to produce Al_2O_3 at STP?

3. Calculate the mass of calcium carbide that is needed to produce 100 cm^3 of acetylene according to the following equation.

$$CaC_2 + 2H_2O \rightarrow C_2H_2 + Ca(OH)_2$$

4. How many millilitres of sulphur dioxide are formed when 12.5 g of iron sulphide ore (pyrite) reacts with oxygen according to the equation at STP?

$$4FeS_2 + 11O_2 \rightarrow 2Fe_2O_3 + 8SO_2$$

4.6.5 Limiting and Excess Reactants

When all the reactants are completely consumed in a chemical reaction, then such reactants are said to be in stoichiometric proportions. But, practically these types of chemical reactions do not always occur. In many cases, an excess of one or more

reactants is encountered in the reaction and the other reactant is completely converted into products. Thus, the reactant that is completely consumed in the reaction is known as the limiting reactant, because it limits or determines the amount of products that can be formed.

For example, consider the following reaction:

$$\underset{1\text{ mole}}{H_2} + \underset{1\text{ mole}}{Cl_2} \rightarrow \underset{2\text{ moles}}{2HCl}$$

According to the equation, 1 mole of H_2 reacts with 1 mole of chlorine to produce 2 mole of HCl. Thus, all the reactants are completely consumed and only products appear. However if 1 mole of H_2 reacts with 1.5 mole of Cl_2, there is insufficient H_2 to react with all of the Cl_2. Therefore, Cl_2 will be in excess and H_2 will be the limiting reactant. Only 2 moles of HCl are formed and at the end of the reaction 0.5 mole of Cl_2 remains unreacted.

Example 8

How much ammonia is produced if 10 g of hydrogen reacts with 18 g of nitrogen?

Solution:

$$3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$$

First determine the number of moles;

$$\text{Moles of } H_2 = \frac{10\text{ g}}{2\text{ g/mol}} = 5\text{ mol}$$

$$\text{Moles of } N_2 = \frac{18\text{ g}}{28\text{ g/mol}} = 0.64\text{ mol}$$

Now, calculate the number of moles or masses of the product that would be formed by each reactant.

The reactant that gives the smallest amount of product is the limiting reactant.

i. **Using the quantity of H_2**

$$\overset{5\text{ mol}}{\underset{3\text{ mol}}{3H_2}} + N_2 \rightarrow \overset{x}{\underset{2\text{ mol}}{2NH_3}}$$

$$\frac{5\text{ mol}}{3\text{ mol}} = \frac{x}{2\text{ mol}}$$

$x = 3.33\text{ mol } NH_3$

$\text{Mass of } NH_3 = 3.33\text{ mol} \times 17\text{ g/mol}$

$= 56.6\text{ g}$

ii. **Using the quantity of N_2**

$$3H_2 + \overset{0.64\text{ mol}}{\underset{1\text{ mol}}{N_2}} \rightarrow \overset{x}{\underset{2\text{ mol}}{2NH_3}}$$

$$\frac{0.64\text{ mol}}{1\text{ mol}} = \frac{x}{2\text{ mol}}$$

$x = 1.28\text{ mol } NH_3$

$\text{Mass of } NH_3 = 1.28\text{ mol} \times 17\text{ g/mol}$

$= 21.8\text{ g}$

Therefore, the limiting reactant is nitrogen, because it gives less amount of NH_3, *i.e.*, 21.8 g NH_3. In the reaction, 0.64 mole (18 g) of N_2 is consumed. Hydrogen is in excess. The amount of hydrogen consumed will be:

$$\overset{x}{3H_2} + \underset{28\,g}{\overset{18\,g}{N_2}} \rightarrow 2NH_3 \qquad \text{or} \qquad \overset{x}{3H_2} + \underset{1\,mole}{\overset{0.64\,mol}{N_2}} \rightarrow 2NH_3$$

$$\frac{x}{6} = \frac{18}{28} \qquad \frac{x}{3\,mol} = \frac{0.64\,mol}{1\,mol}$$

x = 3.86 g of H_2 $\qquad$ x = 1.92 mol of H_2

Therefore, 3.86 g or 1.92 moles of H_2 is used in the reaction, and 6.14 g or 3.08 moles of H_2 is left unreacted.

Example 9

In the chemistry laboratory, a student performed a displacement reaction by adding 9.5 g of zinc into 9.5 g of HCl in a beaker. What weight of $ZnCl_2$ will be produced?

Solution:

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

$$\text{Moles of Zn} = \frac{9.5\,g}{65\,g/mol} = 0.15\ mol$$

$$\text{Moles of HCl} = \frac{9.5\,g}{36.5\,g/mol} = 0.26\ mol$$

Even though the given masses of the two reactants are the same, they are not mixed in equimolar ratio as shown above. Thus, the limiting reactant must be determined first.

i. **Using the quantity of Zn**

$$\underset{1\,mol}{\overset{0.15\,mol}{Zn}} + 2HCl \rightarrow \underset{1\,mol}{\overset{x}{ZnCl_2}} + H_2$$

$$\frac{0.15\,mol}{1\,mol} = \frac{x}{1\,mol}$$

x = 0.15 mol $ZnCl_2$

ii. **Using the quantity of HCl**

$$Zn + \underset{2\,mol}{\overset{0.26\,mol}{2HCl}} \rightarrow \underset{1\,mol}{\overset{x}{ZnCl_2}} + H_2$$

$$\frac{0.26\,mol}{2\,mol} = \frac{x}{1\,mol}$$

x = 0.13 mol $ZnCl_2$

Hence, the limiting reactant is HCl.

Mass of $ZnCl_2$ = 0.13 mol × 136 g/mol = 17.68 g $ZnCl_2$

Exercise 4.11

1. If 6.5 g of zinc reacts with 5.0 g of HCl, according to the following reaction.

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

   a Which substance is the limiting reactant?

   b How many grams of the reactant remains unreacted?

   c How many grams of hydrogen would be produced?

2. What mass of Na_2SO_4 is produced if 49 g of H_2SO_4 reacts with 80 g of NaOH?

3. If 20 g of $CaCO_3$ and 25 g of HCl are mixed, what mass of CO_2 is produced?

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$$

4. If 3 moles of calcium reacts with 3 moles of oxygen, then

   a Which substance is the limiting reactant?

   b How many moles of calcium oxide are formed?

5. For the reaction:

$$2Al + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3H_2$$

   How many grams of hydrogen are produced if 0.8 mole of aluminium reacts with 1.0 mole of sulphuric acid?

4.6.6 Theoretical, Actual and Percentage Yields

The measured amount of product obtained in any chemical reaction is known as the actual yield. The theoretical yield is the calculated amount of product that would be obtained if the reaction proceeds completely. The actual yield (*experimentally determined yield*) of a product is usually less than the theoretical yield (*calculated yield*).

The percentage yield is the ratio of the actual yield to the theoretical yield multiplied by 100.

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

Example 10

25 grams of methane gas (CH_4) burns in oxygen according to the following reaction:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

What is the percentage yield if 60.3 grams of carbon dioxide is produced?

Solution:

The actual yield is 60.3 g of CO_2.

Determine the theoretical yield using mass-mass relationship

$$\overset{25\,g}{\underset{16\,g}{CH_4}} + 2O_2 \rightarrow \overset{x}{\underset{44\,g}{CO_2}} + 2H_2O$$

$$\frac{25\,g}{16\,g} = \frac{x}{44}$$

x = 68.75 g of CO_2 (*theoretical yield*)

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

$$= \frac{60.3\,g}{68.75\,g} \times 100\% = 87.7\ \%$$

Exercise 4.12

1. When 20 g of sulphur dioxide reacts with oxygen, 23 g of sulphur trioxide is formed. What is the percentage yield?
2. When 14.5 g of SO_2 reacts with 21 g of O_2, what will be the theoretical yield and percentage yield of the reaction if the actual yield is 12 g?
3. In the reaction:

$$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$$

When 52.7 g of octane (C_8H_{18}) burns in oxygen, the percentage yield of carbon dioxide is 82.5%. What is the actual yield in grams?

4.7 OXIDATION AND REDUCTION REACTIONS

Competencies

***By the end of this section, you will be able to*:**

- define redox reactions;
- define the terms oxidation and reduction in terms of electron transfer;
- define oxidation number (oxidation state),
- state oxidation number rules,
- determine the oxidation number of an element in a given formula;
- describe the oxidizing and reducing agents;
- analyze a given redox reaction by specifying the substance reduced and the substance oxidized, and also the oxidizing and reducing agents; and
- Distinguish between redox and non-redox reactions.

Activity 4.10



Form a group and discuss the following phenomenon:

When you dry your meal dishes with a towel, the towel can be termed as a drying agent and the dish as wetting agent. What happens to the towel and the dish, in terms of getting wet and dry, after cleaning? Relate this idea with oxidizing agent and reducing agent, oxidation and reduction of substances. Present your conclusion to the class.

In our day to day activity, we are familiar with the chemical processes like rusting of iron, burning of substances, breathing of air, digestion of food and so on. All such types of processes or reactions are known as oxidation and reduction or redox reactions.

4.7.1 Oxidation-Reduction

Can oxidation take place without reduction?

Oxidation: is the process in which a substance loses electrons in a chemical reaction.

For example, in the reaction

$$2Na + Cl_2 \rightarrow 2NaCl$$

Each sodium atom has lost one electron and has turned to a sodium ion. Hence, sodium is oxidized.

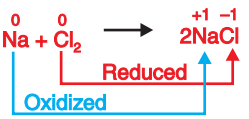
$$Na + 1e^- \rightarrow Na^+$$

Reduction: is the process in which a substance gains electrons in a chemical reaction.

For example, in the above reaction each chlorine atom has gained an electron and has changed to chloride ion. Thus, chlorine is reduced;

$$Cl_2 + 2e^- \rightarrow Cl^-$$

The processes of oxidation and reduction always occur simultaneously because if one substance loses electrons, the other substance must gain these electrons. Since the process of oxidation and reduction involves the transfer of electrons, it also results in the changes of oxidation number. Thus, oxidation and reduction can also be defined in terms of oxidation number. Oxidation is an increase in the oxidation number of an element and reduction is a decrease in the oxidation number. For example, in the reaction,



The oxidation number of sodium is increased from 0 to +1 and thus sodium is oxidized. The oxidation number of chlorine is decreased from 0 to –1, and therefore chlorine is reduced.

4.7.2 Oxidation Number or Oxidation State

Oxidation number or oxidation state is the number of electrons that an atom appears to have gained or lost when it is combined with other atoms.

Rules for Assigning Oxidation Numbers

Rule 1: The oxidation number of all elements in free state is zero. This rule is also applied for diatomic or polyatomic elements.

Example: The oxidation number of Na = 0, Cu = 0, Cl in Cl_2 = 0, O in O_3 = 0, S in S_8 = 0.

Rule 2: The oxidation number of a monatomic ion is equal to the charge on the ion.

Example: $Na^+ = +1$, $Mg^{2+} = +2$, $S^{2-} = -2$.

Rule 3: The oxidation number of oxygen in a compound is usually – 2 except in the following cases:

Exceptions

The oxidation number of oxygen in:

i) peroxides is –1. **Example:** Na_2O_2

ii) superoxides is –1/2. **Example:** KO_2

iii) oxygen diflouride is +2. **Example:** OF_2

Rule 4: The oxidation number of hydrogen in its entire compounds is +1 except in metal hydrides, (like NaH, CaH_2 and AlH_3), where its oxidation number is –1.

Rule 5: The sum of the oxidation number of all the atoms in a neutral compound is zero.

Example: $\overset{+1}{H_2}\overset{+6}{S}\overset{-2}{O_4}$ ((+2) + (+6) + (–8) = 0)

Rule 6: In a polyatomic ion, the sum of the oxidation numbers of the constituent atoms equals the charge on the ion.

Example: $(\overset{+6}{S}\overset{-2}{O_4})^{-2}$ ((+6) + (–8) = –2)

Rule 7: Elements of group IA have +1 and group IIA have +2 oxidation states in all of their compounds.

Rule 8: In a compound, the more electronegative element is assigned a negative oxidation number, and the less electronegative element is assigned a positive oxidation number.

Example: $\overset{+3}{N}\overset{-1}{Cl_3}$ (*chlorine is more electronegative than nitrogen*)

Example 1

What is the oxidation number of chromium in $Na_2Cr_2O_7$?

Solution:

The oxidation number of O is –2 (*Rule 3*)

The oxidation number of Na is +1 (*Rule 7*)

Let the oxidation number of Cr be *x*.

$\overset{+1}{Na_2}\ \overset{x}{Cr_2}\ \overset{-2}{O_7}$

Since the sum of the oxidation numbers of Na, Cr, and O in $Na_2Cr_2O_7$ is 0 (*Rule 5*)

Then, $\overset{+1}{Na}_2\ \overset{x}{Cr}_2\ \overset{-2}{O}_7$

$$(1 \times 2) + (x \times 2) + (-2 \times 7) = 0$$
$$2 + 2x - 14 = 0$$
$$x = +6$$

Therefore, the oxidation number of Cr in $Na_2Cr_2O_7$ is +6.

Example 2

What is the oxidation number of manganese in MnO_4^-?

Solution:

Let the oxidation number of Mn be *x*.

$\left(\overset{x}{Mn}\ \overset{-2}{O}_4\right)^-$

The sum of the oxidation numbers of Mn and O in MnO_4^- is –1 (*Rule 6*)

$$x + (-2 \times 4) = -1$$
$$x - 8 = -1$$
$$x = +7$$

Therefore, the oxidation number of Mn in MnO_4^- is +7.

Example 3

Determine the oxidation number of phosphorus in $Ca(H_2PO_4)_2$.

Solution:

The oxidation number of Ca is +2 (*Rule 7*).

Let, the oxidation number of P be *x*.

$$\overset{+2}{Ca}\left(\overset{+1}{H}_2\ \overset{x-2}{PO}_4\right)_2$$

$$+2 + (4 \times (+1)) + (2 \times x) + (8 \times (-2)) = 0$$
$$2 + 4 + 2x - 16 = 0$$
$$2x - 10 = 0 \quad \text{or} \quad x = +5$$

Hence, the oxidation number of P in $Ca(H_2PO_4)_2$ is +5.

Exercise 4.13

1. Determine the oxidation number of the specified element in each of the following:

   a C in $H_2C_2O_4$ b N in NH_4F

   c S in $Na_2S_4O_6$ d P in $Ca_3(PO_4)_2$

   e H in AlH_3 f N in NH_4HCO_3

   g Fe in $K_4[Fe(CN)_6]$

2. Determine the oxidation number of the specified element in each of the following.

   a S in S^{-2} b Cl in ClO_3^-

   c N in NH_4^+ d P in PO_4^{-3}

   e Cr in $Cr_2O_7^{-2}$ f S in $S_2O_8^{-2}$

3. Determine whether the following processes are oxidation or reduction reactions:

   a $Cu^2 + 2e^- \rightarrow Cu$ b $K \rightarrow K^+ + e^-$

   c $O + 2e^- \rightarrow O^{2-}$ d $S^{2-} \rightarrow S + 2e^-$

   e $Fe^{2+} \rightarrow Fe^{3+} + e^-$ f $N + 3e^- \rightarrow N^{3-}$

4.7.3 Oxidizing and Reducing Agents

In a redox reaction, the substance that causes another substance to get oxidized, but itself gets reduced, is known as an oxidizing agent, or oxidant. In the same manner, the substance that causes another substance to get reduced, but itself oxidized, is referred to as a reducing agent or reductant.

- Oxidizing agents are substances that:
  - are reduced (*gain electrons*)
  - contain elements whose oxidation number decreases
- Reducing agents are substances that:
  - are oxidized (*lose electrons*)
  - contain elements whose oxidation number increases

Tests for an oxidizing agent are accomplished by mixing it with a substance which is easily oxidized to give a visible colour change when the reaction takes place.

For example,

i) Permanganate ion (MnO_4^-) in acidic solution changes colour from purple to colourless.

$$MnO_4^- \rightarrow Mn^{2+}$$

ii) Dichromate in acidic solution changes colour from orange to green.

$$Cr_2O_7^{2-} \rightarrow Cr^{3+}$$

Other common oxidizing agents are chlorine, potassium chromate, sodium chlorate and manganese (IV) oxide.

Similarly, certain reducing agents undergo a visible colour change with a substance which is easily reduced.

For example,

i) A moist starch solution changes potassium iodide paper to blue-black to show that iodine is formed, $2I^- \rightarrow I_2$. That is potassium iodide is a reducing agent.

ii) Hydrogen sulphide bubbled through a solution of an oxidizing agent forms a yellow precipitate, $S^{2-} \rightarrow S$. That is H_2S is a reducing agent.

Other common reducing agents are carbon, carbon monoxide, sodium thiosulphate, sodium sulphite and iron (II) salts.

The oxidizing or reducing ability of substances depend on many factors. Some of these are:

- **Electronegativity:** Elements with high electronegativity such as F_2, O_2, N_2 and Cl_2 are good oxidizing agents. Elements with low electronegativity for example, metallic elements like Na, K, Mg and Al are good reducing agents.
- **Oxidation states:** In a compound or ion, if one of its elements is in a higher oxidation state, then it is an oxidizing agent. Similarly, if an element of a compound or ion is in a lower oxidation state, then it is a reducing agent..

Examples

$\overset{+7}{K}\overset{+7}{Mn}O_4, \overset{+7}{Na}\overset{+7}{Cl}O_4, \overset{+6}{K_2}\overset{+6}{Cr_2}O_7$ ... are oxidizing agents

$\overset{+2}{Fe}S, \overset{+2}{C}O, \overset{+4}{Na_2}\overset{+4}{S}O_3$ ... are reducing agents

4.7.4 Analysing Redox Reactions

Oxidation and reduction reactions are called Redox reactions. Oxidation and reduction reaction take place simultaneously in a given reaction.

Activity 4.11



Form a group and discuss the following:

1. Why a reducing agent undergoes oxidation?
2. Why must every redox reaction involve both an oxidizing agent and a reducing agent?

Present your discussion to the class.

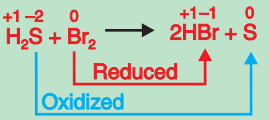
Example 4

Identify the oxidizing agent, reducing agent, the substance oxidized and reduced in the following reaction.

$$H_2S + Br_2 \rightarrow 2HBr + S$$

Solution:

Let us assign oxidation number to all the elements of the reactants and products.



In the reaction, the S atom in H_2S increases its oxidation number from –2 to 0. Hence, S is oxidized and H_2S is a reducing agent. The oxidation number of Br is decreased from 0 to –1. Thus, Br_2 is reduced and is an oxidizing agent.

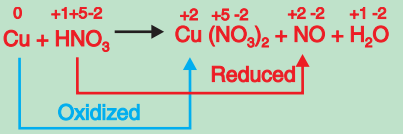
Example 5

Write the oxidizing and reducing agents for the reaction given below:

$$Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO + H_2O$$

Solution:

Write the oxidation numbers of each element and identify the substances which undergo a change in oxidization number.



Therefore, copper is a reducing agent, and HNO_3 is an oxidizing agent.

Non-redox Reactions

So far we have discussed oxidation and reduction reaction or redox reactions. However, there are also reactions in which oxidation and reduction do not occur and such types of reactions are known as non-redox reactions.

In non-redox reactions, no electrons are exchanged between the reacting substances. Therefore, the oxidation numbers of the atoms do not change in the reaction. Usually such types of reactions involve the exchange of positive and negative ions. Most of the double displacement reactions and acid-base reactions are not oxidation-reduction reactions.

Examples

$\overset{+2\ -1}{BaCl_2} + \overset{+1\ +6-2}{Na_2SO_4} \rightarrow \overset{+2+6-2}{BaSO_4}\downarrow + \overset{+1\ -1}{2NaCl}$

$\overset{+1\ -1}{HCl} + \overset{+1\ -2+1}{NaOH} \rightarrow \overset{+1\ -1}{NaCl} + \overset{+1\ -2}{H_2O}$

Exercise 4.14

1. In each of the following equations, identify the substance oxidized, the substance reduced, the oxidizing agent and reducing agent.

   a $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

   b $CuO + H_2 \rightarrow Cu + H_2O$

   c $3NO_2 + H_2O \rightarrow 2HNO_3 + NO$

   d $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$

   e $Fe^{2+} \rightarrow Fe^{3+} + e^-$

   f $3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$

   g $SO_4^{2-} + I^- + H^+ \rightarrow H_2S + I_2 + H_2O$

   h $MnO_4^- + H^+ \rightarrow Mn^{2+} + H_2O$

2. In each of the following, is the change indicated oxidation, reduction or no change at all?

   a $Cl_2 \rightarrow Cl^-$ b $MnO_4^- \rightarrow MnO_2$

   c $CrO_2 \rightarrow CrO_4^{2-}$ d $SO_4^{2-} \rightarrow SO_3^{2-}$

   e $Cr_2O_7^{2-} \rightarrow CrO_4^{2-}$

4.8 RATES OF CHEMICAL REACTIONS AND CHEMICAL EQUILIBRIUM

Competencies

By the end of this section, you will be able to:

- define rate of reaction;
- describe rate of reaction using a graph;
- carry out an experiment to illustrate the relative rate of reactions;
- list the pre-conditions for a chemical reaction to occur;
- explain how collision, activation energy and proper orientation of reactants cause a chemical reaction to occur;
- list factors that affect rate of chemical reaction;
- explain the effects of changes in temperature, concentration or pressure and surface area on the rates of a chemical reaction;
- explain the effect of catalysts on the rates of chemical reaction;
- carry out an activity on how the factors affect the rate of chemical reaction;
- define the terms reversible reaction and irreversible reaction;
- define chemical equilibrium;
- describe the characteristics of chemical equilibrium;
- write the expression for equilibrium constant of a reversible reaction;
- state Le Châtelier's principle; and
- use Le Châtelier's principle to explain the effect of changes in temperature, pressure and concentration of reactants at equilibrium.

Activity 4.12



Discuss each of the following phenomena in groups and present your findings to the class:

1. Why do some reactions take place rapidly and others slowly? Give examples of fast and slow reactions.
2. Does sugar dissolve faster in hot or in cold tea? Why?

4.8.1 Reaction Rate

Experiment 4.9



Measuring Rate of a Reaction

Objective: To measure the rate of the reaction between HCl and $CaCO_3$.

Apparatus: Balance, conical flask, cotton wool, stop watch.

Chemicals: HCl, $CaCO_3$

Procedure:

1. Take 50 mL of dilute hydrochloric acid into a conical flask.
2. Place the conical flask on a laboratory balance.
3. Add 25 g of calcium carbonate (*marble chips*) into the flask. Plug the cotton wool immediately as shown in Figure 4.11. This will help to prevent escape of the acid spray.
4. Read and record the mass of the flask and its content at one minute intervals until the reaction is over. (*Use a stop watch*).

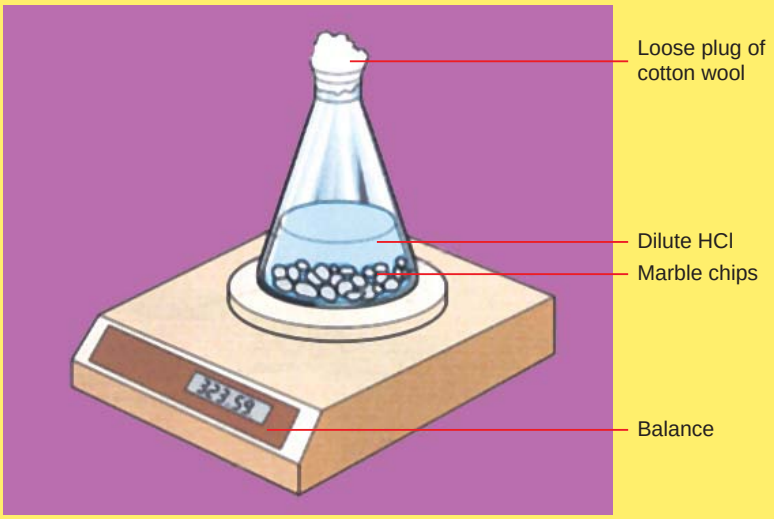


Figure 4.11 Measuring the rate of the reaction between HCl and $CaCO_3$.

Observations and analysis:

1. What happens to the mass during the reaction?
2. Write the balanced chemical equation.

Use the following table to record your results.

Time (min)	0	1	2	3	4	5	6	7
Mass (g)								
Loss in mass (g)								

3. Plot a graph between time (*x*-axis) and loss in mass (*y*-axis), and draw a smooth curve through maximum points.
4. Why the graph is steep in the beginning but horizontal at the end of the reaction?
5. At what time does the reaction stop?

Every chemical reaction proceeds at different rates or speed. Some reactions proceed very slowly and may take a number of days to complete; while others are very rapid, requiring only a few seconds.

The rate of a chemical reaction measures the decrease in concentration of a reactant or the increase in concentration of a product per unit time. This means that the rate of a reaction determines how fast the concentration of a reactant or product changes with time. The rate of a reaction is obtained by determining the concentration of reactants or products during the reaction. Methods for determining the concentration of reactants or products depend on the type of reactions. Some of the methods are:

a Colour (*changes in colour*)
b Pressure (*increase or decrease in pressure, particularly in gases*)
c Volume (*increase or decrease in size, particularly in gases*)
d Mass (*gain or loss in weight*)
e Amount of precipitate formed

Generally, the rate of a reaction can be obtained by measuring either one of the above changes in properties of substances and consequently relating to changes in their concentrations during the course of the reaction.

$$\text{Rate of reaction} = \frac{\text{Change in concentration of substance}}{\text{Change in time}} = \frac{\Delta C}{\Delta t}$$

From this expression, it follows that the rate of a reaction is inversely proportional to the time taken by the reaction.

$$\text{Rate} \propto \frac{1}{\text{Time}}$$

Figure 4.12 illustrates the changes of the rate of a chemical reaction with time. A reaction becomes slower as reactants are consumed. The reaction rate curve becomes less steep until it becomes a horizontal straight line. No more reactant is used up at this point.

Note that the rate of a reaction is the slope of the tangent to the curve at any particular time.

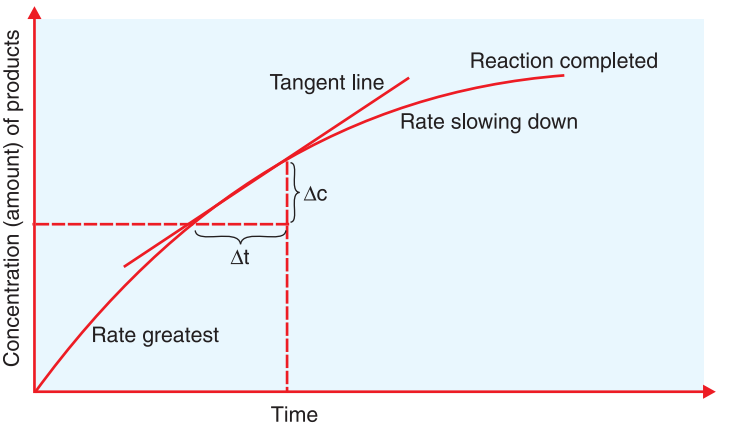


Figure 4.12 The change in concentration of product with time.

Reading Check

When a clean piece of magnesium ribbon is added to excess dilute hydrochloric acid, hydrogen gas is evolved. When a graph of volume versus time is drawn, show that the total volume of the gas evolved can be measured at fixed intervals.

Pre-conditions for a Chemical Reaction

Chemical reactions are usually explained by the collision theory. The assumption of the collision theory is that chemical reactions take place due to the collision between molecules.

1. *Collisions between reactants*

The first precondition for a reaction to occur is the direct contact of the reacting substance with each other. However, all collisions between molecules are not necessarily effective in bringing a reaction.

2. *Activation energy*

If the collisions between the reactant molecules do not have sufficient energy, then no reaction will occur. Therefore, for the reaction to take place collision must always occur with sufficient energy to break the bonds in the reactants and form new bonds in the product. Thus, minimum amount of energy needed for the reaction is known as activation energy.

3. *Proper Orientation*

Collision of molecules with sufficient activation energy will not bring a reaction if the reacting molecules are poorly oriented. Thus, the collision between molecules should have the proper orientation.

Consider the reaction between H_2 and Cl_2 molecules:

$$H_2 + Cl_2 \rightarrow 2HCl$$

For the reaction to occur, first H_2 and Cl_2 molecules must collide with each other. For the collision to be effective, these colliding molecules (H_2 and Cl_2) must have sufficient energy to break the H – H and Cl – Cl bonds and consequently to form new H – Cl bonds.

Unless, the H_2 and Cl_2 molecules are oriented in proper positions, there is no product formed. Therefore, as shown in Figure 4.13(c) the H_2 and Cl_2 molecules rearrange themselves so as to form a new H – Cl molecule.

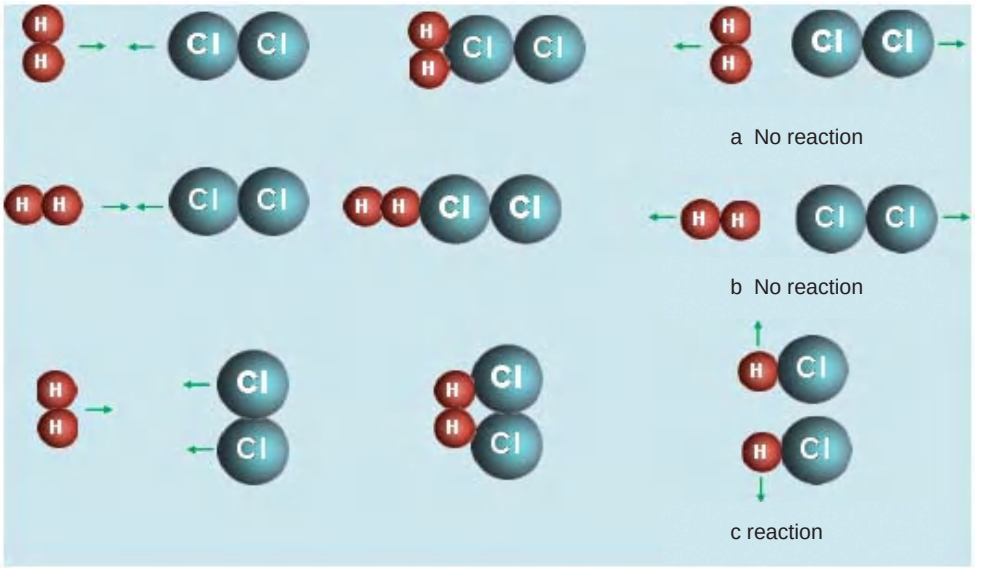


Figure 4.13 Molecular collisions and chemical reactions.

Factors Affecting the Rates of Chemical Reaction

Activity 4.13



Form a group and discuss each of the following:

1. How is the burning of charcoal affected by:
   a increasing the amount of air used
   b adding more charcoal.
2. How can you increase the rate of combustion of a given block of wood ?

Present your findings to the class.

Activity 4.14



Form a group and compare the rate of combustion of the following substances:

a paper
b kerosene
c wood charcoal
d copper

Present your findings to the class.

The rate of a chemical reaction depends on nature of reactants, temperature, and concentration of reactants, surface area and catalysts.

The collision theory assumes that the rate of a reaction depends on the number (*frequency*) of collisions of particles and those factors which affect the frequency of collisions will also affect the rate of a reaction. Now let us discuss each factor.

1. *Nature of the reactants*

The rate of a reaction is influenced by the type and nature of the reacting substances. For example, the following reactions have different rates due to the nature of the reactants, Mg, Fe and Cu.

$Mg + 2HCl \rightarrow MgCl_2 + H_2$ ... (*very fast reaction*)

$Fe + 2HCl \rightarrow FeCl_2 + H_2$ ... (*slow reaction*)

$Cu + HCl \rightarrow$ *No reaction*

2. *Temperature*

An increase in temperature increases the rate of a reaction. This is because as the temperature increases, the average kinetic energy of the particles increases which in turn increases the number of effective collisions.

In general, for many chemical reactions, the rate of a reaction doubles for every 10°C rise in temperature.

3. *Concentration of reactants*

The number of collisions is proportional to the concentration of reactants. The higher the concentration of the reactants, the more collisions between the reacting particles and thus the higher the rate of the reaction.

For example, if you heat a piece of steel wool in air (21% oxygen by volume) it burns slowly. But in pure oxygen (100% oxygen by volume) it bursts in to a dazzling white flame. This indicates that the rate of burning increases as the concentration of oxygen is higher.

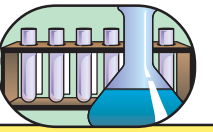
Activity 4.15



Form a group and discuss the following idea. Present your discussion to the class.

If a certain reaction is carried out with water as a solvent, what will be the rate of the reaction if more water is added to the reaction vessel? Explain.

Experiment 4.10



Investigating the effect of concentration on reaction rate

Objective: To determine the rate of reaction of magnesium with 0.1M and 5M of sulphuric acid.

Apparatus: Beakers.

Chemicals: H_2SO_4.

Procedure:

a
1. Take 20 mL of 0.1M H_2SO_4 into the first beaker.
2. Add 1 cm long magnesium ribbon into the beaker.
3. Note how fast the reaction occurs.

b
1. Take 20 mL of 5M H_2SO_4 into the second beaker.
2. Add 1cm long magnesium ribbon into the beaker as shown in Figure 4.14.
3. Observe how fast the reaction occurs.

Observations and Analysis:

1. In which of the reactions does the evolution of gas bubble faster? (*a*) or (*b*).
2. Write the balanced chemical equation.
3. What do you conclude from the experiment?

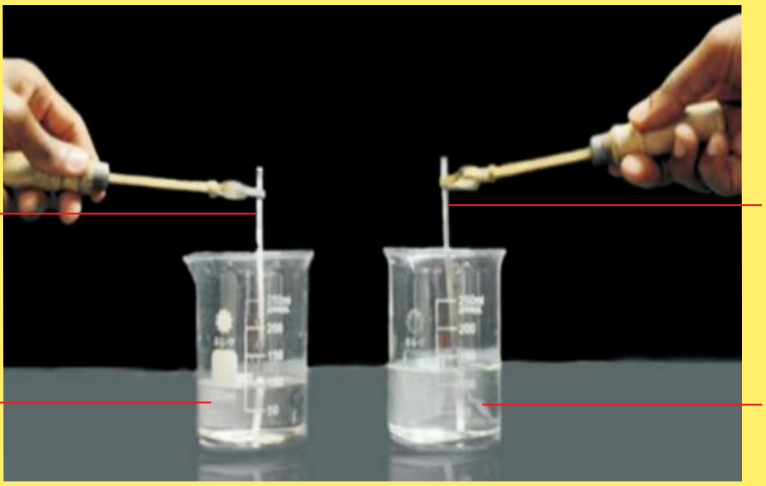
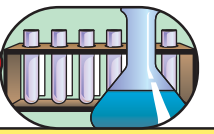


Figure 4.14 The effect of concentration on the rate of a reaction.

4. *Surface area*

When the reactants are in different phases, be it solid, liquid or gas, then the surface area of the substances affect the rate of the reaction. The higher the surface area of reactants, the faster is the rate of the reaction. This is because more contact results in more collisions between each small particle of reactants.

Experiment 4.11



Investigating the effect of surface area on reaction rate

Objective: To determine the rate of reaction of a lump and powdered calcium carbonate with hydrochloric acid.

Apparatus: Beaker, dish, and grinder.

Chemicals: Calcium carbonate and dilute hydrochloric acid.

Procedure:

a
1. Take a 5 g of calcium carbonate and put it in a beaker.
2. Add 100 mL dilute hydrochloric acid into the beaker carefully.
3. Observe how fast the reaction occurs.

b
1. Add 5 g of calcium carbonate into a dish and grind until it becomes powder.
2. Put this powdered calcium carbonate into the beaker as shown in Fig. 4.15.
3. Add 100 mL dilute hydrochloric acid carefully.
4. Observe how fast the reaction occurs.

Observations and Analysis:

1. Which of the reaction is faster? (*a*) or (*b*).
2. What do you conclude from the experiment?

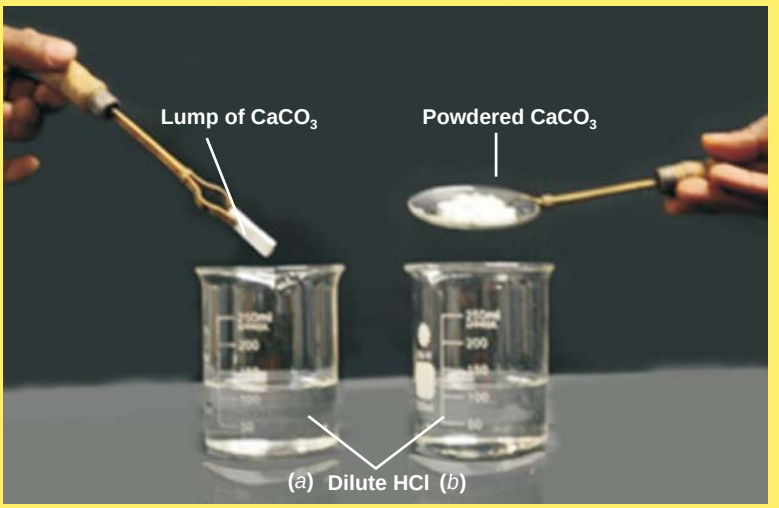


Figure 4.15 Effect of surface area on the rate of a reaction.

5. *Catalysts*

A catalyst is a substance that changes the rate of a chemical reaction without itself being consumed in the reaction. For example, the decomposition of potassium chlorate, $KClO_3$ into KCl and O_2 is made faster in the presence of MnO_2 catalyst.

$$2KClO_3 \xrightarrow[\text{Heat}]{MnO_2} 2KCl + 3O_2$$

A catalyst speeds up the rate of a reaction by providing an alternative reaction path with lower activation energy. Lower activation energy for a reaction corresponds to the higher reaction rate. Figure 4.16 illustrates how the presence of catalysts speeds up a reaction.

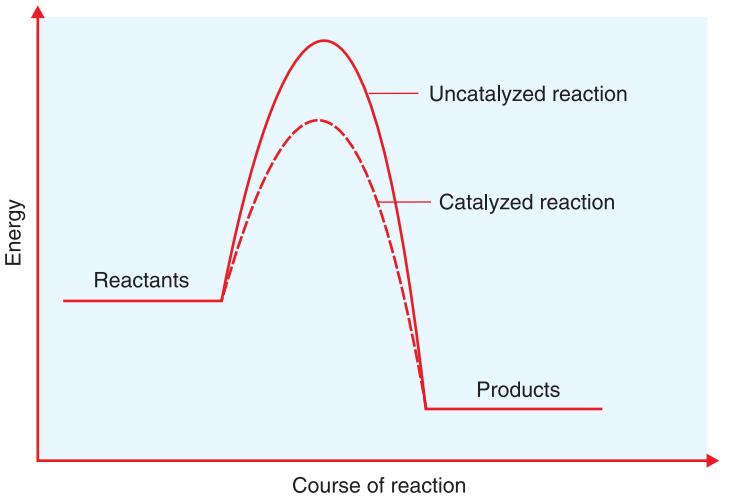


Figure 4.16 Activation energy and catalysts.

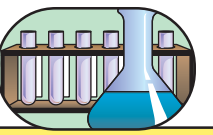
Activity 4.16



Form a group and discuss the following concept. Present your discussion to the class.

Do the factors that affect the rate of a chemical reaction influence a physical change in the same manner? Explain, by giving appropriate example.

Experiment 4.12



Investigating the effect of a catalyst on the Rate of a Chemical Reaction

Objective: To determine the effect of MnO_2 catalyst on the rate of decomposition of H_2O_2.

Apparatus: Test tube, test-tube rack.

Chemicals: MnO_2, H_2O_2.

***Procedure*:**

1. Take 5 mL of hydrogen peroxide in each of the two test tubes.
2. Add a small amount of coarse manganese (IV) oxide to the first test tube (*a*). What do you observe?
3. Leave the second test tube (*b*) without adding any chemical.
4. What change do you observe in each test tube?
5. Introduce a glowing splint in the test tube and test for the evolution of a gas.

Observations and Analysis:

1. In which of the test tubes does the reaction occur at a fast rate? Why?
2. Which gas is evolved in the test tube?
3. What do you conclude from the experiment?

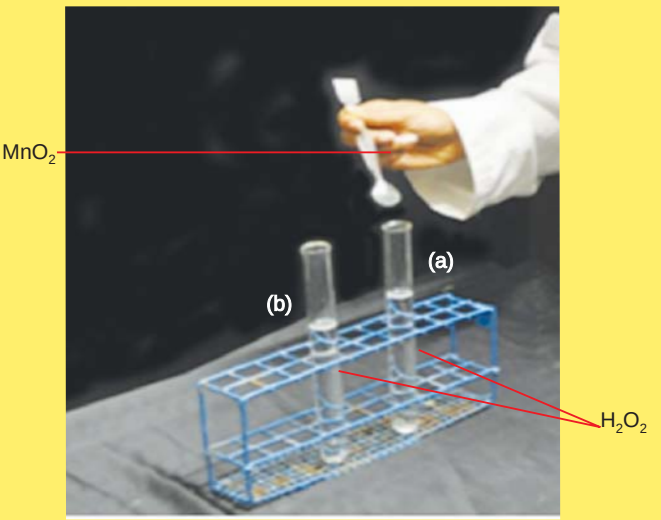


Figure 4.17 Effect of a catalyst on the decomposition of H_2O_2.

Reading Check

What is the difference between positive and negative catalysts? Explain the importance of negative catalysts by giving examples.

4.8.2 Chemical Equilibrium

In the previous chapters, we came across chemical reactions in which all the reactants are completely converted to products. Such types of reactions are known as irreversible reactions. Irreversible reactions proceed only in one direction (forward direction) and expressed by a single arrow (→).

Examples

$2Na + Cl_2 \rightarrow 2NaCl$

$2KClO_3 \rightarrow 2KCl + 3O_2$

However, there are many chemical reactions that do not proceed to completion. The products at the same time react to give (produce) the reactants. These are called reversible reactions.

Reversible reactions take place in both the forward and backward directions under the same conditions. A double arrow ($\rightleftarrows$) or ($\rightleftharpoons$) pointing in opposite directions is used in such reaction equations.

Example

$N_2 + 3H_2 \rightleftharpoons 2NH_3$

The forward reaction proceeds from left to right and the reaction that goes from right to left is the reverse reaction.

Does a reaction stop if it attains equilibrium?

Chemical equilibrium is the state of a chemical system in which the rates of the forward and reverse reactions are equal. At the state of chemical equilibrium, there is no net change in the concentrations of reactants and products because the system is in dynamic equilibrium. Dynamic equilibrium means the reaction does not stop and both the forward and the backward reactions continue at equal rates.

At equilibrium, Rate of forward reaction = Rate of reverse reaction

The law of chemical equilibrium can be expressed mathematically using the molar concentrations of reactants and products at equilibrium. The concentration of species is denoted by enclosing the formula in square bracket [ ].

Thus, for the reversible reaction:

$aA + bB \rightleftharpoons cC + dD$

Rate of forward reaction = $K_f [A]^a [B]^b$

Rate of reverse reaction = $K_r [C]^c [D]^d$

where K_f and K_r are rate constants for the forward and reverse reactions respectively.

Since at equilibrium the rate of the forward reaction equals the rate of the reverse reaction, it follows:

$$K_f\,[A]^a\,[B]^b = K_r\,[C]^c\,[D]^d$$

$$\frac{K_f}{K_r} = \frac{[C]^c\,[D]^d}{[A]^a\,[B]^b}$$

Solving for the constants, K_f/K_r, gives a new constant, termed as the equilibrium constant, K_{eq}.

Therefore, $$K_{eq} = \frac{K_f}{K_r} = \frac{[C]^c\,[D]^d}{[A]^a\,[B]^b}$$

Example

For the reaction,

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

Rate of forward reaction $= K_f\,[N_2][H_2]^3$

Rate of reverse reaction $= K_r\,[NH_3]^2$.

$$K_{eq} = \frac{K_f}{K_r} = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

The rates of the forward and reverse reactions are also illustrated by the following graph.

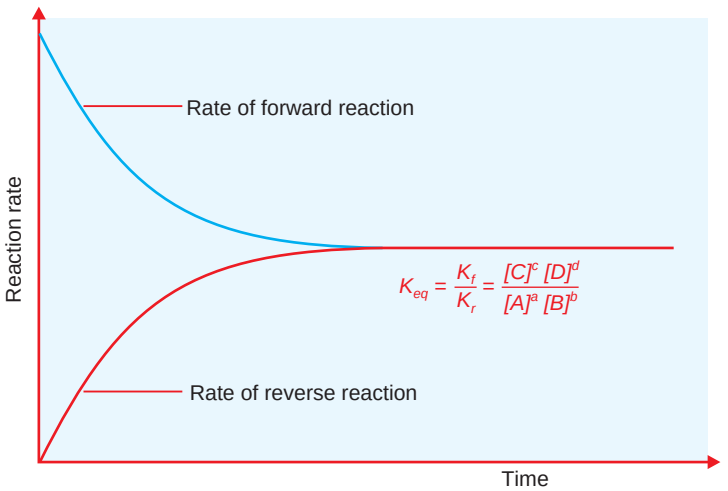


Figure 4.18 Change of the rate of the forward and reverse reactions with time.

As it is noted in the figure the rate of the forward reaction decreases with time as the concentrations of the reactants, A and B decrease with time. The reverse reaction rate starts at zero and increases as more of the products, C and D are produced. However, at equilibrium the forward and the reverse reaction rates are equal.

Example

The following equilibrium has been studied at 230°C.

$$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$$

In one experiment, the concentrations of the reacting species at equilibrium are found to be [NO] = 0.0542 M, $[O_2]$ = 0.127 M and $[NO_2]$ = 15.5 M. Calculate the equilibrium constant of the reaction at this temperature.

Solution :

The equilibrium constant is given by

$$K = \frac{[NO_2]^2}{[NO]^2[O_2]}$$

Substituting the concentration, we find that

$$K = \frac{(15.5)^2}{(0.0542)^2(0.127)} = 6.44 \times 10^5\ M^{-1}$$

Exercise 4.15

Give appropriate answers to the following questions.

1. What is the reason for all collisions between reactant molecules not to lead to products? Explain on the basis of the collision theory.
2. Explain why the rates of a reaction change with time.
3. Write the equilibrium constant expression for each of the following reactions.
   a $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$
   b $CH_4(g) + Cl_2(g) \rightleftharpoons CH_3Cl(g) + HCl(g)$
   c $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$
   d $FeO(s) + CO(g) \rightleftharpoons Fe(s) + CO_2(g)$
4. At the start of a reaction there are 0.0249 mol of N_2, 3.21×10^{-2} mol H_2 and 6.42×10^{-4} mol of NH_3 in a 3.50 L reaction vessel at 375°C.

   If the equilibrium constant, *K*, for the reaction:

   $$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

   is 1.2 at this temperature, decide whether the system is at equilibrium or not. If it is not, predict in which direction, the net reaction will proceed.

Activity 4.17



Form a group and discuss the importance of equilibrium in the study of chemical reactions.
Present your discussion to the class.

Factors that Affect Chemical Equilibrium

Activity 4.18



Form a group and try to explore at least two properties that can be utilize to determine the state of chemical equilibrium in system.
Present your findings to the class.

A system remains at a state of chemical equilibrium if there is no change in the external conditions that disturb the equilibrium. But the point of equilibrium could be affected due to any external factors like temperature, pressure, concentration and so on. How a system at equilibrium adjusts itself to any of these changes is stated by the French chemist Henri Le Chatelier in 1888.

Le Chatelier states that if *a stress is applied to a system in equilibrium, the system will respond in such a way to counteract the stress*. The stress could be change in temperature, concentration or pressure.

The factors affecting chemical equilibrium and their effects:

1. *Effect of temperature*

The effect of temperature changes on equilibrium depends on whether the reaction is exothermic or endothermic. An increase in the temperature of a system will favour an endothermic reaction and a decrease in temperature favors an exothermic reaction. For example, consider the following reaction:

$$H_2O(g) + CO(g) \rightleftharpoons H_2(g) + CO_2(g);\ \Delta H = -41\ kJ$$

Since the reaction is exothermic,

i) if temperature is increased, the system will shift to the left.

ii) if temperature is decreased, the system will shift to the right and a high yield of products (H_2 and CO_2) is obtained.

2. *Effect of Pressure or Volume*

Pressure changes only affect equilibrium reaction involving gaseous reactants and products. The effect of pressure on liquids and solids is negligible. An increase in pressure (or decrease in volume) on a gaseous system shifts the equilibrium in the direction of forming smaller number of moles of gas. On the contrary, decreasing the pressure shifts the equilibrium in the direction of forming more number of moles of the gas. For example, in the reaction,

$$\underbrace{2CO(g) + O_2(g)}_{\text{3 moles of gases}} \rightleftharpoons \underbrace{2CO_2(g)}_{\text{2 moles of gases}}$$

There are 3 moles of reactants and 2 moles of product.

Therefore, increasing the pressure of the system shifts the equilibrium to the forward direction. This results in a higher yield of CO_2. Decreasing the pressure of the gaseous mixture shifts the equilibrium in the reverse direction.

When the number of moles of reactants and products are equal, pressure has no effect on the equilibrium. For example, in the following reaction;

$$\underbrace{H_2(g) + I_2(g)}_{\text{2 moles of gases}} \rightleftharpoons \underbrace{2HI(g)}_{\text{2 moles of gases}}$$

the number of moles of reactants and products are equal (2 mol each) and hence no effect of pressure on the equilibrium.

Exercise 4.16

For the following equilibrium system, how would the position of the equilibrium be changed if:

a the temperature is increased; b the pressure is decreased?

c the temperature is increased, and d the pressure is decreased

1. $N_2(g) + O_2(g) \rightleftharpoons 2NO(g) + \text{heat}$
2. $2H_2O(g) \rightleftharpoons 2H_2(g) + O_2(g),\ \Delta H = +241.7\ kJ$
3. $N_2(g) + 2H_2O(g) + \text{heat} \rightleftharpoons 2NO(g) + 2H_2(g)$
4. $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) + \text{heat}$
5. $N_2O_4(g) + \text{heat} \rightleftharpoons 2NO_2(g)$
6. $PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g) + 21\ kJ$

3. *Effect of Concentration*

According to Le Chatelier's principle, any change in the concentration of a reactant or product will lead to a change in the concentration of the substances on the other side of the equation.

If one of the reactants is added to the equilibrium mixture, the system shifts to the forward direction and a high yield of product is obtained. To the contrary, if more product is added to the system, the equilibrium shifts to the reverse direction. For example, in the reaction,

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

If the concentration of N_2 or H_2 is increased (i.e., if more N_2 or H_2 is added), the equilibrium will shift to the right direction and more NH_3 will be produced.

4. *Effect of catalysts*

Catalysts change the speed of both the forward and reverse reactions equally. However, catalysts do not affect the state of chemical equilibrium of a reaction. This means that the position of equilibrium will not be shifted due to the presence of a catalyst.

How can Le-Chatelier's principle help in maximizing the yields of products in industrial production?

Many industrial reactions are reversible reactions. The Haber and contact processes provide excellent illustrations of the effects of temperature, pressure and catalyst on the equilibrium systems.

A Haber process (the *industrial production of ammonia*). In the Haber process, ammonia (NH_3) is industrially manufactured using gaseous nitrogen and hydrogen.

$$N_2(g) + 3H_2 \underset{\text{Catalyst}}{\overset{\text{Fe}}{\rightleftharpoons}} 2NH_3(g);\ \Delta H = -92\ \text{kJ/mol}$$

High yield of ammonia is produced by decreasing the temperature and increasing the pressure of the system.

However, the yield per unit time is highly increased by using an iron catalyst.

B Contact process (*the industrial production of sulphuric acid*). Similarly, in the contact process, sulphuric acid (H_2SO_4) is commercially manufactured by the reaction of SO_2 and O_2. The resulting SO_3 is then converted to H_2SO_4.

$$SO_2(g) + O_2 \overset{V_2O_5}{\rightleftharpoons} 2SO_3(g) + \text{heat}$$

To get high yield of SO_3, how would you adjust the temperature, pressure and concentration in the above reaction?

Check list

Key terms of the unit

- *Activation energy*
- *Catalysts*
- *Chemical equilibrium*
- *Chemical reaction*
- *Collision theory*
- *Combination reaction*
- *Decomposition reaction*
- *Double displacement reaction*
- *Dynamic equilibrium*
- *Endothermic reaction*
- *Energy changes*
- *Energy diagrams*
- *Enthalpy*
- *Equilibrium constant*
- *Exothermic reaction*
- *Heat of reaction*
- *Irreversible reaction*
- *Le Chatelier's principle*
- *Limiting reactants*
- *Oxidation*
- *Oxidizing agents*
- *Percentage yield*
- *Products*
- *Reactants*
- *Reaction rates*
- *Redox reaction*
- *Reducing agents*
- *Reduction*
- *Reversible reaction*
- *Single displacement reaction*
- *Stoichiometry*
- *Theoretical and actual yield*

Unit Summary

- *Chemical reactions are represented by chemical equations.*
- *The three basic laws of chemical reactions are: the law of conservation of mass, the law of definite proportion and the law of multiple proportions.*
- *A balanced chemical equation is an equation in which all the number of atoms of reactants and products are equal.*
- *Most of the chemical reactions are accompanied by energy changes.*
- *Exothermic reactions release heat energy to the surrounding.*
- *Endothermic reactions absorb heat energy from the surrounding.*
- *Heat of reaction or change in enthalpy is the amount of heat energy liberated or absorbed by a chemical reaction.*
- *Chemical reactions are classified into combination, decomposition, single displacement and double displacement reactions.*

- *Stoichiometry is the quantitative relationship between reactants and products.*
- *Mass-mass problems, mass-volume and volume-volume problems are the main types of stoichiometric calculations.*
- *Oxidation is the process where a substance loses electron (s) or increases its oxidation number.*
- *Reduction is the process where a substance gains electron (s) or decreases its oxidation number.*
- *Oxidizing agents are the substances reduced, and reducing agents are the substances oxidized.*
- *Rate of reaction is the change in concentration of reactants or products with time.*
- *Reversible reactions take place in both the forward and backward directions.*
- *Irreversible reactions proceed only in the forward direction.*
- *Chemical equilibrium is the state of a chemical reaction in which the rate of the forward and reverse reactions is equal.*
- *Le Chatlier's principle states that when an equilibrium is subjected to a stress, the system will shift in a direction so as to relieve the stress.*
- *The chemical equilibrium is affected by temperature, pressure, and concentrations.*

REVIEW EXERCISE ON UNIT 4

Part I: True-False Type Questions

1. In any chemical reaction, each type of atoms is conserved.
2. In a balanced chemical equation, both sides of the equation have the same number of moles.
3. The oxidizing agent is oxidized by the reducing agent.
4. Most metallic elements are strong reducing agents, whereas most non-metallic elements are strong oxidizing agents.
5. The higher the percentage yield of a chemical reaction, the more successful the reaction.
6. The rate of a reaction is directly proportional to time.
7. A catalyst increases the reaction rates by changing the reaction mechanism.
8. The concentrations of reactants and products are equal at the state of chemical equilibrium.

9. A chemical equilibrium involves two opposite reactions, where one is endothermic and the other is exothermic.

10. The amount of product obtained at equilibrium directly indicates how fast equilibrium is attained.

Part II: Write the missing words in your exercise book

11. The oxidation number of sulphur in $CuSO_4.5H_2O$ is _____ .

12. The value of the equilibrium constant (K) is changed or affected by _____ .

13. Out of all factors, the rate of a heterogeneous reaction is highly influenced by _____ .

Part III: Problems to Solve

14. How many grams of oxygen can be prepared by the decomposition of 12 grams of mercury (II) oxide?

15. 25 g of NH_3 is mixed with 4 moles of O_2 in the given reaction:

$$4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(l)$$

a Which is the limiting reactant?

b What mass of NO is formed?

c What mass of H_2O is formed?

16. Consider the following reaction:

$$2N_2O_5(g) \rightleftharpoons 4NO_2(g) + O_2(g)$$

When 40 g of N_2O_5 decomposes, 4.5 g of O_2 is formed. What is the percent yield?

17. Consider the following equilibrium:

$$N_2(g) + 2H_2O(g) + \text{heat} \rightleftharpoons 2NO(g) + 2H_2(g)$$

How would the equilibrium of the system be affected by the following changes?

a Increasing the temperature.

b Increasing the concentration of N_2.

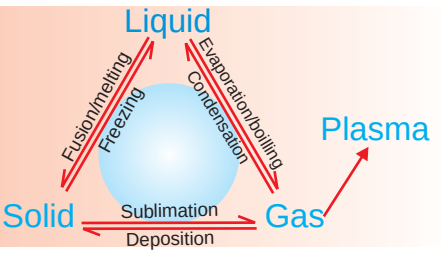
c Removing all NO and H_2.

d Compressing the reaction vessel.

e Decreasing the volume of the reaction vessel.

f Decreasing the amount of N_2 or H_2O.

UNIT 5



Physical States of Matter

Unit Outcomes

After completing this unit, you will be able to:

- *understand the kinetic molecular theory and properties of the three physical states of matter;*
- *know the behavior of gases by using the variables volume, temperature, pressure and number of moles;*
- *know terms like ideal gas, diffusion, evaporation, boiling, condensation, vapor pressure, boiling point, molar heat of vaporization, molar heat of condensation, melting, fusion, sublimation, melting point, freezing point, molar heat of fusion, molar heat of solidification;*
- *understand gas laws;*
- *develop skills in solving problems to which the gas laws apply;*
- *perform activities to illustrate gas laws;*
- *carry out experiments to determine the boiling points of liquids and the melting point of solids;*
- *demonstrate an experiment to show phase changes;and*
- *demonstrate scientific inquiry skills: observing, predicting, comparing and contrasting, measuring, interpreting data, drawing conclusion, applying concepts and making generalizations.*

MAIN CONTENTS

5.1 Introduction
5.2 The kinetic theory and properties of matter
5.3 The gaseous state
5.4 The liquid state
5.5 The solid state
– Unit Summary
– Review Exercise

Start up Activity

Objective of the activity

Scientists have conducted different activities and experiments to describe the different physical states of matter. The present simple activity will help you to form an idea about the existence of different states of matter (solid, liquid and gas) by using ice (*solid form of water*).

Procedure

1. Take a few pieces of ice in an evaporating dish. Ensure that the ice is present in solid form.
2. Heat the evaporating dish gently and observe the changes.
3. Cover the evaporating dish with a watch glass and continue heating it for more time. Record your observations.

Analysis

1. What happens to ice when it is heated?
2. When the ice is heated in the evaporating dish covered with watch glass, some droplets of water appear on it. What do you conclude from this observation? Can you name this phenomenon?
3. Name the different forms of water that you notice in the above activity.
4. Form a conclusion about the ice as it exists in solid, liquid, and gaseous state. Record your observations and present to the class.

5.1 INTRODUCTION

Competencies

By the end of this section, you will be able to:

- Name the three physical states of matter.

Activity 5.1



Discuss each of the following in your group and present your discussion to the class.

1. Name two examples for each of solids, liquids and gases.
2. What happens when you heat an ice cube?

You recall that all object around us is called matter. Matter is defined as anything that occupies space and has mass. It can exist in the form of gas, liquid and solid. The simplest example is the water we use in our daily life. The three physical states of water are:

- Steam, water in the form of gas.
- Water, in the form of liquid.
- Ice, water in the form of solid.

The changes of the states of matter are our every day experience. That is, ice melts and water freezes; water boils and steam condenses.

The physical state of a given sample of matter depends on the temperature and pressure. Changing these conditions or variables may change the behaviour of the substances as solids, liquids or gases.

Solid

A solid has a definite shape and a definite volume. Solids are almost completely incompressible and have very high average density. A high average density reflects the fact that the particles within solids are usually packed closer than those in liquids or gases do. The tightly packed particles of solids are also highly organized.

The particles of a solid, whether they are atoms, ions or molecules only vibrate about a fixed point with respect to the neighboring particles. Because of these, the particles maintain a fixed position; for example substances like metals, wood, coal and stone, are solids.

Liquid

A liquid has a definite volume, but does not have a definite shape. Liquids take the shape of their container. This is explained in terms of arrangement of particles. In the liquid state, particles vibrate about a point, and constantly shift their positions relative to their neighbours. At room temperature, water, ethanol, benzene and oil are liquids.

Gas

A gas has neither a definite shape nor a definite volume. This is because its particles are virtually independent of one another. For example, air, hydrogen, oxygen, carbon dioxide and nitrogen are gases.

Plasma

Besides, solid, liquid and gas, there exists a fourth-state of matter at very high temperature (million degrees Celsius). At such high temperatures molecules cannot exist. Most or all of the atoms are stripped of their electrons. This state of matter, a gaseous mixture of positive ions and electrons, is called plasma. Because of the extreme temperatures needed for fusion, no material can exist in the plasma state.

Exercise 5.1

1. Can oxygen exist as a liquid and solid?
2. Compare and contrast the three states of matter.
3. What is dry ice?

5.2 KINETIC THEORY AND PROPERTIES OF MATTER

Competencies

***By the end of this section, you will be able to*:**

- give examples for each of the three physical states of matter;
- state kinetic theory of matter;
- explain the properties of the three physical states of matter in terms of kinetic theory; and
- compare and contrast the three physical states of matter.

Activity 5.2



Form a group and perform the following task. Present your findings to the class.

1. Select any three different substances; one existing in the solid state, the second in the liquid state and the third in the gaseous state at room temperature. Use the following table to explain the motion, distance and attraction between particles.

	Substances	Motion of particles	Distance between particles	Attraction between particles
1				
2				
3				

2. In which state do the particles possess the highest kinetic energy?

5.2.1 The Kinetic Theory of Matter

The three states of matter in which substances are chemically the same but physically different are explained by the kinetic theory of matter. The kinetic theory of matter gives an explanation of the nature of the motion and the heat energy. According to the kinetic theory of matter, every substance consists of a very large number of very small particles called ions, atoms and molecules. The particles are in a state of continuous and random motion with all possible velocities. The motion of the particles increases with a rise in temperature.

Generally, the kinetic theory of matter is based on the following three assumptions:

1. All matter is composed of particles which are in constant motion.
2. The particles possess kinetic energy and potential energy.
3. The difference between the three states of matter is due to their energy contents and the motion of the particles.

5.2.2 Properties of Matter

Activity 5.3



Form a group and discuss each of the following idea.Present your discussion to the class.

1. Compare and contrast the density of solid, liquid and gaseous forms of a substance.
2. Discuss the compressibility of solid, liquid and gaseous forms of water.

As discussed earlier, matter exists as gas, liquid, and solid. Their properties are explained in terms of the kinetic theory as follows:

Properties of Gases

From the kinetic molecular theory of gases, the following general properties of gases can be summarized.

1. Gases have no definite shape and definite volume. This is because gases assume the volume and shape of their containers.
2. Gases can be easily compressed. By applying pressure to the walls of a flexible container, gases can be compressed; the compression results in a decrease in volume. This happens due to the large spaces between the particles of gases.
3. Gases have low densities compared with liquids and solids. This is due to the fact that the particles of a gas are very far apart and the number of molecules per unit volume is very small. A small mass of a gas in a large volume results in a very low density.
4. Gases exert pressure in all directions. Gases that are confined in a container exert pressure on the walls of their container. This pressure is due to collisions between gas molecules and the walls of the container.
5. Gases easily flow and diffuse through one another. A gas moves freely and randomly throughout in a given space.

Properties of Liquids

Liquids can be characterized by the following properties.

1. Liquids have a definite volume, but have no definite shape. They assume the shapes of their container. Lack of a definite shape for liquid substances arises from its low intermolecular forces as compared to that of solids.
2. Liquids have higher densities than gases. Their density is a result of the close arrangement of liquid particles. Thus, the particles of liquids are closer than those of gases. This accounts for the higher densities of liquids as compared to gases.
3. Liquids are slightly compressible. With very little free spaces between their particles liquids resist an applied external force and thus are compressed very slightly.
4. Liquids are fluids. A fluid is a substance that can easily flow. Most liquids naturally flow downhill because of gravity. Because liquids flow readily the molecules of a liquid can mix with each other. They flow much more slowly than gases.

Properties of Solids

1. Solids have a definite shape and a definite volume. This is due to the strong force of attraction that holds the particles of solids together.

2. Solids generally have higher densities than gases and liquids. The particles of solids are very close to each other. There is almost no empty space between the particles of solids. This closeness of particles makes solids to have more particles (mass) per unit volume, and hence solids have high density.
3. Solids are extremely difficult to compress. This is because of the high interparticle forces, and a very short distance between the particles.
4. Solids are not fluids. That is they normally do not flow. This is due to the fact that solid particles are rigidly held in position by strong forces that cause the restricted motion of their particles.

Activity 5.4



Form a group and perform the following activity:

Take a balloon and blow air into it. What happens to the volume of balloon. Explain your observations and findings to the class.

Exercise 5.2

Arrange the three states of matter in order of increasing:

a intermolecular force

b density

c compressibility

d kinetic energy

5.3 THE GASEOUS STATE

Competencies

By the end of this section, you will be able to:

- explain the assumptions of kinetic molecular theory of gases;
- describe the properties of gases using kinetic molecular theory;
- describe the behavior of gases by using the variables *V* (volume), *T* (Temperature) *P* (pressure) and *n* (number of moles);
- state Boyle's law;
- perform an activity to show changes in volume and pressure of gases to illustrates Boyle's law;

- apply Boyle's law in solving problems;
- state Charles' law;
- perform an activity to show changes in volume and temperature of gases to illustrate Charles' law;
- apply Charles' law in solving problems;
- derive combined gas law equation from Boyle's law and Charles' law;
- use the combined gas law to calculate changes in volume, pressure or temperature;
- define an ideal gas;
- derive an ideal gas equation from Boyle's law, Charles' law and Avogadro's law;
- compare the nature of real gases with ideal gases;
- solve problems related to ideal gas equation;
- define diffusion;
- state Graham's law of diffusion;
- carry out an activity to compare the rate of diffusion of two different gases;
- apply Graham's law of diffusion in solving problems.

Activity 5.5



Form a group and discuss each of the following phenomenon. Present your conclusion to the class.

Regarding the pressure in the tyres, what would you recommend to the driver of a car who is taking a trip to an area that is experiencing very cold temperatures, more or less air? Explain.

5.3.1 The Kinetic Molecular Theory of Gases

The particles in an ideal gas are very widely spaced and they are in a constant random motion. The pressure of a gas is the result of continuous collisions between the particles and the walls of their container.

Assumptions of the kinetic molecular theory of gases

1. The particles are in a state of constant, continuous, rapid, random motion and, therefore, possess kinetic energy. The motion is constantly interrupted by collisions with molecules or with the container. The pressure of a gas is the effect of these molecular impacts.

2. The volume of the particles is negligible compared to the total volume of the gas. Gases are composed of separate, tiny invisible particles called molecules. Since these molecules are so far apart, the total volume of the molecules is extremely small compared with the total volume of the gas. Therefore, under ordinary conditions, the gas consists chiefly of empty space. This assumption explains why gases are so easily compressed and why they can mix so readily.

3. The attractive forces between the particles are negligible. There are no forces of attraction or repulsion between gas particles. You can think of an ideal gas molecule as behaving like small billiard balls. When they collide, they do not stick together but immediately bounce apart.

4. The average kinetic energy of gas particles depends on the temperature of the gas. At any particular moment, the molecules in a gas have different velocities. The mathematical formula for kinetic energy is K.E. = ½ mv^2, where m is mass and v is velocity of gas molecules. Because the molecules have different velocities, they have different kinetic energies. However, it is assumed that the average kinetic energy of the molecules is directly proportional to the absolute (Kelvin) temperature of the gas.

5.3.2 The Gas Laws

Activity 5.6



Form a group and discuss the following phenomena in terms of the gas laws. Present your conclusion to the class.

a The increase in pressure in an automobile tire on a hot day.

b The loud noise heard when a light bulb shatters.

The gas laws are the products of many experiments on the physical properties of gases, which were carried out over hundreds of years. The observation of Boyle and other scientists led to the development of the Gas Laws. The gas laws express mathematical relationships between the volume, temperature, pressure, and quantity of a gas.

Pressure: pressure is defined as the force applied per unit area.

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

Pressure is one of the measurable properties of gases. Thus, the pressure of a gas can be expressed in unit of atmosphere, Pascal, torr, millimetre of mercury. The SI unit of pressure is Pascal (Pa), and is defined as one Newton per square metre.

$1\ Pa = 1\ N/m^2$ and

$1\ atm = 760\ mmHg = 76\ cmHg = 760\ torr = 101325\ Pa = 101.325\ kPa$

Volume: Volume is the space taken up by a body. The SI unit of volume is the cubic metre (m^3). Volume is also expressed in cubic centimetre (cm^3) and cubic decimetre (dm^3). Other common units of volume are millilitre (mL) and litre (L).

$$1\ cm^3 = (1\times10^{-2}\ m)^3 = 1\times10^{-6}\ m^3$$

$$1\ dm^3 = (1\times10^{-1}\ m)^3 = 1\times10^{-3}\ m^3 = 1\ L$$

A litre is equivalent to one cubic decimeter: The relation is given as follows

$$1\ L = 1000\ mL = 1000\ cm^3 = 1\ dm^3$$

Temperature: Temperature is the degree of hotness or coldness of a body. Three temperature scales are commonly used. These are °F (degree Fahrenheit), °C (degree Celsius) and K (Kelvin). In all gas calculations, we use the Kelvin scale of temperature.

We use the following formulae for all necessary inter-conversions:

$$K = °C + 273$$

$$°C = (°F - 32)\frac{5}{9}$$

$$°F = \left(\frac{9}{5}\times °C\right) + 32$$

Exercise 5.3

Convert the following:

a. 500 mmHg into atm, torr, and cmHg

b. 100 dm^3 into mL, cm^3, L, m^3

c. 54°C into K and °F.

Molar Volume and Standard Conditions (STP)

The conditions of a pressure of 1 atmosphere and a temperature of 0°C (273.14 K) are called standard temperature and pressure or STP for gases. At STP the volume of one mole of any gas is equal to 22.4 litres. This volume is known as molar gas volume.

Quantity of gas: The quantity of a gas is expressed in mole (*n*). Mole is the quantity of gas in terms of number of particles. It is the number of atoms or molecules in 1 gram-atom or 1 gram-molecule of an element or a compound.

1. *Boyle's Law*

Activity 5.7



Form a group and discuss the following phenomenon. Present your discussion to the class.

Explain why a helium weather balloon expands as it rises more in the air. (Assume the temperature remains constant.)

The first quantitative experiments on gases were performed by the Irish chemist, Robert Boyle (1627-1691). His experiment helped to analyze the relationship between the volume and pressure of a fixed amount of a gas at constant temperature. Decreasing the external pressure, causes the gas to expand and to increase in volume. Correspondingly, increasing the external pressure allows the gas to contract and decrease in volume. This is shown in Figure 5.1.

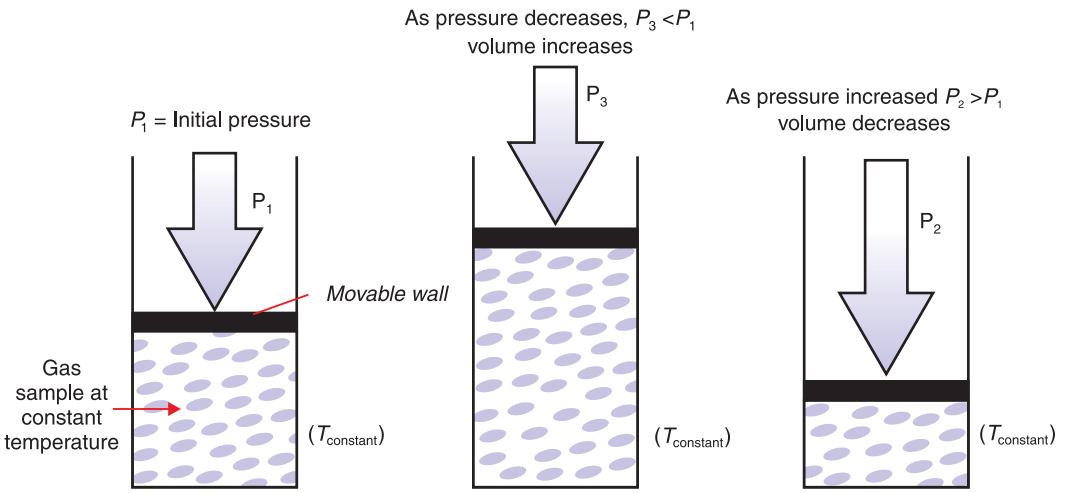


Figure 5.1 The relation between pressure and volume.

Boyle studied the relationship between the pressure of the trapped gas and its volume. Accordingly, he discovered that at constant temperature doubling the pressure on a

sample of gas reduces its volume by one-half. Tripling the gas pressure reduces its volume to one-third of the original. Generally, the volume of a gas decreases, as the pressure on the gas increases. This volume-pressure relationship is illustrated in Table 5.1.

Table 5.1 Pressure and volume data for a gas at constant mass and temperature.

Presssure (atm)	Volume (mL)	$P \times V$
0.5	1200	600
1.0	600	600
2.0	300	600
3.0	200	600
4.0	150	600
5.0	120	600
6.0	100	600

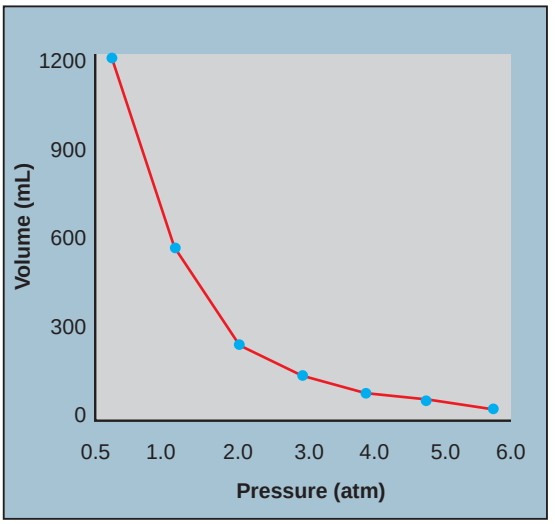


Figure 5.2 Volume versus pressure graph for a gas at constant temperature and mass.

The general volume-pressure relationship illustrated above is called Boyle's law. Boyle's law states that *the volume of a fixed mass of gas is inversely proportional to the pressure at a constant temperature*. The inverse relationship between pressure and volume is mathematically given as

$$V \ \alpha \ \frac{1}{P} \text{ (at constant } T \text{ and } n)$$

From which follows,

$$V = k\frac{1}{P} \quad \text{or} \quad PV = k;$$

where k is a constant at a specific temperature for a given sample of gas.

If P_1 and V_1 represent the initial conditions; and P_2 and V_2 represent the new or final conditions, Boyle's law can be written as:

$$P_1V_1 = P_2V_2;$$

Experiment 5.1



Effect of Pressure on the Volume of Gas

Objective: To observe the relationships between the volume and pressure of a gas at constant temperature.

Apparatus: U-tubes, ruler, rubber tube, burette, glass tube.

Chemicals: Mercury.

Procedure:

1. Join two tubes by a rubber tubing to give a U-arrangement as shown in Figure 5.3, and then partially fill these two tubes with mercury.
2. Put a ruler in the middle of the tube.
3. The first arm of the tube (A) contains air and is sealed by a tap.
4. By moving the second arm of the tube (B) up and down, the volume of air in the first tube can be varied.
5. The pressure exerted on the air is obtained from the difference in height of mercury in the two arms of the tube.

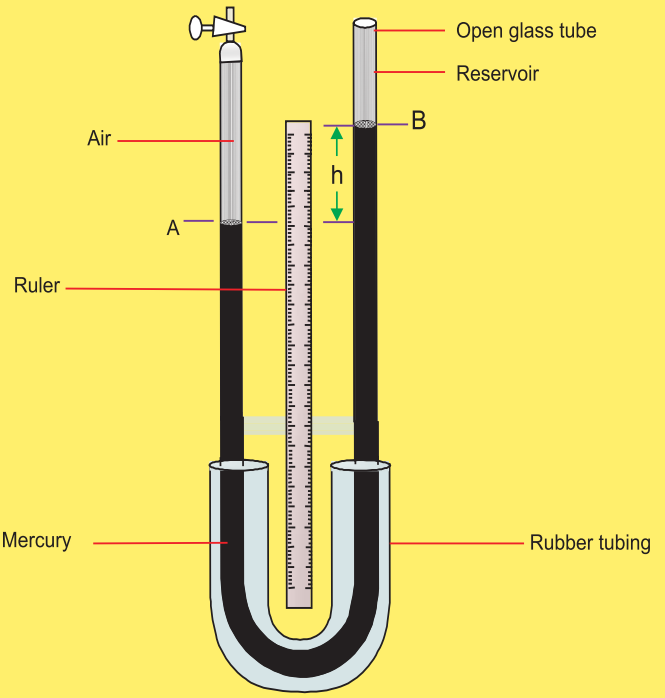


Figure 5.3 Effect of pressure on the volume of a gas at constant temperature.

Observations and analysis:

1. Plot a graph taking pressure on the vertical axis *versus* volume on the horizontal axis and comment on the shape of the graph.
2. What can you conclude from this experiment?

Activity 5.8



Discuss the following activity in your group and present your discussion to the class. Plot a graph of pressure, P, versus 1/V, using the following data:

(Note: Calculate 1/V values from the given volume by yourself)

Pressure(atm)	1	2	3	4	5
Volume(mL)	20	10	6.7	5	4

1. What does the graph look like?
2. What does the graph indicate?
3. What can you conclude from the graph?

Example 1

An inflated balloon has a volume of 0.55 L at sea level (1.0 atm) and is allowed to rise to a height of 6.5 km, where the pressure is about 0.40 atm. Assuming that the temperature remains constant, what is the final volume of the balloon?

Solution:

Givens:

Initial conditions	**Final conditions**
$P_1 = 1.0$ atm	$P_2 = 0.40$ atm
$V_1 = 0.55$ L	$V_2 = ?$

Use Boyles' law equation: $P_1V_1 = P_2V_2$

Therefore, $V_2 = \frac{P_1V_1}{P_2} = 0.55\,\text{L} \times \frac{1.0\ \cancel{\text{atm}}}{0.40\ \cancel{\text{atm}}} = 1.4$ L.

Exercise 5.4

1. A certain gas occupies a volume of 10.0 m^3 at a pressure of 100.0 kPa. If its volume is increased to 20 m^3, what would be the new pressure of the gas assuming temperature remains constant?
2. A laboratory experiment shows 4.0 litres of helium gas trapped in a cylinder at a pressure of 7.0 atm. The pressure is decreased, at a constant temperature, to 2.00 atm. What is the new volume?

2. *Charles' Law*

Historical Note



Jacques Alexandre Ceser Charles

The French scientist Charles was most famous in his lifetime for his experiment in ballooning. The first such flights were made in 1783, using a balloon of linen paper filled with hot air. Charles filled a silk balloon with hydrogen. He developed the gas law what we call Charles' law. The volume of samples of gases increases with increasing temperature (at constant pressure).

Activity 5.9



Form a group and discuss the following activity. Present your discussion to the class.

1. Have you ever wondered what makes a pop corn 'pop'?
2. Explain Charles' law in terms of kinetic molecular theory.

The French physicist, Jacques Charles (1746-1823), was the first person to fill a balloon with hydrogen gas and made the first solo balloon flight.

In 1787, Jacques Charles investigated quantitative relationship between the volume and temperature of a fixed quantity of gas which is held at constant pressure.

This can be related to real life by using an empty balloon. Accordingly, if we take a balloon filled with air and place it on boiling water, the volume of the balloon visibly increases as shown in Figure 5.4(b). If we take the balloon and place it in the freezer

compartment of our refrigerator, its volume shrinks drastically as the air inside cools (Figure 5.4(c)).

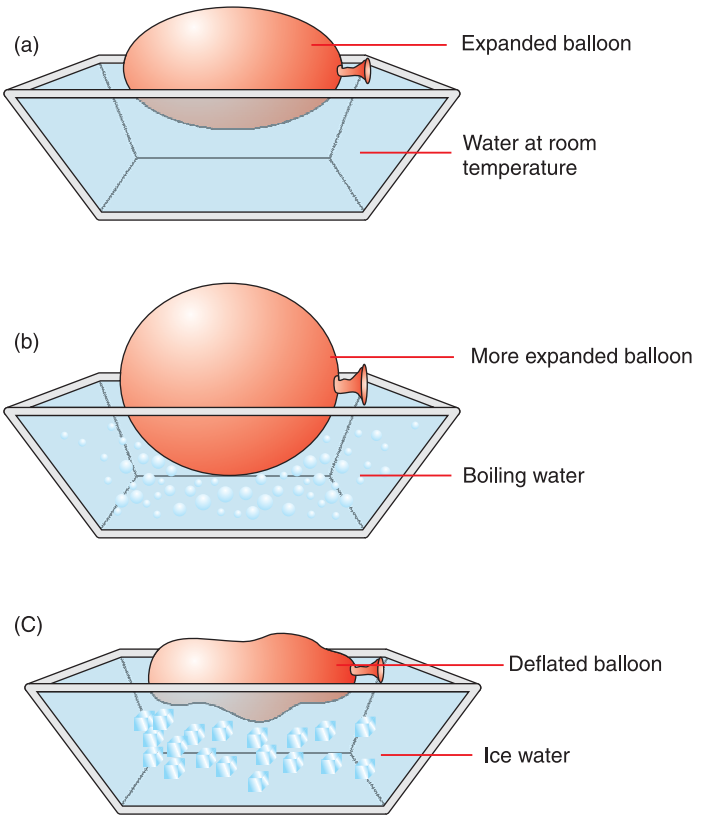


Figure 5.4 Relationship between the volume of air in the balloon and its temperature.

From your observation what can you conclude about the effect of temperature on the volume of a gas?

In 1848, Lord Kelvin realized that a temperature of -273.15ºC is considered as absolute zero. Absolute zero is theoretically the lowest attainable temperature. Then he set up an absolute temperature scale, or the Kelvin temperature scale, with absolute zero as the starting point on the Kelvin scale.

The average kinetic energy of gas molecules is closely related to the Kelvin temperature. The volume of a gas and Kelvin temperature are directly proportional to each other. For example, doubling the Kelvin temperature causes the volume of a gas to double, and reducing the Kelvin temperature by half causes the volume of a gas to decrease by half. This relationship between Kelvin temperature and the volume of a gas is known as Charles' law.

Charles' law states that *the volume of a fixed mass of gas at constant pressure varies directly with the Kelvin temperature.*

Mathematically;

$$V \propto T \text{ at constant } P \text{ and } n:$$

$$V = kT, \quad \text{or} \quad \frac{V}{T} = k$$

The value of T is the Kelvin temperature, and k is a constant. The value of k depends only on the quantity of gas and the pressure. The ratio V/T for any set of volume-temperature values always equals the same k. Charles' law can be applied directly to volume-temperature problems using the relationship:

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

where V_1 and T_1 represent the initial condition; V_2 and T_2 represent the new condition.

Charles found that the volume of a gas at constant pressure increases linearly with the temperature of the gas. That is, a plot of the volume of a gas versus its temperature (K) gives a straight line assuming the pressure remains constant. Figure 5.5 illustrates the relationship between the volume of a gas and Kelvin temperature. If the graph is extrapolated to give zero volume for a gas, the temperature reached is –273.15°C. This temperature is the lowest temperature attained by a gas and it is called absolute zero.

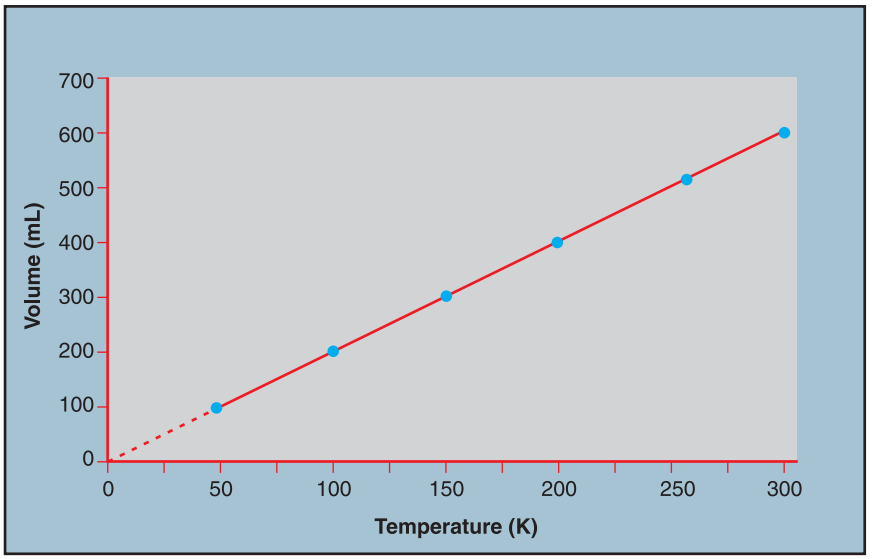


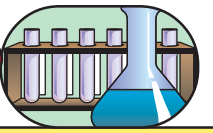
Figure 5.5 A plot of volume versus the Kelvin temperature.

Note that the ratio V/T is constant for every plot of the curve. Figure 5.5 is drawn using the data given in Table 5.2 for the same sample of gas. In all cases, V/T is constant as noted in the table.

Table 5.2 Volume-Temperature Data for a Gas sample (at constant mass and pressure).

Volume (mL)	Temperature in Kelvin	V/T
600	300	2
500	250	2
400	200	2
300	150	2
200	100	2
100	50	2

Experiment 5.2



Effect of Temperature on the Volume of Gas

Objective: To observe the changes in volume of a gas as temperature changes.

Apparatus: Round bottomed flask, beaker, delivery tube and burner.

Procedure:

1. Set up the apparatus as shown in Figure 5.6.
2. Warm the flask gently with a low Bunsen flame.
3. Cool the flask and note what happens.
4. Record your observation.

Observations and analysis:

1. What do you observe from the experiment?
2. What is your conclusion from this activity?

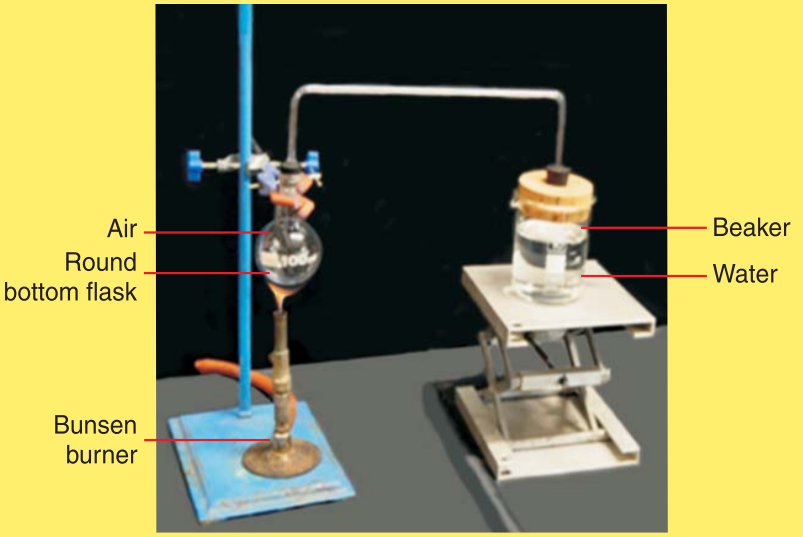


Figure 5.6 Relationship between temperature and volume of a gas.

Activity 5.10



Form a group and perform the following activity and present to the class.

Given the following data at a constant pressure:

Volume of Nitrogen gas (L)	Temperature (K)
4.28 L	303
5.79 L	410
7.77 L	550

a Draw a graph of the relationship between volume and temperature.

b Calculate the expected volume of the gas when the temperature reaches 700K.

c Explain the relationship between temperature and volume.

Example 2

A quantity of gas occupies a volume of 804 cm^3 at a temperature of 127°C. At what temperature will the volume of the gas be 603 cm^3, assuming that there is no change in the pressure?

Solution:

Given:

Initial conditions	Final conditions
T_1 = 127°C = 127 + 273 = 400 K V_1 = 804 cm^3	T_2 = ? V_2 = 603 cm^3

Based on Charles' equation $\left(\frac{V_1}{T_1}=\frac{V_2}{T_2}\right)$, the unknown variable ($T_2$) can be calculated as:

$$\frac{V_1}{T_1}=\frac{V_2}{T_2} \quad \Rightarrow \quad T_2 = T_1 \times \frac{V_2}{V_1}$$

$$= 400\,\text{K} \times \frac{603\,\text{cm}^{\cancel{3}}}{804\,\text{cm}^{\cancel{3}}} = 300\,\text{K}$$

Example 3

A gas at 65°C occupies 4.22 L. What will be the volume of the gas at a temperature of 36.9°C, assuming a constant pressure?

Solution:

Given:

Initial conditions	Final conditions
T_1 = 65°C = 65 + 273 = 338 K	T_2 = 36.9°C + 273 = 309.9 K
V_1 = 4.22 L	V_2 = ?

From **Charles'** law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \Rightarrow \quad V_2 = \frac{T_1}{V_1} \times T_2$$

$$= \frac{4.22\ \text{L} \times 309.9\ \text{K}}{338\ \text{K}} = 3.87\ \text{L}$$

Exercise 5.5

1. At constant pressure, by what fraction of its volume will a quantity of gas change if the temperature changes from -173°C to 27°C?
2. At what temperature will the volume of a gas be
   a halved, b doubled,
   c tripled at constant pressure if the original temperature is 17°C?
3. At 25°C and 1 atm a gas occupies a volume of 1.5 dm^3. What volume will it occupy at 100°C and 1 atm?

3. *The Combined Gas Law*

A sample of a gas often undergoes changes in temperature, pressure, and volume. When this happens, the three variables must be dealt with at the same time.

Boyle's law and Charles' law can be combined to give one expression called the combined gas law. The combined gas law expresses the relationship between pressure, volume, and temperature of a fixed amount of gas.

Derivation of the combined gas law:

Boyle's law: $V \propto 1/P$

Charles' law: $V \propto T$

Then, $V \propto T/P$ (*combined*)

$V = kT/P$ (where k *is a constant*)

It follows,

$$\frac{P_1V_1}{T_1} = k \quad \text{and} \quad \frac{P_2V_2}{T_2} = k$$

Since in each case k is constant, the combined gas law equation is given as follows:

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

Where P_1, V_1 and T_1 are the initial pressure, volume and temperature; P_2, V_2 and T_2 are the final pressure, volume and temperature of the gas respectively.

Example 4

A 300 cm^3 sample of a gas exerts a pressure of 60.0 kPa at 27°C. What pressure would it exert in a 200 cm^3 container at 20°C?

Solution:

Given:

Initial Conditions	V_1 = 300 cm^3	T_1 = 27 + 273 = 300 K	P_1= 60.0 kPa
Final Conditions	V_2 =200 cm^3	T_2 = 20 + 273 = 293 K	P_2 = ?

Using the combined gas law, $\left(\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}\right)$.

$$\Rightarrow \quad P_2 = \frac{P_1V_1}{T_1} \times \frac{T_2}{V_2} = \frac{60.0\ \text{kPa} \times 300\ \text{cm}^3}{300\ \text{K}} \times \frac{293\ \text{K}}{200\ \text{cm}^3}$$

$$P_2 = 87.9\ \text{kPa}$$

Exercise 5.6

If a 50 cm^3 sample of gas exerts a pressure of 60.0 kPa at 35°C, what volume will it occupy at STP?

4. *Avogadro's law*

Activity 5.11



Form a group and discuss the following phenomena. Present your discussion to the class.

Suppose while you are playing a football in your school football team, the ball is accidentally deflated. Then immediately you fill the ball with air using air pump.

i. Why did the ball become strong enough?

ii. What happened to the number of particles in the ball?

iii. Which gas law can be obeyed? Explain.

The relationship between the volume of a gas and its number of molecules was explained by Avogadro. Avogadro's law states that *equal volumes of different gases, under the same conditions of temperature and pressure, contain the same number of molecules.* Thus, according to the law the volume of a gas is proportional to the number of molecules (*moles*) of the gas at STP.

Mathematically, $V \propto n$; where V is the volume and n is number of moles.

5. *The Ideal Gas Equation*

Activity 5.12



Form a group and discuss the following phenomenon. Present your discussion to the class.

A balloon can burst when too much air is added into it. Describe what happens to the pressure, volume, temperature, number of moles, and the balloon itself as it is inflated and finally bursts. Can you derive an equation which describes these relationships?

An ideal gas is a hypothetical gas that obeys the gas laws. Real gases only obey the ideal gas laws closely at high temperature and low pressure. Under these conditions, their particles are very far apart. The ideal gas law is a combination of Boyle's law, Charles' law and Avogadro's law.

We have considered the three laws that describe the behavior of gases as revealed by experimental observations:

Boyle's law: $V \propto \frac{1}{P}$ (*at constant* T *and* n)

Charles' law: $V \propto T$ (*at constant* P *and* n)

Avogadro's law: $V \propto n$ (*at constant* P *and* T)

This relationship indicates how the volume of gas depends on pressure, temperature and number of moles.

$$V \propto \frac{nT}{P}$$

or $$V = R\frac{nT}{P}$$

where R, is a proportionality constant called the gas constant.

$$PV = nRT \quad (\textit{the ideal gas equation})$$

Thus, the ideal gas equation describes the relationship among the four variables P, V, T and n. An ideal gas is a gas whose pressure-volume-temperature behavior can be completely explained by the ideal gas equation.

At STP, the values of R can be calculated from the ideal gas equation.

$$R = \frac{PV}{nT} = \frac{(1\,\text{atm})\,(22.414\,\text{L})}{(1\,\text{mol})(273.15\,\text{K})}$$

$$= 0.082057 \frac{\text{L.atm}}{\text{K.mol}} = 8.314 \text{ L. kPa/K.mol} = 8.314 \text{ J/mol.K}$$

For calculations, we round off the value of R to three significant figures (0.0821 L.atm/K.mol) and use 22.4 L for the molar volume of a gas at STP.

Example 5

Calculate the volume (in liters) occupied by 7.4 g of CO_2 at STP?

Solution:

The ideal gas equation is given as

$$PV = nRT$$

$$V = \frac{nRT}{P} \left(\text{since } n = m/M \text{ by rearranging } V = \frac{mRT}{M\,P} \right)$$

$$= \frac{7.4\,\text{g}}{44\,\text{g/mol}} \times 0.082 \frac{\text{L.atm}}{\text{K mol}} \frac{273\,\text{K}}{1\,\text{atm}} = 3.77 \text{ L}$$

Example 6

At STP, 0.280 L of a gas weighs 0.400 g. Calculate the molar mass of the gas.

Solution:

Given: V = 0.280 L, m = 0.400 g

At conditions of standard temperature and pressure

$$T = 273 \text{ K, P} = 1 \text{ atm}$$

$$R = 0.082 \text{ L.atm/K.mol}$$

$$\Rightarrow \quad PV = nRT$$

$$PV = \frac{m}{M} \cdot RT \text{ (Since } n = \frac{m}{M}\text{)};$$

$$\Rightarrow \quad M = \frac{mRT}{PV} = \frac{0.400\,\text{g} \times 0.082\,\text{L. atm/K.mol} \times 273\,\text{K}}{1\,\text{atm} \times 0.280\,\text{L}} = 31.98 \text{g/mol}$$

Exercise 5.7

The density of a gas at a pressure of 1.34 atm and a temperature of 303 K is found to be 1.77 g/L. What is the molar mass of this gas?

6. *Graham's Law of Diffusion*

Activity 5.13



Form a group and discuss the following phenomenon. Present your discussion to the class.

Explain why a helium-filled balloon deflated over time faster than an air-filled balloon does. (Hint: A balloon has many invisible pinholes).

We have seen that a gas tends to expand and occupy any space available to it. This spreading of gas molecules is called diffusion.

How do you compare the rate of diffusion of molecules with different densities?

Thomas Graham (1805 - 1869), an English chemist, studied the rate of diffusion of different gases. He found that gases having low densities diffuse faster than gases with high densities. He was able to describe quantitatively the relationship between the density of a gas and its rate of diffusion. In 1829, he announced what is known as Graham's law of diffusion.

Graham's law of diffusion states that *at constant temperature and pressure, the rate of diffusion of a gas, r, is inversely proportional to the square root of its density, d, or molar mass, M.*

Mathematically it can be expressed as:

$$r \propto \sqrt{\frac{1}{d}} \quad \text{or} \quad r \propto \sqrt{\frac{1}{M}};$$

where r is the rate of diffusion, d is the density and M is the molecular mass of the gas.

For two gases (Gas 1 *and* Gas 2), their rates of diffusion can be given as:

$$r_1 \propto \sqrt{\frac{1}{d_1}} \quad or \quad r_1 \propto \sqrt{\frac{1}{M_1}}$$

and

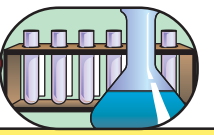
$$r_2 \propto \sqrt{\frac{1}{d_2}} \quad or \quad r_2 \propto \sqrt{\frac{1}{M_2}}$$

Rearranging these relationships gives the following expression

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} \quad or \quad \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}};$$

where r_1, d_1 and M_1 represent the rate of diffusion, density and molecular mass of gas 1. r_2, d_2 and M_2 represent the rate of diffusion, density and molecular mass of gas 2.

Experiment 5.3



Determination of Diffusion of Gases

Objective: To compare the rates of diffusion of HCl and NH_3 gases.
Apparatus: Glass tube, cork, cotton and stand.
Chemicals: conc NH_3, conc. HCl.

***Procedure*:**

1. Set up the apparatus as shown in Figure 5.7.
2. Insert pieces of cotton at the two ends of the tube.
3. Add 8 drops of concentrated ammonia on the cotton at one end and 8 drops of concentrated hydrochloric acid at the other end of the tube at the same time. Immediately close the two ends with corks.
4. Watch till a white ring is formed and record the time at which the white ring is formed.
5. Measure the distances between the white ring and the two ends.

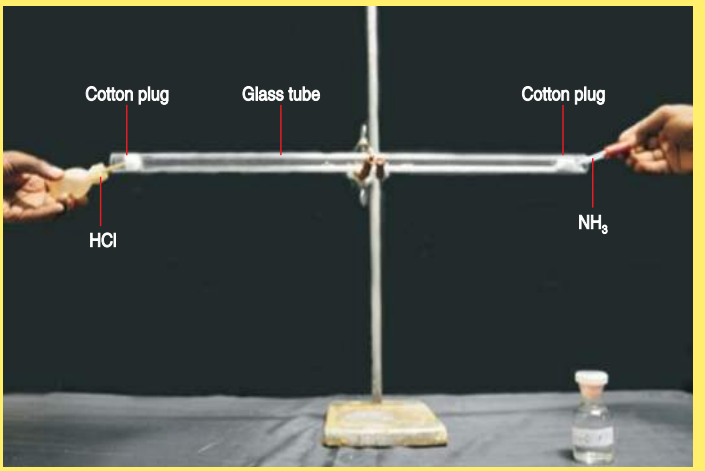


Figure 5.7 Determination of diffusion of gases.

Observations and analysis:

1. Which gas has moved the shorter distance to the white ring?
2. How do you compare the rate of diffusion of the two gases?
3. Which gas diffuses faster? HCl or NH_3?
4. Write your conclusions about the experiment.

Example 7

Which gas will diffuse faster, ammonia or carbon dioxide? What is the relative rate of diffusion?

Solution:

The molecular weight of CO_2 is 44 g/mol and that of NH_3 is 17 g/mol. Therefore, NH_3 diffuses faster than CO_2.

We can calculate the rate of diffusion as follows:

Let the rate of diffusion of NH_3 be r_{NH_3}

Let the rate of diffusion of CO_2 be r_{CO_2}

$$\frac{r_{NH_3}}{r_{CO_2}} = \sqrt{\frac{M_{CO_2}}{M_{NH_3}}} = \sqrt{\frac{44}{17}} = 1.6$$

This means rate of diffusion of NH_3 is 1.6 times that of CO_2.

Example 8

The rate of diffusion of methane (CH_4) is twice that of an unknown gas. What is the molecular mass of the gas?

Solution:

Let r_{CH_4} and r_x be the rates of diffusion of CH_4 and the unknown gas respectively.

Let M_{CH_4} and M_x be the molecular masses of CH_4 and the unknown gas respectively.

The rate of diffusion of CH_4 is two times faster than the unknown gas. This can be written mathematically as $r_{CH_4} = 2r_x$.

Now, substitute $2r_x$ in place of r_{CH_4} and solve for M_x using Graham's law.

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} \Rightarrow \frac{r_{CH_4}}{r_x} = \sqrt{\frac{M_x}{M_{CH_4}}}$$

$$\frac{2r_X}{r_X} = \sqrt{\frac{M_X}{16}} \Rightarrow M_X = 64$$

Therefore, the molecular mass of the unknown gas is 64.

Also the rate at which a gas diffuses is inversely proportional to the time taken. Mathematically,

$$r \propto \frac{1}{t}$$

If two different gases (gas 1 and gas 2) under the same conditions of temperature and pressure diffuse through a porous container, then the time required to diffuse for the two gases can be given by the following formula:

$$\frac{r_1}{r_2} = \frac{t_2}{t_1} = \sqrt{\frac{M_2}{M_1}}$$

where t_1 and t_2 are the time taken, r_1 and r_2 are the rates, M_1 and M_2 are the molecular masses of Gas 1 and Gas 2 respectively.

Activity 5.14



Form a group and discuss the kinetic molecular theory to explain the compression and expansion of gases. Present your findings to the class.

Reading Check

What names are given for the following ideal gas relationships?

a Volume and moles at constant temperature and pressure.

b Volume and pressure at constant temperature and moles.

c Volume and Kelvin temperature at constant pressure and moles.

5.4 THE LIQUID STATE

Competencies

By the end of this section, you will be able to:

- explain the terms: evaporation, boiling, condensation, vapor pressure; boiling point, molar heat of vaporization and molar heat of condensation;
- carry out an activity to demonstrate the concept of vapor pressure; and
- carry out an activity to determine the boiling points of water and ethanol.

Activity 5.15



Discuss the following activities in your group and present to the class.

1. Why some liquids are volatile and others are not?
2. What is the relationship between altitude and boiling point of a liquid?

You recall that liquids have a definite volume and an indefinite shape. They take the shape of their containers to the level they fill. On the average, liquids are more dense than gases, but less dense than solids.

As in a gas, particles in a liquid are in constant motion. However, the particles in a liquid are closer together than those in a gas. The attractive forces between particles in a liquid are more effective than between particles in a gas. This attraction between liquid particles is caused by the intermolecular forces (*dipole-dipole forces, London dispersion forces, and hydrogen bonding*).

Liquids are more ordered than gases because of the stronger intermolecular forces and the lower mobility of liquid particles. Accordingly, liquid particles are not bound together in fixed positions.

Energy Changes in Liquids

Activity 5.16



Form a group and discuss the following phenomenon:

When you take bath with hot water in your bathroom, the water collects on the mirror of the bathroom.

Present your discussion to the class.

The process by which a liquid changes to a gas is known as vaporization. Evaporation is the process by which liquid molecules break freely from the liquid surface and enter the vapor phase. Evaporation is explained in terms of the energy possessed by the molecules on the surface of the liquid.

In an open container, evaporation continues until all of the liquid enters the vapor phase. However, liquids in a closed container behave differently. The volume of the liquid decreases for a period of time, and remains unchanged. In closed containers, the vapor cannot escape. As the vapor concentration increases, some of the vapor molecules lose energy and return to the liquid state. When a vapor returns to the liquid state, it is said to condense. The process is called condensation. Evaporation and condensation are opposing processes.

$$\text{Liquid} \underset{\text{Condensation}}{\overset{\text{Evaporation}}{\rightleftharpoons}} \text{Gas}$$

The rate of evaporation of a liquid depends on three factors. These are temperature, intermolecular forces, and surface area of the liquid.

An increase in temperature increases the average kinetic energy of the molecules and this increases the tendency to change into the gaseous state. Some liquids evaporate readily at room temperature. Such liquids are said to be volatile. The volatile liquids have relatively weak forces of attraction between particles. Diethyl ether, ethyl alcohol, benzene and acetone are volatile liquids.

Non-volatile liquids have a little tendency to evaporate at a given temperature. They have relatively stronger attractive forces between their molecules, e.g., sulphuric acid, water, and molten ionic compounds.

Vapour pressure: The partial pressure of the vapour above a liquid is called vapour pressure. The vapour pressure of a liquid depends up on the temperature. At a given temperature, vapour pressure is constant. As the temperature increases, the vapour pressure of a liquid also increases due to high rate of evaporation.

Vapour pressure depends also on the strength of the intermolecular forces between the particles of the liquid. The stronger the intermolecular forces, the lower the vapour pressure will be, because fewer particles will have enough kinetic energy to overcome the attractive force at a given temperature. For example, water and alcohol have relatively low vapour pressure. On the contrary, liquids with low intermolecular forces have high vapour pressures at room temperature. For example, diethyl ether, a non-polar molecule with relatively weak dispersion forces, has a relatively higher vapour pressure.

Boiling and Boiling Point

Activity 5.17



Form a group and discuss the following ideas. Present your discussion to the class.

1. How does the kinetic-molecular theory explain why atmospheric pressure is greater at lower altitude than at a higher altitude?
2. Ice melts normally at 0°C. What happens to the melting point of ice in the presence of impurities? Does it melt below 0°C, above 0°C or exactly at 0°C? Explain.

What is the difference between evaporation and boiling? How does boiling point depend on the external pressure?

Boiling is the change of a liquid to bubbles of vapour that appear throughout the liquids. It is the conversion of liquid to vapour within the liquid as well as at its surface. It occurs when the equilibrium vapour pressure of the liquid equals the atmospheric pressure. During evaporation only molecules at the surface escape into the vapour phase, but at the boiling point the molecules within the liquid have sufficient energy to overcome the intermolecular attractive forces of their neighbors, so bubbles of vapour are released at the surface. It is the formation of vapour bubbles within the liquid itself that characterizes boiling and distinguishes it from evaporation.

If the temperature of the liquid is increased, the equilibrium vapour pressure also increases. Finally, the boiling point is reached. The boiling point of a liquid is the temperature at which the equilibrium vapour pressure of the liquid equals the atmospheric pressure. Therefore, the lower the atmospheric pressure, the lower the boiling point will be.

The boiling point of a liquid can be reduced as lowering the external pressure, because the vapour pressure of the liquid equals the external pressure at a lower temperature.

If the external pressure is 1.0 atmosphere (760 mmHg), the boiling point is called normal boiling point. For instance water boils, at 100°C at 1.0 atmospheric pressure. Thus, the normal boiling point of water is 100°C.

Where will water boil first? In Addis Ababa** or **Hawassa? Why?

Experiment 5.4



Observing the Vapour Pressure of liquid

Objective: To observe the vapour pressure of liquid.

Apparatus: Erlenmeyer flask, rubber bung, U-tube and burner.

Procedure:

1. Set up the apparatus as shown in Figure 5.8.
2. Add about 100 mL of water into the Erlenmeyer flask and put a stopper. Heat the flask to expel the air above the water in the flask.
3. Half fill the U-tube with water.
4. Connect the U-tube to Erlenmeyer flask and note the water level in the two arms of the U-tube.
5. Heat the flask gently and observe the water level changes in the arms of the U-tube.

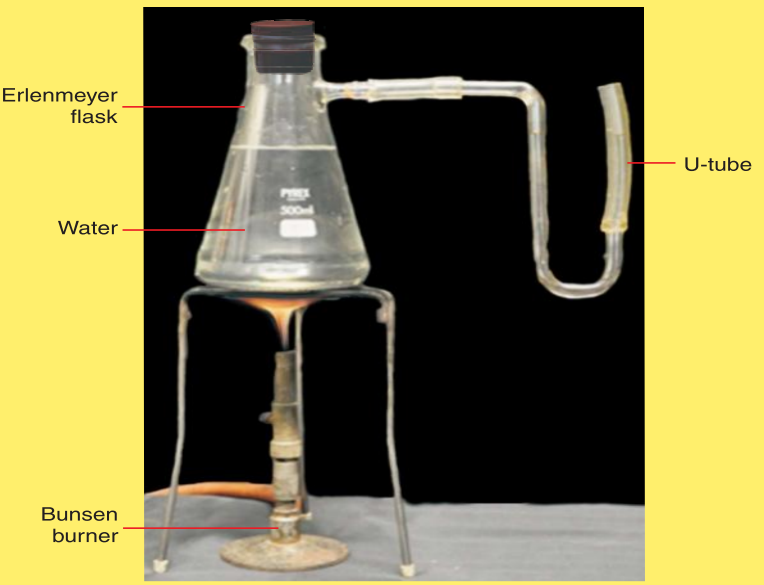


Fig. 5.8 Determination of vapour pressure.

Observations and analysis:

1. What do you observe? Give an explanation for the observation.

Experiment 5.5



Determining of Boiling Point

Objectives: To determine the boiling point of water.

Apparatus: Test tube, stopper, thermometer, beaker, burner, clamp, stand and base.

***Procedure*:**

1. Half fill the test tube with a sample of pure water and add some porclein chips.
2. Take a rubber stopper and pierce a thin opening on the side of the rubber stopper to allow the vapour to escape.
3. Fit the thermometer with the rubber stopper and insert it in the test tube.
4. Put the test test tube in a beaker containing oil as shown in Figure 5.9.
5. Heat the oil in beaker gently and record the temperature at which the water boils.

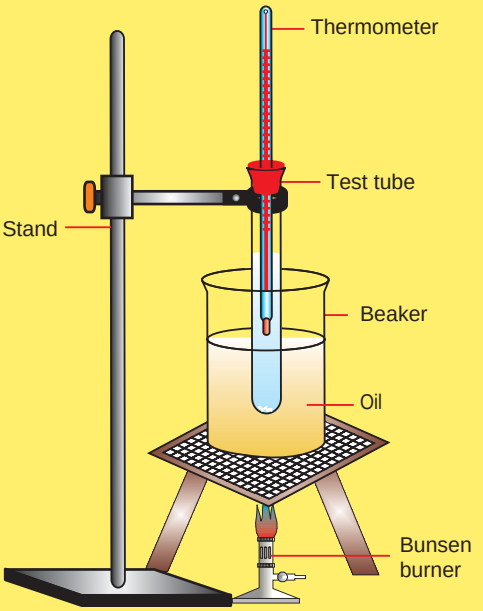


Figure 5.9 Determination of boiling point.

Observations and analysis:

1. What is the boiling point of water that you obtained from the experiment?
2. Does the temperature from the thermometer reading increase after the water starts to boil?
3. Explain why the thermometer was not put into the liquid.
4. Explain the purpose of adding porclein chips.

Boiling of a liquid requires a certain amount of heat energy to break the forces of attraction that holds the molecules together. The amount of heat energy necessary to bring about the vaporization of a fixed amount of a liquid at a fixed temperature to the gaseous state is called the heat of vaporization. For example, the heat of vaporization per mole of water at 298 K and 1 atmosphere is 44.0 kJ. This is called the molar heat of vaporization (ΔH_{vap}) of water.

Molar heat of vaporization (ΔH_{vap}) and molar heat of condensation (ΔH_{cond}) are equal in magnitude but opposite in sign, *i.e.*, $\Delta H_{vap} = -\Delta H_{cond}$. Vaporization is an endothermic processes where as condensation is an exothermic process.

Reading Check

1. What are the effects of impurity on the boiling point of liquids?
2. Why does the boiling point of liquid decrease as altitude increases?

5.5 THE SOLID STATE

Competencies

By the end of this section, you will be able to:

- explain the terms melting, fusion, sublimation, melting point, freezing point, molar heat of solidification;
- describe phase changes;
- explain temperature changes associated with phase changes;
- determine melting point of ice; and
- demonstrate an experiment to show the phase changes from ice to liquid water and then to water vapor.

Activity 5.18



Form a group and discuss the following:

1. When the crystals of iodine are warmed, they disappear into vapours without being changed into liquid.
2. When ethyl alcohol is taken in an open container it disappeares after sometime.

Present your discussion to the class.

When a solid is continuously heated its ordered crystalline structure is disturbed. The particles attain their freedom of motion gradually and melting (or fusion) takes place where the solid is converted into the liquid state. The temperature at which a crystalline solid is converted to a liquid is known as the melting point.

In contrast when a liquid is cooled, the molecules become closer and closer, and the intermolecular forces of attraction become stronger and stronger. The particles arrange themselves into a regular pattern in the liquid. This process is called freezing or solidification. Note that the freezing point of a liquid is the same as the melting point of a crystalline solid. This means that:

- Both melting and freezing take place at the same temperature.
- Both liquid and solid phases coexist at equilibrium with each other at the melting or freezing point.

To clarify this, let us consider ice, which melts at 0°C and water freezes at 0°C. Ice and water coexist in equilibrium with each other at 0°C as follows.

$$\text{Ice} \underset{\text{freezing, 0 °C}}{\overset{\text{melting, 0 °C}}{\rightleftharpoons}} \text{Liquid water}$$

The amount of heat needed to convert one gram of a solid to a liquid at the melting point is called heat of fusion. The molar heat of fusion or molar enthalpy of fusion (ΔH_{fus}) is the quantity of heat needed to convert one mole of a solid at its melting point to the liquid state. For example, the molar heat of fusion of ice is 6.01 kJ at 0°C. This is the amount of energy needed to break the attractive forces in the solid at the melting point. Melting requires the supply of energy; therefore, it is an endothermic process.

During the process of solidification, an amount of heat equal to the heat of fusion must be liberated. This quantity of heat liberated, which is exactly equal to the heat of fusion, is called the heat of solidification or heat of crystallization.

The molar heat of crystallization (ΔH_{cryst}) is the quantity of energy that must be removed from one mole of a liquid to convert it to the solid state at its freezing point.

$$\Delta H_{cryst} = -\Delta H_{fus}$$

Although the motion of the particles in a solid are more restricted than those in a liquid, many solids have a significant vapour pressure and evaporate directly from the solid to the vapour state without passing through the liquid state. This process is called sublimation.

$$\text{Solid} \underset{\text{deposition}}{\overset{\text{sublimation}}{\rightleftharpoons}} \text{Vapour (gas)}$$

During the sublimation process heat energy is absorbed. That is, it is an endothermic process.

Molar heat of sublimation (ΔH_{fus}) is the quantity of heat required to convert one mole of a solid to a gas at its sublimation point. The heat (enthalpy) of sublimation is related to the enthalpies of fusion and vaporization by:

$$\Delta H_{sub} = \Delta H_{fus} + \Delta H_{vap}$$

Phase Changes and Energy Changes in Solids

A phase is any part of a system that has uniform composition and properties. A state of matter represents a phase. Most solid substances undergo two changes of state when heated. A solid change to a liquid at the melting point, and the liquid changes to vapour at the boiling point. To understand state changes, we will consider the heating curve for a substance given in Figure 5.10.

A heating curve is a plot of temperature verses the uniform addition of heat. This can be illustrated for a hypothetical substance, in which the temperature of the substance is on the vertical axis and the passage of time during which heat is added to the substance is on the horizontal axis. Figure 5.10 shows the changes in the temperature and phases of a pure substance as it is heated, beginning with a solid and continuing to the gaseous state as described.

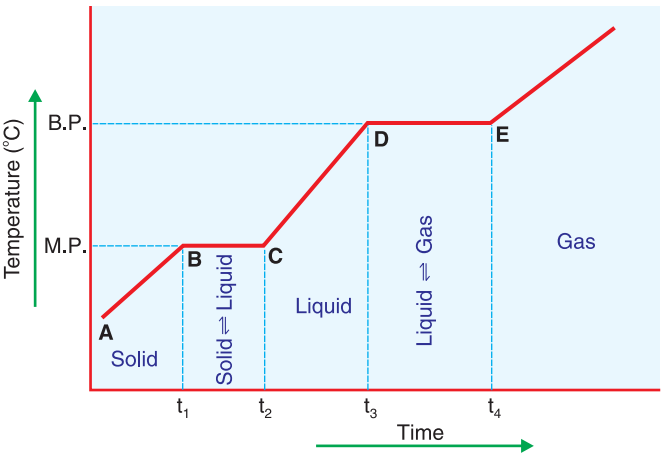


Figure 5.10 Heating curve.

Initially, the substance exists in the solid state, and the addition of heat increases its temperature. When the solid is heated, its temperature rises (A to B) until it reaches the melting point (point B), and the temperature remains constant (B to C) until all the

solid is converted to a liquid (point C). The added heat energy is used to break the intermolecular forces, thus disrupting the solid structure. At point C phase change is completed. Once melting is completed, heating of the liquid raises its temperature (C to D) until the boiling point is reached at point D. In region (D to E) the addition of heat is utilized to break the intermolecular forces of the liquid to change it to a gas. When all the liquid has been converted into a gas, addition of heat simply raises the temperature of the gas.

Reading Check

Freezing and melting point of ice takes place at 0°C. Rationalize this.

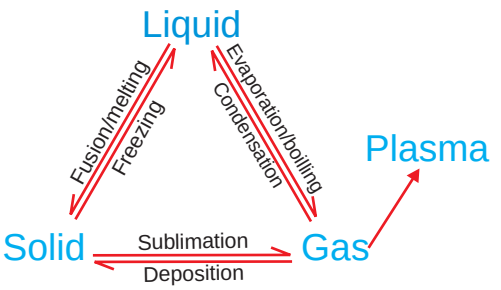
Check list

Key terms of the unit

- *Absolute zero*
- *Atmospheric pressure*
- *Avogadro's law*
- *Boiling*
- *Boiling point*
- *Boyle's law*
- *Charles' law*
- *Collision*
- *Combined gas law*
- *Condensation*
- *Evaporation*
- *Fluid*
- *Freezing*
- *Freezing point*
- *Gas constant*
- *Gas laws*
- *Gases*
- *Graham's law*
- *Heat of condensation*
- *Heat of vaporization*
- *Heating curve*
- *Ideal gas equation*
- *Kelvin temperature*
- *Kinetic molecular theory of gases*
- *Liquids*
- *Melting*
- *Melting point*
- *Molar heat of condensation*
- *Molar heat of fusion*
- *Molar heat of sublimation*
- *Non-volatile liquid*
- *Normal boiling point*
- *Phase change*
- *Properties of gases*
- *Solidification*
- *Standard temperature and pressure (STP)*
- *Sublimation*
- *Vapor pressure*
- *Volatile liquid*

Unit Summary

- *Matter is anything that has mass and occupies space.*
- *Matter exists in the form of a solid, a liquid or a gas.*
- *Solids have a definite volume and a definite shape.*
- *A liquid has no definite shape, it takes the shape of its container.*
- *A gas has neither a definite volume nor a definite shape.*
- *Gases and liquids are fluids where as solids are not. Fluidity is the tendency to flow.*
- *When energy is supplied to a solid, it melts and changes to a liquid, the particles move faster. Additional energy will make the liquid boil and form a gas.*
- *In the gas the particles are much more widely spaced and move much faster than in liquid and solid.*
- *Some solids directly change to gases by the process of sublimation.*
- *Phase change can be illustrated as:*



- *General (combined) gas equation:* $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
- *Ideal gas equation:* $PV = nRT$
- *Grahm's Law of diffusion:* $\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}}$
- *Melting point = Freezing point*
- *Boiling point = Condensation point*
- *Sublimation point = Deposition point.*

REVIEW EXERCISE ON UNIT 5

Part I: Assert your answers based on the given instructions:

We have three physical states of substances namely solid, liquid, and gas. Identify the following as solid, liquid or gas based on the properties given to explain their characteristics.

1. They have a definite shape and a definite volume.
2. Their molecules are highly disordered.
3. The motion of their molecules is highly restricted.
4. They can be easily compressed.
5. They have a tendency to flow.
6. They can take the shape of their container.
7. They can move in all direction at high speed.
8. They can easily diffuse through each other.
9. They can sublime.
10. They have less density relative to the other states.

Part II: Matching-type questions

A	B
11. Melting point	**A** Solid → gas
12. Heat of fusion	**B** Liquid → gas
13. Heat of sublimation	**C** Gas → solid
14. Sublimation point	**D** Solid → liquid
15. Melting	**E** Liquid → solid
16. Freezing	**F** Same as freezing point
17. Sublimation	**G** Same as deposition point
18. Deposition	**H** Same as heat of crystallization
	I Similar to fusion
	J Similar to crystallization
	K Similar to heat of deposition

Part III: Write the missing words in your exercise book

19. In a solid the particles are very close together and can only _____ about a fixed position.
20. Sublimation occurs when a solid changes directly to a _____ passing the _____ state.
21. The melting point of a solid is the same as _____.
22. When the water boils, its vapour pressure is equals to _____.
23. The temperature at which a crystalline solid is converted to a liquid, is called _____.
24. The lowest attainable temperature is _____.
25. "Equal volumes of different gases at the same temperature and pressure contain equal numbers of molecules"; this is a statement of _____ law.

Part IV: Short-answers type questions

26. What is the difference between volatile and non-volatile substances? Give an example of each.
27. Why is the temperature of a substance constant at its melting point; even though heat is added to it?
28. What factors affect the rate of evaporation of a liquid?

Part V: Problems to solve

29. Convert the following pressure measurements:
    a 720 mmHg to atm
    b 1.25 atm to mmHg
    c 542 mmHg to atm
    d 740 mmHg to kPa
    e 700 kPa to atm
30. A 2.50 L container is filled with 175 g of argon:
    a if the pressure is 10.0 atm, what is the temperature?
    b if the temperature is 22 K, what is the pressure?
31. If 0.500 mole of nitrogen gas occupies a volume of 11.2 L at 0°C; what volume will 2.00 mole of nitrogen gas occupy at the same temperature and pressure?
32. A certain gas is found in the exhaust of automobiles and power plants. Consider a 1.53 L sample of a gas at a pressure of 5.6×10^4 Pa. If the pressure is changed to 1.5×10^4 Pa at a constant temperature, what will be the new volume of the gas?

33. What mass of helium in grams is required to fill 15.0 L balloon to a pressure of 1.1 atm at 25°C?

34. A sample of a certain gas has a volume of 1.25 L at –125°C and 5.0 atm. The gas is compressed at 50.0 atm to a volume of 325 mL. What is the final temperature of the helium gas in °C?

35. If 140 cm^3 of methane diffuse into air in 72 seconds, how long will it take 210 cm^3 of sulphur dioxide to diffuse under the same conditions?

36. The temperature of 2.5 L of a gas initially at STP is raised to 250°C at constant volume; calculate the final pressure of the gas in atmosphere.

37. A sample of air occupies 3.8 L when the pressure is 1.2 atm, assuming the temperature remains constant:

    a what volume does it occupy at 6.6 atm?

    b what pressure is required in order to compress it to 0.075 L?

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