

CHEMISTRY

A Textbook for Grade 10



C10TB

AUTHORS

Dr. Geetanjali

Professor

Department of Chemistry, Kirori Mal College, University of Delhi

Dr. Ram Singh

Professor

Department of Applied Chemistry, Delhi Technological University

REVIEWERS

Dominic Dugbe Doe

Chemistry Specialist

Kandakai Massaquoi

Chemistry Specialist

Patrick A. Anderson

Sr. Director

Division of Technical and Vocational Education, MoE



Ministry of Education
Monrovia, Republic of Liberia



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Foreword

Liberia, having gone through a period of utmost turmoil till 2003, due to the civil wars, is still reeling under its effect and the added trauma of Ebola in 2014 and effects of the COVID-19 outbreak in 2020. The Liberian government, in the past decade, has made valiant efforts to bring order to the lives of its people. In one such effort, the Ministry of Education (MoE) brought changes to the National Curriculum Framework which are relevant to the present generation, and which would prepare them to meet the challenges of the changing trends of the world. The National Curriculum Framework (NCF) 2018 recommends a change in basic assumptions in the teaching learning process from behaviorist to constructivist approach — moving from hardcore print material to the digital world. Keeping in consideration the sociocultural context and varied experiences of learners as laid down in the Framework, our Teaching Learning Materials are expected to be competent to use multiple methods and techniques like e-learning resources, energized textbooks, and readily available reference material to engage the learners.

As a first initiative, the MoE, through its World Bank-funded Improving Results in Secondary Education (IRISE) project, has adapted textbooks for Grades 10 to 12 in five subjects — English Language and Literature, Mathematics, Biology, Physics and Chemistry.

The National Curriculum Framework, 2018, recommends that children’s learning at school is a reflection of their life outside the school and shows them the path to become a responsible citizen who makes knowledge-based choices. This principle marks a departure from the legacy of teacher centered learning to student centered learning. The syllabi and textbooks developed on the basis of the NCF indicate a serious attempt to implement the idea of Activity Base Learning (ABL). We hope these measures will take us ahead in the direction of building a system of education as outlined in the NCF.

Combined with the efforts by the school principals and teachers this will encourage children to reflect on their own learning and to pursue imaginative activities and questions. With this in mind, perhaps for the first time in our country, we are able to provide separate subject specific textbooks accompanied with guides for teachers for 10–12 grades. Not only have these been developed, adapted and modified to the Liberian context, each of the eight Minimum Learning Competencies (MLCs) have been included in each textbook. So as to reach every high school student, for the first time in the country’s history we have included the digitized form of the textbook accessible by a Quick Response (QR) code given in each book. Not only does it have the digitized textbook, but it provides additional learning materials for use by students, teachers and interested persons. The links to these e-resources and digitized material is being made available on the MoE’s website.

The Textbooks and Teacher Guides have reached the hands of the students after a rigorous quality evaluation by carefully handpicked subject specialists by the MoE, to whom the Ministry expresses gratitude. For the success of this project, I acknowledge the contributions of the IRISE Project Team in the World Bank, and in particular, the Task-Team Leaders; the Project Implementation Team in Liberia headed by its Coordinator Abraham A. Kiazolu II, supported by the Executive Director of the Center of Excellence for Curriculum Development and Textbooks Research, Mrs. Julia K. Sandiman-Gbeyai and her technical working group (TWG), and the International Textbook Consultant and Advisor, Dr Shveta Uppal engaged by the MoE. These notwithstanding would not have been possible without the guidance of the Senior Management Team (SMT) of the Ministry of Education, and in particular, the Deputy Ministers for Instructions, Administration, and Planning, Research and Development, respectively.

Professor Dao Ansu Sonii, Sr.
Minister of Education
Republic of Liberia

Monrovia, Republic of Liberia
January 24, 2023

Acknowledgments

The development of textbooks contributes to the quality of teaching and learning that go on in the classroom.

The Ministry of Education (MoE) has aligned its Curriculum for Grades 10–12 to the National Curriculum Framework (NCF) of 2018. To ensure the provision of Teaching Learning Materials (TLMs) that support the revised curriculum, the Ministry has sought, reviewed and adapted a new set of textbooks and teacher guides along with digitized contents and e-learning resources for the five core subjects taught at the Senior Secondary education level, namely English Language and Literature, Mathematics, Biology, Chemistry and Physics, through an internationally competitive bidding process from the market supported by the World Bank funded Improving Results in Secondary Education (IRISE) Project.

With profound gratitude and honor, we recognize the Senior Management Team of the Ministry, headed by the Coach, Professor D. Ansu Sonii, Sr., for the strategic decision to make teaching learning materials available and accessible to all in the Liberian Senior Secondary School System, and for providing directions through the process of securing these textbooks and other teaching learning materials for our students and teachers. Our special thanks and appreciation to the World Bank for the financial support towards this policy intervention, and its education task-team including Alonso Sanchez, Oni Lusk-Stover and Binta B. Massaquoi for all their technical inputs offered throughout the process to ensure the kind of quality TLMs the Liberian students deserve are made available for improved learning outcomes.

We would like to specifically recognize the invaluable contributions of the 15 subject experts selected by the MoE from across the various education systems and the West African Examinations Council (WAEC) to evaluate, review and sign off on these teaching learning materials. They didn't just deliver according to our expectations, but also ensured the contextual relevance of the materials to the Liberian Secondary

Education Curriculum and its minimum learning competencies (MLCs). These subject experts include Professor Isaac Saye-Lakpoh Zawolo – *Superintendent* of the Monrovia Consolidated School System (MCSS), Mr. Matthew V.Z. Darblo, Sr. – *Mathematics Instructor* at the University of Liberia (UL), Mr. Charles Tieh Bropleh – *Mathematics Specialist* (MoE), Mrs. Linda Y. Dean – *English Specialist*, Mr. Hassan M. Bangura *English Language and Literature Expert*, Mr. J. Emmanuel Milton *English Specialist* (MoE), Mr. Moses K.M. Togbah – *Physics Specialist*, Mr. Prince A. Dossen – *Physics Specialist*, Mr. Benjamin Koryah – *Physics Instructor* at the University of Liberia (UL), Mr. Dominic Dugbe Doe – *Chemistry Specialist*, Mr. Patrick A. Anderson, Sr. – *Director* of the Division of Technical and Vocational Education (MoE), Mr. Kandakai Massaquoi – *Chemistry Specialist*, Ms. Patricia N. Doe – *Head* of Biology Department, African Methodist Episcopal University (AMEU), Mr. Job Carpenter – *Biology Specialist* and Mr. Prince Philip K.A. Aderibigbe – *Biology Specialist*.

The MoE is sincerely grateful to Dr Shveta Uppal, the *International Textbook Consultant* engaged by the IRISE Project to provide technical guidance and quality assurance support to the revising of the Textbooks Management Guidelines (TMG) and the procurement process leading to the provision of textbooks, teacher guides, digital contents and e-learning resources for the Senior Secondary School System in Liberia in accordance with the revised TMG. Heartfelt thanks and appreciations also to the *Executive Director* for the Center of Excellence for Curriculum Development and Textbooks Research, Mrs. Julia K. Sandiman-Gbeyai, and members of her Technical Working Group (TWG) for taking up the responsibility to lead the process of making textbooks and other TLMs available to Liberian students and teachers.

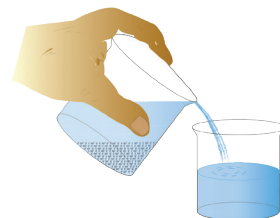
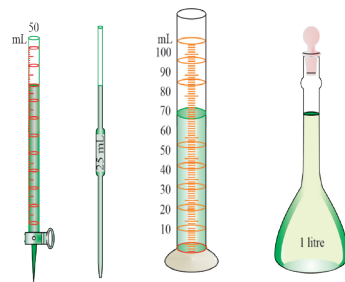
Lastly, we acknowledge the IRISE Project Delivery Team led by Mr. Abraham A. Kiazolu, II – *Project Coordinator*, Mr. Fuseini A. Abu – *International Procurement Specialist* and Mr. Lawrence S. Taylor – *Project Control Specialist* who coordinated the entire process.

We remain grateful to you all!

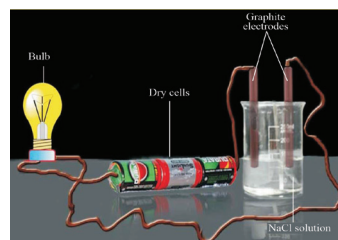
Hon. Alexander N. Duopu, Sr.,
Deputy Minister for Instruction
Ministry of Education, Republic of Liberia
#The Teacher

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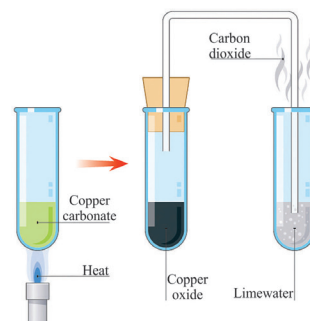


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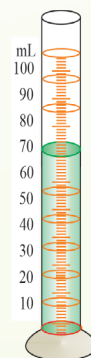
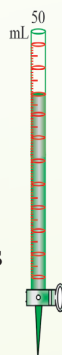
CHAPTER

1

INTRODUCTION TO CHEMISTRY

Chapter Contents

- 1.1 Historical Development of Chemistry
- 1.2 Definition and Branches of Chemistry
- 1.3 The Scientific Method
- 1.4 Units of Measurements
- 1.5 Chemical Symbols, Formula and Naming compounds
- 1.6 Laboratory apparatus and safety measures
 - Key Terms
 - Summary
 - Exercises



Chapter Outcomes

Upon completion of this chapter, students will be able to:

- demonstrate knowledge of the origins and various stages in the development of chemistry;
- express appreciation for the scientific method;
- explain the word Chemistry and other related terminologies;
- distinguish the systems of units of measurement;
- solve simple conversion problem;
- discuss the origin of symbols of element;
- apply the symbols to write the formula and the names of compounds;
- apply the laboratory safety rules and;
- identify apparatus in the Laboratory.

1.1 HISTORICAL DEVELOPMENT OF CHEMISTRY

ACTIVITY 1

Form a group and perform the following tasks.

Collect the following materials which are commonly found in your environment and bring them to your class : common salt, water, sand, charcoal, ash, paper, sugar, chalk and soap. Observe the forms, colors, odors of each material. Also notice what you feel while rubbing it between your thumb and forefingers.

1. How do these materials differ from each other?
2. Why does each material differ from the others? After the discussion share your ideas with the rest of the class

Contributors of Chemistry

It was not until the era of the ancient Greeks that we have any record of how people tried to explain the chemical changes they observed and used. At that time, natural objects were thought to consist of only four basic elements: earth, air, fire, and water. Then, in the fourth century BC, two Greek philosophers, Democritus and Leucippus, suggested that matter was not infinitely divisible into smaller particles but instead consisted of fundamental, indivisible particles called **atoms**.



Democritus

Over the next two millennia, alchemists, who engaged in a form of chemistry and speculative philosophy during the Middle Ages and Renaissance, achieved many advances in chemistry. In particular, alchemists wanted to find a way to transform cheaper metals into gold. Although most alchemists did not approach chemistry systematically and many appear to have been outright frauds, alchemists in China, the Arab kingdoms, and medieval Europe made major contributions, including the discovery of elements such as quicksilver (mercury) and the preparation of several strong acids.

Modern Chemistry

The 16th and 17th centuries saw the beginnings of what we now recognize as modern chemistry. During this period, great advances were made in *metallurgy*, the extraction of metals from ores, and the first systematic quantitative experiments were carried out. In 1661, the Englishman Robert Boyle (1627–91) published *The Sceptical Chymist*, which described the relationship between the pressure and the



Robert Boyle

volume of air. More important, Boyle defined an element as a substance that cannot be broken down into two or more simpler substances by chemical means. This led to the identification of a large number of elements, many of which were metals. Ironically, Boyle himself never thought that metals were elements.

In the 18th century, the English clergyman Joseph Priestley (1733–1804) discovered oxygen gas and found that many carbon-containing materials burn vigorously in oxygen



Joseph Priestley

present in the atmosphere. This process is called combustion. Priestley also discovered that the gas produced by fermenting beer, which we now know to be carbon dioxide, is the same as one of the gaseous products of combustion. Priestley's studies of this gas did not continue as he would have liked, however. After he fell into a vat of fermenting beer, brewers prohibited him from working in their factories. Although Priestley did not understand its identity, he found that carbon dioxide dissolved in water to produce seltzer water. In essence, he may be considered the founder of the multibillion-dollar carbonated soft drink industry.



Antoine Lavoisier

Despite the pioneering studies of Priestley and others, a clear understanding of combustion remained elusive. In the late 18th century, however, the French scientist Antoine Lavoisier (1743–94) showed that combustion is the reaction of a carbon-containing substance with oxygen to form carbon dioxide and water and that life depends on a similar reaction, which today we call *respiration*. Lavoisier also wrote the first modern chemistry text and is widely regarded as the father of modern chemistry.

His most important contribution was the law of conservation of mass, which states that in any chemical reaction, the mass of the substances that react equals the mass of the products that are formed. That is, in a chemical reaction, mass is neither lost nor destroyed.

Can you mention some other scientists who contribute to the development of chemistry?

1.2 DEFINITION AND BRANCHES OF CHEMISTRY

Chemistry is the science that deals with matter and the changes that it undergoes. It is a study of the composition, structure, and properties of matter and of the changes that occur in matter.

Perhaps the only permanent thing in the world is change. Iron rusts, snow melts, paints peel off and fire woods burn. We grow up, we grow old. Living plants and animals undergo ceaseless change, and even dead animals and plants continue to change as they decay. Such changes fascinated people and inspired them to look more closely at nature's way of working.

Understanding change is closely related to understanding the nature and composition of matter—the physical material of the universe. Matter is anything that occupies space and has mass.

It has long been known that matter can change or be made to change from one form to another. These changes are broadly classified into **chemical** and **physical changes**. Chemical changes, more commonly called as chemical reactions are processes whereby one substance is transformed into another as a result of combination or dissociation of atoms. We can describe the transformation both qualitatively and quantitatively with the help of chemical equations for the reaction. Some of the examples of chemical change include oxidation of matter (rusting, burning), fermentation, changing milk into yogurt, and addition of water to calcium oxide.

Matter also undergoes other kinds of changes called **physical changes**. These changes differ from chemical reactions in that the involved substances do not change their identities. Each retains its composition. Most physical changes are accompanied by changes in physical state, such as the melting of solids and the boiling of liquids. For example, water remains H_2O whether it is in solid state (ice), liquid water or gaseous state (steam). Physical change also involves making or separating mixtures. Dissolving table salt (NaCl) in water is a physical change.

ACTIVITY 2

Form a group and discuss the general concepts of the following phrases and the interrelationship between them.

1. Natural science
2. Study of materials
3. Behavior of materials
4. Study of nature
5. From your discussion define:
 - (a) natural science
 - (b) chemistry

Branches of chemistry

The study of modern chemistry has many branches, but can generally be broken down into five main disciplines, or areas of study. These are:

- (i) **Physical chemistry:** Physical chemistry is the study of macroscopic properties, atomic properties, and phenomena in chemical systems. A physical chemist may study such things as the rates of chemical reactions, the energy transfers that occur in reactions, or the physical structure of materials at the molecular level.
- (ii) **Organic chemistry:** Organic chemistry deals with the study of substances containing carbon. Carbon is one of the most abundant elements on Earth and is capable of forming a tremendously vast number of chemicals (over twenty million so far). Most of the chemicals found in all living organisms are based on carbon.
- (iii) **Inorganic chemistry:** Inorganic chemistry is the study of substances that, in general, are not primarily based on carbon. Inorganic chemicals are commonly found in rocks and minerals. One current important area of inorganic chemistry deals with the design and properties of materials involved in energy and information technology.
- (iv) **Analytical chemistry:** Analytical chemistry involves the study of the composition of matter. It focuses on separating, identifying, and quantifying chemicals in samples of matter. An analytical chemist may use complex instruments to analyze an unknown material in order to determine its various components.
- (v) **Biochemistry:** Biochemistry is the study of chemical processes that occur in living things. Research may cover anything from basic cellular processes up to understanding disease states so that better treatments can be developed.

1.3 THE SCIENTIFIC METHOD

The scientific method is a process for gathering data and processing information. It provides well-defined steps to standardize how scientific knowledge is gathered through a logical, rational problem-solving method. Figure 1 shows the steps of the scientific method.

Making observations. Observations may be qualitative (the sky is blue; water is a liquid) or quantitative (water boils at 100°C; a certain chemistry book weighs 2 kilograms). A qualitative observation does not involve a number. A quantitative observation (called a measurement) involves both a number and a unit.

Formulating hypotheses. A hypothesis is a possible explanation for an observation.

Performing experiments. An experiment is carried out to test a hypothesis. This involves gathering new information that enables a scientist to decide whether the hypothesis is valid—that is, whether it is supported by the new information learned from the experiment. Experiments always produce new observations, and this brings the process back to the beginning again.

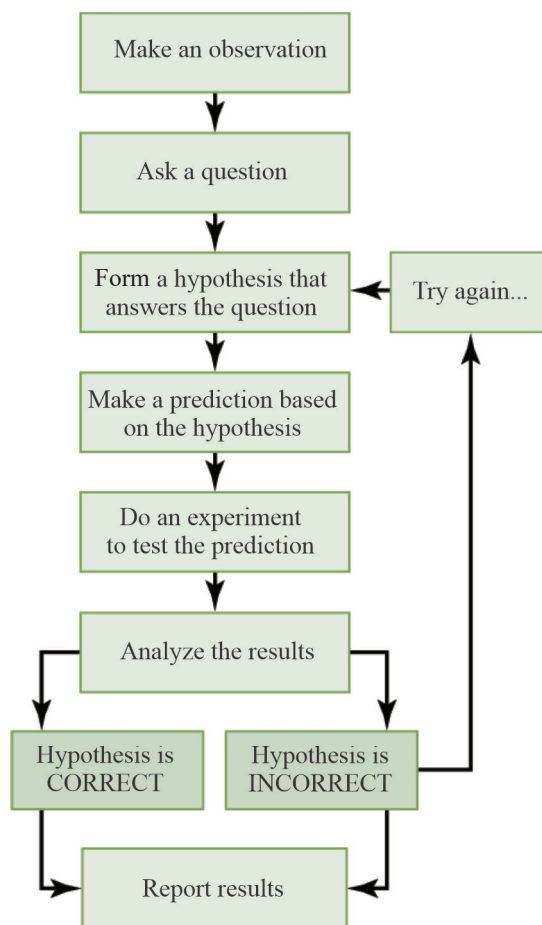


Figure 1. Steps in Scientific Method

ACTIVITY 3

- Form a group and imagine/enact that you are a group of scientists who have discovered a new chemical compound of great use in curing cancer.
 - How would you present your work to other scientists of the world?
 - Is there a particular format to write scientific reports? Discuss the essential features of a well-designed experiment.
- Consider the following:
 - Strike a match stick to light it.
 - Record all your observations in writing.
 - Examine your written observations and consider their objectivity.

Which of these do you think are just descriptions of your observations? In contrast, which are the ideas that you formed based on your observations? Finally, which of your data could be subjective or partly subjective? Why?

Present your observations to your classmates. As you do so, describe the objectivity or subjectivity of each statement that you make. In particular, discuss those that you inferred based on observation.

Scientists seek general relations that unify their observations. A concise verbal statement or a mathematical equation that summarizes a broad variety of observations and experience is known as *scientific law*. A familiar example is the law of gravity. It summarizes the experience that what goes up must come down.

Scientists also seek to understand laws. A tentative explanation of a law is called a *hypothesis*. A hypothesis is useful if it can be used to make predictions that can be tested by further experiments and can thereby be verified.

A hypothesis that continually withstands such tests is called a *theory*. A theory is an explanation of the general principles of certain phenomena that has considerable evidence or facts to support it. It may serve to unify a broad area and may provide a basis for explaining many laws.

There is no fool-proof, step-by-step scientific method that people use. Their approaches depend on their temperaments, circumstances and training. Rarely will two people approach the same problem in the same way. Scientific progress is not smooth, certain and predictable. The path of any scientific study is likely to be irregular and uncertain. Progress is often slow, and many promising leads turn out to be dead ends. Serendipity (fortunate accidental discovery) as well as perseverance has played an important role in the development of science.

1.4 UNITS OF MEASUREMENTS

ACTIVITY 4

Form a group and perform the following activity.

Take a small piece of magnesium ribbon. Measure its mass and length. Now put it in 20 mL of dilute hydrochloric acid and measure the time required for magnesium ribbon to dissolve completely. Record the temperature of solution before and after putting the magnesium ribbon in it. Discuss the following questions:

1. What instruments/equipments did you use for measuring the physical quantities?
2. What units did you use to express them?

3. Which is the appropriate unit to express the time taken for the above reaction to go to completion?
4. Explain the difference between heat and temperature.
5. Which basic SI units are appropriate to express the: a) length of a race track, b) average room temperature, and c) time duration for the earth to have one rotation around its axis?

Systems of Units

Making observations is fundamental to all science. A quantitative observation, or measurement, always consists of two parts: a number and a scale (called a unit). Both parts must be present for the measurement to be meaningful.

In this textbook we discuss measurements of mass, length, time, temperature, and volume of a substance, among others. Scientists recognized long ago that standard systems of units had to be adopted if measurements were to be useful. If every scientist had a different set of units, complete chaos would result. Unfortunately, different standards were adopted in different parts of the world. The two major systems are the English system used in the United States and the Metric System used by most of the rest of the industrialized world.

In 1960, an International agreement set up a system of units called the International System (*le Système International* in French), or the SI system. This system is based on the metric system and units derived from the metric system. All other SI units of measurement can be derived from these base units. The fundamental SI units are listed in Table 1.

Table 1 The fundamental SI units

Property	Name of Unit	Abbreviation
Length	meter	m
Mass	kilogram	kg
Time	second	s
Amount of substance	mole	mol
Temperature	kelvin	K
Electric current	ampere	A
Luminous intensity	candela	Cd

We use prefixes to indicate decimal multiples or fractions of the base units. The international system uses a series of prefixes to indicate decimal fractions or multiples of various units by powers of 10. All numbers can be expressed in the form of $a \times 10^b$, where ‘a’ is a number between 1 and 10, and the exponent ‘b’ is an integer. This feature makes it easy to convert from one unit to another. Some of the most commonly encountered prefixes in chemistry are listed in Table 2.

Table 2 Some common SI prefixes

Prefix	Meaning	Symbol	Multiple/Fraction
tera	trillion	T	10^{12}
giga	billion	G	10^9
mega	million	M	10^6
kilo	thousand	k	10^3
deci	tenth of	d	10^{-1}
centi	hundredth of	c	10^{-2}
milli	thousands of	m	10^{-3}
micro	millionth of	μ	10^{-6}
nano	billionth of	n	10^{-9}
pico	trillionth of	p	10^{-12}

When we solve numerical problems, we use an approach to units called dimensional analysis. Dimensional analysis was developed to ensure that our answers yield proper units. It also offers a systematic approach to solve numerical problems and check our solutions for possible errors.

In dimensional analysis, we carry units through all calculations. As we work, we multiply units together, divide them by each other, and ‘cancel’ them.

The key to use dimensional analysis is the correct use of conversion factors in order to change one unit into another. A conversion factor is a fraction whose numerator and denominator are the same physical quantity expressed in different units. For example, 100 cm and 1 m are the same length, $100 \text{ cm} = 1 \text{ m}$. This relationship allows us to write two conversion factors:

$$\frac{100\text{cm}}{1\text{m}} \text{ and } \frac{1\text{m}}{100\text{cm}}$$

The first of these factors is used when we want to convert meters to centimeters and second centimeters to meters. For example, the length in centimeters of an object that is 8.50 m long is given by:

$$\text{Number of centimetres} = 8.50 \text{ m} \times \frac{100 \text{ cm}}{1 \text{ m}} = 850 \text{ cm}$$

Note that the unit of meter in the denominator of the conversion factor cancels the unit of meter in the measurement given (8.50 m). The centimeter in the numerator of the conversion factor becomes the unit of the final answer. In general, the units multiply and divide as follows:

$$\text{given unit} \times \frac{\text{desired unit}}{\text{given unit}} = \text{desired unit}$$

If the desired units are not obtained in a calculation, then we know that we made an error somewhere. Careful inspection of units often reveals the source of the error.

Example

If you had a sample of a substance with a mass of 0.0034 g, and you wanted to express that mass in mg, you could use the following dimensional analysis. The given quantity is the mass of 0.0034 g. The quantity that you want to find is the mass in mg, and we know that 1 mg = 10^{-3} g. Expressing this as a proportion or ratio, there is one mg per 10^{-3} grams, or 1000 mg/1 g.

$$\text{Therefore, } 0.0034\text{g} \times (1000 \text{ mg}/1 \text{ g}) = 3.4 \text{ mg}$$

Exercise

1. A piece of aluminium foil is 8.0×10^{-5} cm thick. What is its thickness in micrometers?
2. Convert 75 ng to mg.
3. Convert 6.75 m^3 to μL .
4. Convert each of the following measurements to a unit that replaces the power of ten by a prefix.
 - (a) 3.22×10^{-6} s
 - (b) 9.56×10^{-3} m
 - (c) 1.07×10^3 g
5. Calculate the mass in grams of two cubic inches (2.00 in^3) of gold. Density of gold = 19.3 g cm^{-3}

Basic SI Units

Mass

Mass of an object is the amount of matter present in it. It is measured with an analytical balance and in contrast to weight, mass is not affected by gravity. The basic unit of mass in the metric system is gram.

Length

The SI base unit of length is the meter (m). To measure length much larger than the meter, we often use the kilometer (km). In the laboratory, lengths smaller than a meter are often most convenient. For example, the centimeter (cm) and the millimeter (mm). On the submicroscopic scale, the micrometer (μm), the nanometer (nm), etc., are used.

Time

The SI base unit for measuring intervals of time is the second (s). Short times are expressed through the usual SI prefixes: milliseconds (ms), microsecond (μs), nanosecond (ns), and picosecond (ps). Long time intervals, on the other hand, are usually expressed in traditional, non-SI units: minute (min), hour (h), day (d), and year (y).

Temperature

Temperature is a measure of the average energy of motion or kinetic energy, of a single particle in a system. The instrument for measuring temperature is called thermometer. From common experience, we know that if two objects at different temperatures are brought together, heat flows from the warmer to the colder object. For example, if you touch a hot test tube, heat will flow from the test tube to your hand. If the test tube is hot enough, your hand will get burned. The temperature of the warmer object drops and that of the colder object increases, until finally the two objects are at the same temperature (thermal equilibrium). Temperature is therefore a property that tells us in what direction heat flows.

The SI basic unit of temperature is the Kelvin (K). For most routine laboratory work, we can use a more familiar temperature scale: the Celsius scale. On this temperature scale, the freezing point of water is 0°C , and its boiling point is 100°C . Another temperature scale, probably unfamiliar to most people, is the Fahrenheit scale. The relationship between these three temperature scales is given below:

$$[^{\circ}\text{C}] = ([^{\circ}\text{F}] - 32) \times 5/9 \quad \text{or} \quad [^{\circ}\text{F}] = [^{\circ}\text{C}] \times 9/5 + 32$$

$$\text{K} = ^{\circ}\text{C} + 273.15$$

The Kelvin scale assigns a value of zero Kelvin (0K) to the lowest possible temperature, which is called **absolute zero** and corresponds to -273.15°C . Note that the term absolute zero is used because this is a hypothetical temperature characterized by complete absence of thermal (kinetic) energy.

Example

- Solder is an alloy made of tin and lead that is used in electronic circuits. A certain solder has a melting point of 224°C . What is its melting point in degrees Fahrenheit?
- Helium has the lowest boiling point of all the elements which is 2452°F . Convert this temperature to $^{\circ}\text{Celsius}$.
- Mercury, the only metal that exists as a liquid at room temperature, melts at -38.9°C . Convert its melting point to kelvins.

Solution

These three parts require that we carry out temperature conversions.

- This conversion is carried out by writing

$$9^{\circ}\text{F} / 5^{\circ}\text{C} \times (224^{\circ}\text{C}) + 32^{\circ}\text{F} = 435^{\circ}\text{F}$$
- $[^{\circ}\text{C}] = ([^{\circ}\text{F}] - 32) \times 5/9$

$$= (-452^{\circ}\text{F} - 32) \times 5^{\circ}\text{C} / 9^{\circ}\text{F} = -269^{\circ}\text{C}$$
- The melting point of mercury in kelvins is given by

$$\text{K} = ^{\circ}\text{C} + 273.15$$

$$= -38.9^{\circ}\text{C} + 273.15 = 234.3\text{K}$$

Exercise

Convert

- 327.5°C (the melting point of lead) to degrees Fahrenheit;
 - 172.9°F (the boiling point of ethanol) to degrees Celsius; and
 - 77 K, the boiling point of liquid nitrogen, to degrees Celsius.
- The average temperature in a certain city, during the summer, is about 25°C . What is the equivalent Kelvin temperature?

3. A parasite that causes trichinosis is killed when meat is cooked to 66°C . Assume you have only a Fahrenheit thermometer. Determine the minimum Fahrenheit temperature to which the meat should be heated when it is being cooked

Mole (Amount of substance)

Chemists use the term mole to represent a large number of small particles such as atoms molecules etc. Just as a dozen implies 12 things, a mole (mol) represents 6.022×10^{23} things. The number 6.022×10^{23} , called Avogadro's number after the 19th-century chemist named Amedeo Avogadro. This is the number used in chemistry to represent macroscopic amounts of atoms and molecules.

Thus, if we have 6.022×10^{23} oxygen atoms, we say we have 1 mol of oxygen atoms. If we have 2 mol of sodium atoms, we have $2 \times (6.022 \times 10^{23})$ sodium atoms, or 1.2044×10^{24} sodium atoms. Similarly, if we have 0.5 mol of benzene molecules, we have $0.5 \times (6.022 \times 10^{23})$ benzene molecules, or 3.011×10^{23} C_6H_6 molecules.

Derived SI Units

Derived units are combinations of SI base units. Units can be multiplied and divided, just as numbers can be multiplied and divided. For example, the area of a square having a side of 2 cm is $2 \text{ cm} \times 2 \text{ cm}$, or 4 cm^2 (read as “four centimeters squared” or “four square centimeters”). Notice that we have squared a length unit, the centimeter, to get a derived unit for area, the square centimeter.

Volume

Volume is an important quantity that uses a derived unit. Volume is the amount of space that a given substance occupies and is defined geometrically as length \times width \times height. Each distance can be expressed using the meter unit, so volume has the derived unit $\text{m} \times \text{m} \times \text{m}$, or m^3 (read as “meters cubed” or “cubic meters”). A cubic meter is a rather large volume, so scientists typically express volumes in terms of 1/1,000 of a cubic meter. This unit has its own name—the liter (L). A liter is $1,000 \text{ cm}^3$. By definition, there are 1,000 mL in 1 L, so 1 milliliter and 1 cubic centimeter represent the same volume.

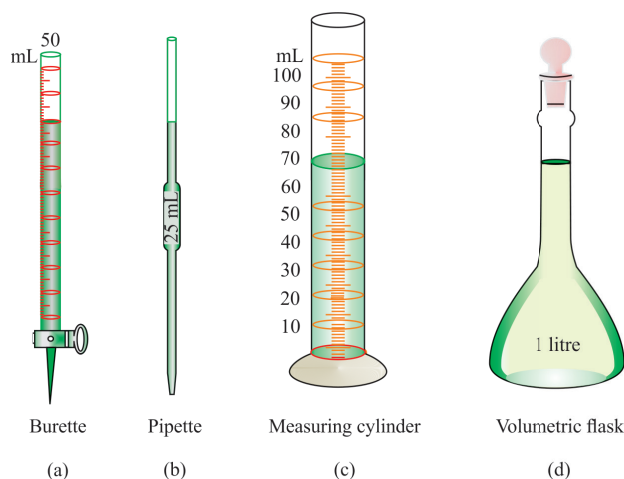


Figure 2. Some commonly used measuring apparatus.

The volume of a liquid can be measured by using graduated (measuring) cylinder, a burette, or a pipette while a volumetric flask is used to take measured volume of the liquid.

Density

Density is the amount of mass in a unit volume of matter. Its symbol is ρ

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \text{ or } \rho = m / v$$

Density can be measured in units of g cm^{-3} , kg m^{-3} or g/ml .

For example, 1.00 g of water occupies a volume of 1.00 cm^3 or 1 mL

$$\rho = m/v = 1.00\text{g}/1.00\text{cm}^3 = 1.00\text{g cm}^{-3}$$

Example

Gold is a precious metal that is chemically unreactive. It is used mainly in jewelry, dentistry, and electronic devices. A piece of gold ingot with a mass of 257 g has a volume of 13.3 cm^3 . Calculate the density of gold.

Solution

We are given the mass and volume and asked to calculate the density.

$$\begin{aligned} \rho &= m/v \\ &= 257\text{g}/13.3\text{cm}^3 = 19.3\text{g/cm}^3 \end{aligned}$$

Exercise

1. Ethanol is used in alcoholic beverages and has a density of 0.789 g/mL. What volume of ethanol (in litres) would have a mass of 500 g?
2. Calculate the density of a rectangular block of metal whose length is 8.335 cm, width is 1.02 cm, height is 0.982 cm and mass is 62.3538 g.
3. A piece of silver metal weighing 194.3 g is placed in a graduated cylinder containing 242.0 mL of water. The volume of water now reads 260.5 mL. Calculate the density of the metal.

Scientific notation and Significant figures

Scientific notation is the representation of a number in the form $A \times 10^n$, where A is a number with a single non-zero digit to the left of the decimal point and n is an integer, or whole number. In scientific notation, the measurement 900 cm precise to two significant figures is written 9.0×10^2 cm. If precise to three significant figures, it is written 9.00×10^2 cm. Scientific notation is also convenient for expressing very large or very small quantities. It is much easier (and simplifies calculations) to write the speed of light as 3.00×10^8 (rather than as 300,000,000) meters per second.

Suppose that we are given a certain number and asked to express it in scientific notation. Basically, this assignment calls for us to find n. We count the number of places that the decimal point must be moved to give the number A (which is between 1 and 10). If the decimal point has to be moved to the left, then n is a positive integer; if it has to be moved to the right, n is a negative integer. The following examples illustrate the use of scientific notation:

1. Express 568.762 in scientific notation:

$$568.762 = 5.68762 \times 10^2$$

Note that the decimal point is moved to the left by two places and $n = 2$.

2. Express 0.00000772 in scientific notation:

$$0.00000772 = 7.72 \times 10^{-6}$$

Here the decimal point is moved to the right by six places and $n = -6$

Keep in mind the following two points. First, $n = 0$ is used for numbers that are not expressed in scientific notation. For example, 74.6×10^0 ($n = 0$) is equivalent to 74.6. Second, the usual practice is to omit the superscript when $n = 1$. Thus, the scientific notation for 74.6 is 7.46×10 and not 7.46×10^1 .

Next, we consider how scientific notation is handled in arithmetic operations.

Addition and Subtraction

To add or subtract using scientific notation, we first write each quantity—say A^1 and A^2 —with the same exponent n . Then we add A^1 and A^2 ; the exponents remain the same. Consider the following examples:

$$\begin{aligned}(7.4 \times 10^3) + (2.1 \times 10^3) &= 9.5 \times 10^3 \\(4.31 \times 10^4) + (3.9 \times 10^3) &= (4.31 \times 10^4) + (0.39 \times 10^4) \\ &= 4.70 \times 10^4 \\(2.22 \times 10^{-2}) - (4.10 \times 10^{-3}) &= (2.22 \times 10^{-2}) - (0.41 \times 10^{-3}) \\ &= 1.81 \times 10^{-2}\end{aligned}$$

Multiplication and Division

To multiply numbers expressed in scientific notation, we multiply A^1 and A^2 in the usual way, but add the exponents together. To divide using scientific notation, we divide A^1 and A^2 as usual and subtract the exponents. The following examples show how these operations are performed:

$$\begin{aligned}(8.0 \times 10^4) \times (5.0 \times 10^2) &= (8.0 \times 5.0)(10^{4+2}) \\ &= 40 \times 10^6 = 4.0 \times 10^7 \\(4.0 \times 10^{-5}) \times (7.0 \times 10^3) &= (4.0 \times 7.0)(10^{-5+3}) \\ &= 28 \times 10^{-2} = 2.8 \times 10^{-1} \\(6.9 \times 10^7) \div (3.0 \times 10^{-5}) &= \frac{6.9 \times 10^7}{3.0 \times 10^{-5}} = \frac{6.9 \times 10^{7+5}}{3.0} \\ &= 2.3 \times 10^{12} \\(8.5 \times 10^4) \div (5.0 \times 10^9) &= \frac{8.5 \times 10^4}{5.0 \times 10^9} = \frac{8.5}{5.0} \times 10^{4-9} \\ &= 1.7 \times 10^{-5}\end{aligned}$$

Significant Figures

Except when all the numbers involved are integers (for example, in counting the number of students in a class), obtaining the exact value of the quantity under investigation is often impossible. For this reason, it is important to indicate the margin of error in a measurement by clearly indicating the number of *significant figures*, which are the meaningful digits in a measured or calculated quantity. When

significant figures are used, the last digit is understood to be uncertain. For example, we might measure the volume of a given amount of liquid using a graduated cylinder with a scale that gives an uncertainty of 1 mL in the measurement. If the volume is found to be 6 mL, then the actual volume is in the range of 5 mL to 7 mL. We represent the volume of the liquid as (6 ± 1) mL. In this case, there is only one significant figure (the digit 6) that is uncertain by either plus or minus 1 mL. For greater accuracy, we might use a graduated cylinder that has finer divisions, so that the volume we measure is now uncertain by only 0.1 mL. If the volume of the liquid is now found to be 6.0 mL, we may express the quantity as (6.0 ± 0.1) mL, and the actual value is somewhere between 5.9 mL and 6.1 mL. We can further improve the measuring device and obtain more significant figures, but in every case, the last digit is always uncertain; the amount of this uncertainty depends on the particular measuring device we use.

The following guidelines apply to determining the number of significant figures in a measured quantity. It has to be decided whether zeros are significant in three different situations.

1. If the zeros precede the first non-zero digit, they are not significant. Such zeros merely locate the decimal point. i.e., they define the magnitude of measurement. For example, 0.004 m has one significant figure, and 0.00016 m has two significant figures.
2. If the zeros are between non-zero digits, they are significant. For example, 204408 kg has six significant figures while 0.05504 has four significant figures.
3. If the zeros follow non-zero digits, there is ambiguity if no decimal point is given. For example, if a volume is given as 200 cm³, there is no way of expressing if the final two zeros are significant. But if the volume is given as 200 cm³, zeros after a non-zero digit preceded by a decimal point make all figures significant. Thus, 200 cm³ has three significant figures. If it is given as 200.0 cm³, it has four significant figures.
4. Non-zero digits are always significant.

Example

Determine the number of significant figures in the following measurements:

1. 394 cm
2. 0.052 kg
3. 5.03 g

4. 2.720×10^{22} atoms
5. 0.714 m
6. 3000 mL.

Solution

1. Three, because each digit is a non-zero digit.
2. Two, because zeros to the left of the first non zero digit do not count as significant figure.
3. Three, because zeros between non-zero digits are significant.
4. Four, because the number is greater than one, all the zeros written to the right of the decimal point count as significant figures.
5. Three
6. This is an ambiguous case. The number of significant figures may be four (3.000×10^3), three (3.00×10^3), two (3.0×10^3), or one (3×10^3). This example illustrates why scientific notation must be used to show the proper number of significant figures.

Exercise

Determine the number of significant figures in each of the following measurements:

- (a) 35 mL,
- (b) 2008 g,
- (c) 0.0580 m³
- (d) 7.2×10^4 molecules
- (e) 830 kg

Rules for Significant Figures in Mathematical Operations:

1. For addition or subtraction, the result has the same number of decimal places as the least precise measurement used in the calculation. Consider these examples:

$$\begin{array}{r} 89.332 \\ +11.1 \\ \hline 90.432 \end{array} \leftarrow \text{one digit after the decimal point}$$

$$90.432 \leftarrow \text{round off to 90.4}$$

$$\begin{array}{r} 2.097 \\ -0.12 \\ \hline 1.977 \end{array} \leftarrow \text{two digits after the decimal point}$$

$$1.977 \leftarrow \text{round off to 1.98}$$

The rounding-off procedure is as follows. To round off a number at a certain point we simply drop the digits that follow if the first of them is less than 5. Thus, 8.724 rounds off to 8.72 if we want only two digits after the decimal point. If the first digit following the point of rounding off is equal to or greater than 5, we add 1 to the preceding digit. Thus, 8.727 rounds off to 8.73, and 0.425 rounds off to 0.43.

2. For multiplication or division, the number of significant figures in the result is the same as the number in the least precise measurement used in the calculation. The following examples illustrate this rule:

$$2.8 \times 4.5039 = 12.61092 \longleftarrow \text{round off to 13}$$

$$\frac{6.85}{112.04} = 0.0611388789 \longleftarrow \text{round off to 0.0611}$$

Example

Carry out the following arithmetic operations to the correct number of significant figures:

- (a) $12,343.2 \text{ g} + 0.1893 \text{ g}$
- (b) $55.67 \text{ L} - 2.386 \text{ L}$
- (c) $7.52 \text{ m} \times 6.923$
- (d) $0.0239 \text{ kg} / 46.5 \text{ mL}$
- (e) $5.21 \times 10^3 \text{ cm} + 2.92 \times 10^2 \text{ cm}$.

Solution

In addition and subtraction, the number of decimal places in the answer is determined by the number having the lowest number of decimal places. In multiplication and division, the

significant number of the answer is determined by the number having the smallest number of significant figures.

$$\begin{array}{r} 12,343.2 \text{ g} \\ + 0.1893 \text{ g} \\ \hline 12,343.3893 \text{ g} \end{array} \longleftarrow \text{round off to } 12,343.4 \text{ g}$$

$$\begin{array}{r} 55.67 \text{ L} \\ - 2.386 \text{ L} \\ \hline 53.284 \text{ L} \end{array} \longleftarrow \text{round off to } 53.28 \text{ L}$$

$$(c) \quad 7.52 \text{ m} \times 6.9232 = 52.06246 \text{ m} \longleftarrow \text{round off to } 52.1 \text{ m}$$

$$(d) \quad \frac{0.0239 \text{ kg}}{46.5 \text{ ml}} = 0.0005139784946 \text{ kg/ml} \longleftarrow \text{round off to } 0.000514 \text{ kg/ml} \text{ or } 5.14 \times 10^{-4} \text{ kg/ml}$$

- (e) First we change $2.92 \times 10^2 \text{ cm}$ to $0.292 \times 10^3 \text{ cm}$ and then carry out the addition $(5.21 \text{ cm} + 0.292 \text{ cm}) \times 10^3$. Following the procedure in (a), we find the answer is $5.50 \times 10^3 \text{ cm}$.

Exercise

- Perform the indicated operations and give answers with the proper number of digits.
 - $451 \text{ g} - 15.46 \text{ g} - 20.3 \text{ g}$
 - $15.436 \text{ L} + 5.3 \text{ L} - 6.24 \text{ L} - 8.177 \text{ L}$
 - $48.2 \text{ m} + 3.82 \text{ m} + 48.4394 \text{ m}$
 - $148 \text{ g} + 2.39 \text{ g} + 0.0124 \text{ g}$
 - $37 \text{ m} + 2.340 \text{ m} + 0.52 \text{ m}$
 - $62.89 \text{ m} + 4.7 \text{ m}$
- Perform the following calculations and round off the answers to the correct number of significant figures (units of measurement have been omitted).
 - $\frac{2.568 \times 5.8}{4.186}$
 - $5.41 - 0.398$
 - $3.38 - 3.01$
 - $4.18 - 58.16 \times (3.38 - 3.01)$
- Carry out these operations as if they were calculations of experimental results, and express each answer in the correct units and with the correct number of significant figures:
 - $7.310 \text{ km} + 5.70 \text{ km}$
 - $(3.26 \times 10^{-3} \text{ mg}) - (7.88 \times 10^{-5} \text{ mg})$
 - $(4.02 \times 10^6 \text{ dm}) + (7.74 \times 10^7 \text{ dm})$
 - $(7.8 \text{ m} - 0.34 \text{ m}) / (1.15 \text{ s} + 0.82 \text{ s})$

Precision and accuracy in measurement and Standard deviation

In discussing measurements and significant figures it is useful to distinguish between accuracy and precision. **Accuracy** tells us how close a measurement is to the true value of the quantity that was measured. To a scientist there is a distinction between accuracy and precision. **Precision** refers to how closely two or more measurements of the same quantity agree with one another.

ACTIVITY 5

Form a group, perform the following activities and discuss each of the following questions. After the discussion, share your ideas with the rest of the class.

Suppose that three students are asked to determine the mass of a piece of copper wire. The results of two successive weighings by each student are

	Student A	Student B	Student C
Reading 1	1.964 g	1.972 g	2.000 g
Reading 2	1.978 g	1.968 g	2.002 g
Average value	1.971 g	1.970 g	2.001 g

If the true mass of the wire is 2.000 g.

1. Calculate the average mass from each set of data, and determine which set is the most accurate
2. Which set of data is the most precise? Is this set also the most accurate?

Which set of data is the least accurate? Is this set also the least precise?

If you are precise, that doesn't necessarily mean you are accurate. However, if you are consistently accurate, you are also precise.



Figure 3.

Exercise

Students A and B determined the mass of an object having an actual mass of 10.24g using two different balances. Each student made three trials as tabulated below,

Table 3

Student	Trial 1	Trial 2	Trial 3
A	10.26	10.24	10.20
B	10.20	10.18	10.22

1. Calculate the average mass from each set of data, and determine which set is the most accurate
2. Which set of data is the most precise? Is this set also the most accurate?
3. Which set of data is the least accurate? Is this set also the least precise?

Standard deviation

Precision and accuracy are linked with two common types of errors called Random and Systematic errors.

Random error makes a measurement less precise but not in any particular direction. In other words, the actual value may be either greater or smaller than the value one records. Random errors arise mostly from inadequacies or limitations in the instrument. On the other hand, this may be a result of how precisely someone can read a meter or a scale.

In general, the more significant figures there are in a measurement quantity, the greater the precision of the measurement. The precision of measurement can be evaluated in terms of standard deviation (average deviation AD) and relative average deviation (RAD) from the mean.

$$AD = \frac{\sum |X_i - \bar{X}|}{n} \text{ and } RAD = \frac{AD}{\bar{X}} \times 100$$

Where \bar{X} is the mean value of a set of measurements, X_i is the individual value, and $\sum |X_i - \bar{X}|$ is the sum of the magnitude of the deviation from the mean, and n is the number of measurements. To illustrate this concept, consider the density of water at 4°C and 1.0 atmosphere. The accepted value of the density of water is 1.0g/mL. Let us assume that two students, student A and B determined the density of water at 4°C and 1.0 atmosphere in six separate experiments and obtained the following values in g/mL.

Table 4

Student	Trial 1	Trial 2	Trial 3	Trial 4	Trial 5	Trial 6	Average
A	0.91	0.92	0.90	0.88	0.89	0.91	0.90
B	1.10	0.99	1.01	0.97	0.91	1.03	1.00

- Which result is precise? Is it also accurate?
- Which result is accurate? Is it also precise?
- Which result has lower values for AD and RAD
- Which result has less uncertainty?

Solution

- The result of student A are closer to each other and hence are precise. But the average value of 0.90g/ml is less than the true value of 1.0g/ml, indicating inaccurate results.
- The result of student B is accurate, because the average value of 1.00g/ml is in good agreement with the true value. But it is less precise because there is less agreement within the set
- To determine AD and RAD we use the equation

$$AD = \frac{\sum |X_i - \bar{X}|}{n} \quad \text{and} \quad RAD = \frac{AD \times 100}{\bar{X}}$$

$$\text{student A : } \bar{X} = 0.90$$

$$\sum |X_i - \bar{X}| = \sum |0.91 - 0.90| + |0.92 - 0.90| + |0.90 - 0.90| + |0.88 - 0.90| +$$

$$|0.89 - 0.90| + |0.91 - 0.90| = 0.0$$

$$AD = \frac{0.7}{6} = 0.1117$$

$$RAD = \frac{0.1117 \times 100}{0.90} = 0.1241$$

$$\text{student B : } \bar{X} = 1.00$$

$$\sum |X_i - \bar{X}| = \sum |1.10 - 1.00| + |0.99 - 1.00| + |1.01 - 1.00| +$$

$$|1.01 - 1.00| + |0.97 - 1.00| + |0.91 - 1.00| + |1.03 - 1.00| = 0.27 \quad \text{and} \quad \sum \frac{|X_i - \bar{X}|}{n} = \frac{0.27}{6}$$

$$RAD = \left(\frac{0.045}{100} \right) \times 100 = 4.5$$

The uncertainty of student A is less than student B, as indicated by the lower values for AD and RAD. This implies that student A has made more precise measurements than student B.

Systematic errors produce values that are either entirely higher or smaller than the actual value. It always affects a result in a particular direction, and skews the accuracy of the experiment in that direction. Systematic errors arise from flaws or defects in the instrument or from errors in the manner that the measurement was taken. For example, when you are taking the initial reading of a burette that is placed well above head height, you might decide to read the top rather than the bottom of the meniscus. Systematic errors can lead to inconsistent results.

The total error, which is the sum of both random and systematic error involved in a measurement may be expressed in terms of accuracy. Mathematically error can be expressed as:

$$\text{Error} = | \text{True value} - 1.00 |$$

The error is stated as positive quantity whether the measured value is higher or lower than the true value. It follows that the percentage error of a measurement is:

$$\% \text{Error} = \frac{\text{Error}}{\text{True value}} \cdot 100$$

1.5 CHEMICAL SYMBOLS, FORMULA AND NAMING COMPOUNDS

ACTIVITY 6

Form a group and do the following tasks. Learning the language of chemistry is like learning English language. Use the following terms to answer questions number 1 and 2 below:

Table 5

Language of chemistry	English language
i) Symbol	a) Sentence
ii) Formula	b) Alphabet
iii) Equation	c) Word

- Match the terms from the language of chemistry with the terms from English language. Support your matching by giving reasons.
- From which terms we must start in learning the language of chemistry and English language. After the discussion, share your ideas with the class.

Origins of chemical symbols

ACTIVITY 7

Discuss in groups and share your ideas with the rest of the class.

1. Why are students given roll numbers by their homeroom teachers?
2. What is the relationship between students' name and their roll numbers?
3. How could you relate roll numbers of students with symbols of elements?

The names of the elements are derived from various sources. Many are derived from Greek, Latin or German words that generally describe some properties of the elements. For example, the word '*iodine*' is taken from the Greek word '*iodes*', meaning violet-like. Iodine, indeed, is violet in vapor state. Some elements are named after the place of their discovery. For example, *americium* is named after America, and *polonium* is named after Poland. Some others are named in commemoration of famous scientists, such as *curium* for the Polish scientist *Madam Merie Curie*, *mendelevium* for the Russian scientist *Mendeleev*, and *einsteinium* for the scientist *Albert Einstein*. A few elements are named after the names of planets. For example, *uranium* is named after the planet *Uranus*, *neptunium* is named after the planet *Neptune*, and plutonium is named after Pluto.

Many signs and symbols were used to represent substances under investigation in the past. Examples of symbols used in the past are shown in Table 5.

Table 6 Dalton's symbol of the elements

Element	Dalton's Symbol	Element	Dalton's Symbol
Hydrogen		Iron	
Carbon		Copper	
Oxygen		Silver	
Phosphorus		Gold	
Sulphur		Lead	
Platinum		Mercury	

Chemists use symbols as abbreviation of the names of elements. A symbol is a shorthand notation of the chemical names of the element. Every element has its own symbol. The symbol is usually the first letter of the element's name or the first letter plus another letter.

In writing symbols, if only one letter is used, it is capitalized. For example: look at the following table:

Table 7 Symbols of some elements with the first letter

Name of the element	Symbol
Hydrogen	H
Oxygen	O
Carbon	C
Sulphur	S
Phosphorus	P
Fluorine	F
Boron	B

To avoid giving identical symbol to two or more different elements a second letter or a letter outstanding in the name is added to the first letter. If two letters are used, the first letter is capitalized and the second one is written in small letter. See Table 8.

Table 8 Symbols of the elements with two letters

Name of the element	Symbol
Calcium	Ca
Chlorine	Cl
Cobalt	Co
Chromium	Cr
Silicon	Si
Magnesium	Mg
Manganese	Mn
Strontium	Sr
Nickel	Ni
Platinum	Pt
Aluminium	Al
Barium	Ba
Bromine	Br
Helium	He
Lithium	Li
Zinc	Zn

The symbols for the above elements are derived from their English names. The symbols of some elements are derived from their Latin names. The following table shows the symbols of some elements derived from their Latin names.

Table 9 Symbols of the elements derived from the Latin names

Element	Latin name	Symbol
Silver	Argentum	Ag
Gold	Aurum	Au
Iron	Ferrum	Fe
Lead	Plumbum	Pb
Mercury	Hydragyrum	Hg
Potassium	Kalium	K
Silver	Argntum	Ag
Sodium	Natrium	Na
Tin	Stannum	Sn

Chemical Formulas

Do you know the difference between CO and CO₂ in Chemistry?

The formula of a substance is the symbolic representation of its composition. Formulas can be classified as formulas of elements and formulas of compounds.

The formula of an element consists of one kind of symbol whereas the formula of a compound contains the symbols of two or more different elements.

The formulas of elements

Only a few elements such as the gases helium, neon and argon consist of single uncombined atoms that move about independent of one another. Other elements, such as the gaseous hydrogen, oxygen, nitrogen etc, consist of pairs of atoms, each moving as a single unit. The element phosphorus consists of units of four atoms. Sulfur consists of eight atoms. Such units are called **molecules**.

A molecule of an element is an atom or a group of atoms that exists freely in nature.

The elements helium, neon, argon, krypton, xenon and radon are known as noble gases. Because they exist uncombined as single atoms, they are monoatomic molecules. Their formulas are the same as their symbols. The hydrogen molecule is a diatomic molecule because it contains only two atoms. It is represented by the

formula H_2 . Other elements that normally exist as diatomic molecules are nitrogen, oxygen, fluorine, chlorine, bromine and iodine. The following table shows the symbols and formulas of diatomic elements.

Table 10 Chemical symbols and formulas of diatomic elements

Name	Chemical symbol	chemical formula
Hydrogen	H	H_2
Nitrogen	N	N_2
Oxygen	O	O_2
Fluorine	F	F_2
Chlorine	Cl	Cl_2
Bromine	Br	Br_2
Iodine	I	I_2

Molecules containing more than two atoms are called polyatomic molecules. Ozone (O_3), Phosphorus (P_4) and Sulfur (S_8) are examples of polyatomic molecules.

Writing formula of compounds

It is easy to write the formula of a compound if we know the combining power of the elements or polyatomic ions involved. The combining power of an element is its relative capacity to combine with other elements. This combining power of an element is called valence number. In compounds containing hydrogen, whose valence number is 1, the valence number of the element is the number of hydrogen atoms that combine with one atom of the element. For example, in HCl one atom of chlorine combines with one atom of hydrogen, the valence number of chlorine is 1. Since one atom of oxygen combines with two atoms of hydrogen to form water, H_2O , the valence number of oxygen is 2. Similarly, one atom of nitrogen combines with three atoms of hydrogen to form ammonia, NH_3 . Therefore, the valence number of nitrogen is 3.

Table 11 Valence numbers of some common elements

Element	Valence number 1		Valence number 2		Valence number 3			
	Name	symbol	Name	Symbol	Name	Symbol		
Metals	Sodium	Na	Calcium	Ca	Aluminum	Al		
	Potassium	K	Magnesium	Mg				
	Silver	Ag	Zinc	Zn			Iron (III)	Fe
	Copper (II)	Cu	Iron (II)	Fe				
Non metals	Hydrogen	H	Oxygen	O	Nitrogen	N		
	Chlorine	Cl						
	Bromine	Br						

The valence numbers of most common elements are either 1, 2, or 3. Some elements have more than one valence number, which is different combining powers under different conditions. Common examples of these elements are iron, copper, lead and tin. One atom of iron can combine with either 2 or 3 atoms of chlorine. If the valence numbers of iron is 2 it forms FeCl_2 and is written as iron (II) chloride. If the valence number of iron is 3, it forms FeCl_3 and is written as iron (III) chloride. Copper and oxygen form Cu_2O and CuO , in which copper has valence numbers 1 and 2 respectively. Valence numbers of some common elements are shown in the Table 10.

Formulas of Binary Compounds

Binary compounds contain atoms of two different elements. Sodium chloride (NaCl), Hydrogen bromide (HBr) and Copper (II) oxide (CuO) are some examples of binary compounds. The formulas of binary compounds in which the metallic and non-metallic elements have the same valence numbers are simply represented by the symbols of the elements. This is because the combining powers of the elements are balanced. For example, The formula of potassium chloride is KCl because the valence numbers of both K and Cl is 1. The formula of calcium oxide is CaO because the valence numbers of both Ca and O is 2.

However, the formulas of binary compounds in which the constituent elements have different valence numbers are written in a way that balances the combining powers

of the elements. For example, let us see how to write the formula of magnesium fluoride. Since Mg has a combining power of 2 and F has a combining power of 1, we need two fluorine atoms to go with one magnesium atom. Hence the formula is MgF_2 .

The other simpler method of writing formulas of binary compounds is shown in the following steps:

Step 1: Write down the symbols for the elements in the compound.

Let us consider some examples:

- | | | | |
|-------|------------------|----|----|
| (i) | Sodium Oxide | Na | O |
| (ii) | Calcium Chloride | Ca | Cl |
| (iii) | Aluminum Oxide | Al | O |

Step 2: Write the valence numbers above the symbols.

- | | |
|-------|--------------------------|
| (i) | Na^1O^2 |
| (ii) | Ca^2Cl^1 |
| (iii) | Al^3O^2 |

Step 3: Now criss cross the valence numbers and put the numbers below the symbols.

- | | | | |
|-------|--|---------------|---|
| (i) | $\begin{array}{cc} 1 & 2 \\ \swarrow & \searrow \\ \text{Na} & \text{O} \end{array} \Rightarrow \text{Na}_2\text{O}$ | \Rightarrow | Na_2O . Thus, the formula of sodium oxide is Na_2O . |
| (ii) | $\begin{array}{cc} 2 & 1 \\ \swarrow & \searrow \\ \text{Ca} & \text{Cl} \end{array}$ | \Rightarrow | CaCl_2 . Thus, the formula of calcium chloride is CaCl_2 . |
| (iii) | $\begin{array}{cc} 3 & 2 \\ \swarrow & \searrow \\ \text{Al} & \text{O} \end{array}$ | \Rightarrow | Al_2O_3 . Thus, the formula of Aluminium oxide is Al_2O_3 . |

Example

Write the formulas of the following binary compounds

- Barium nitride
- Potassium iodide
- Magnesium oxide
- Aluminium chloride

Solution

- Barium has a valence number 2 and nitrogen has a valence number 3. To balance their combining powers, criss cross their valence numbers as follows:



Thus the formula of barium nitride is Ba_3N_2

- Since both potassium and iodine have valence numbers 1, the formula of potassium iodide is KI.
- Since both magnesium and oxygen have valence numbers 2, the formula of magnesium oxide is MgO.
- Aluminium has a valence number 3 and chlorine has a valence number 1. To balance their combining powers, criss cross their valence numbers as follows:



Thus, the formula of aluminum chloride is AlCl_3

Naming Compounds

ACTIVITY 8

Discuss in groups and share your ideas with the class.

- How many compounds do you know so far? Can you memorize their names easily?
- Up to how many names of compounds is it possible to memorize?
- Today the number of known compounds is over 13 million. Do you think there should be a systematic way of naming compounds?

The two constituent elements of binary compounds are usually metals and non-metals. For example, NaBr, CaO and AlCl_3 . There are also binary compounds which consist of non-metallic elements only. For example, H_2O , CO_2 and NH_3 .

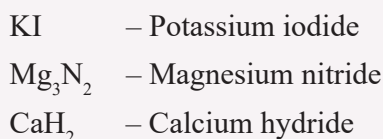
Rules in Naming Binary Compounds:

- The name of a binary compound is the combination of the names of the two constituent elements.
- For binary compounds that consist of metals and non-metals, the metal is named first followed by the non-metal.
- The suffix *-ide* replaces the last letters of the name of the non-metal. For example, chlorine is changed to chloride. See Table 12.

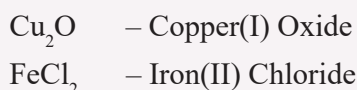
Table 12 Names of non-metallic elements in binary compounds

Non-metallic element	Name in binary compound
Fluorine	Fluoride
Bromine	Bromide
Iodine	Iodide

Oxygen	Oxide
Nitrogen	Nitride
Sulfur	Sulfide
Hydrogen	Hydride

Example

If the metal has variable valence numbers, the valence number of the metal used in the formula should be placed in parenthesis using capital Roman numerals after the name of the metal.

Example

If the binary compound consists of non-metallic elements only, we name the first element in the formula followed by the name of the second element with the suffix -ide.

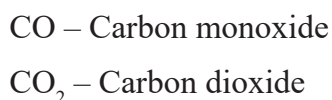
Example

It is quite common for a pair of elements to form several different compounds. In naming such compounds, we use the Greek prefixes to denote the number of atoms of each element. Note that a prefix is a group of letters added at the beginning of a word.

Table 13 Greek prefixes used in naming binary molecular compounds

Prefix	Meaning	Prefix	Meaning
Mono-	1	Penta-	5
Di-	2	Hexa-	6
Tri -	3	Hepta -	7
Tetra-	4	Octa -	8

Consider the following examples:



SO_2 – Sulfur dioxide

SO_3 – Sulfur trioxide

N_2O – Dinitrogen monoxide

NO_2 – Nitrogen dioxide

N_2O_4 – Di nitrogen tetroxide

P_2O_3 – Diphosphorus trioxide

P_2O_5 – Diphosphorus pentoxide

The following guidelines are helpful in naming compounds with prefixes:

- The prefix “mono” may be omitted for the first element. For example, PCl_3 is named phosphorus trichloride, not monophosphorus trichloride. Thus the absence of a prefix for the first element usually means there is only one atom of that element present in the molecule.
- For “oxides”, the ending “a” in the prefix is sometimes omitted. For example, N_2O_4 is named dinitrogen tetroxide rather than dinitrogen tetraoxide.

Exceptions to the use of Greek prefixes are binary molecular compounds containing hydrogen. Traditionally, many of these compounds are called either by their common names or by names that do not specifically indicate the number of H atoms:

H_2O – Water

NH_3 – Ammonia

PH_3 – Phosphine

H_2S – Hydrogen sulphide

Example

Name the following binary molecular compounds.

(a) SiCl_4

(b) N_2O_5

(c) NF_3

Solution

(a) Because there are four chlorine atoms, the name of compound is silicon tetrachloride

- (b) There are two nitrogen atoms and five oxygen atoms. Thus, the name of compound is dinitrogen pentoxide.
- (c) Since there are three fluorine atoms, the name of compound is nitrogen trifluoride.

Exercise

- The valence number of a metallic element X is 3. What is the formula of its oxide?
- Name the following binary compounds
 - Ca_3N_2
 - AlCl_3
 - MgS
 - SO_2
 - NO
 - PCl_3
- Write the formula of the binary compounds made up of
 - Magnesium and chlorine
 - Aluminium and bromine
 - Potassium and oxygen
 - Calcium and oxygen

Qualitative and Quantitative Significance of Symbols and Formulas

Symbols and formulas describe both qualitative and quantitative aspects of the substances.

Qualitatively a symbol represents the identity (kind) of the element. For example, the symbol N represents an atom of nitrogen. No other element can be represented by the symbol N.

Quantitatively a symbol represents the number of atoms of the elements.

Example

⇒ H stands for one atom of hydrogen.

⇒ Na stands for one atom of sodium.

Qualitatively a formula represents the kinds or types of elements involved in forming a compound.

Example

- ⇒ A water molecule, H_2O , contains the elements hydrogen and oxygen.
- ⇒ The compound sodium nitrate, NaNO_3 , contains the elements sodium, nitrogen and oxygen.

A formula stands for one molecule (formula unit) of an element or a compound.

- ⇒ CO_2 represents one molecule of carbon dioxide.
- ⇒ P_4 indicates one molecule of phosphorous.

A number written in front of a symbol or a formula is called a **coefficient**. It shows the number of atoms or molecules or formula units of the substances.

Example

- ⇒ 3Fe represents 3 atoms of iron
- ⇒ 2H means 2 atoms of hydrogen (Note that: 2H differs from H_2)
- ⇒ 5N_2 means 5 molecules of nitrogen

Note that the coefficient multiplies the whole atoms of the elements in the formula, not only the first element. For example, 2HNO_3 represents two molecules of nitric acid which contains 2 atoms of hydrogen, 2 atoms of nitrogen and 6 atoms of oxygen.

Subscripts are used in writing formulas. A subscript is a number written at the bottom right hand side of a symbol. In the formula of an element, the subscript qualitatively shows the element is in a molecular form. For example Cl_2 is a chlorine molecule, P_4 is a phosphorus molecule and S_8 is a sulphur molecule. Quantitatively a subscript in a formula stands for the number of atoms in the formula.

H_2 indicates a molecule of hydrogen which contains 2 atoms. The general notation is shown below using hydrogen molecule as an example.

Symbol \longrightarrow H_2 \longleftarrow subscript

Exercise

- Fill the blank spaces in the table given below:

Table 14

	Chemical symbol or formula	Qualitative meaning	Quantitative meaning
a	$2C_{12}$		
b	$4Cl$		
c	$3S_8$		
d	$5MgF_2$		

1.6 LABORATORY APPARATUS AND SAFETY MEASURES

What is a laboratory apparatus? The following are some of the apparatus commonly used in chemistry laboratories.

Beakers are useful as a reaction container or to hold liquid or solid samples. They are also used to catch liquids from titrations and filtrates from filtering operations.

Bunsen burners are sources of heat.

Burette are used for addition of a precise volume of liquid. The volume of liquid added can be determined to the nearest 0.01 mL with practice.

Clay triangles are placed on a ring attached to a ring stand as a support for a funnel, crucible, or evaporating dish.

Droppers are for addition of liquids drop by drop.

Glass funnels are for funneling liquids from one container to another or for filtering when equipped with filter paper.

Erlenmeyer flasks are useful to contain reactions or to hold liquid samples. They are also useful to catch filtrates.

Graduated cylinders are for measurement of an amount of liquid. The volume of liquid can be estimated to the nearest 0.1 mL with practice.

Hot plates can also be used as sources of heat when an open flame is not desirable.

Pipette are used to dispense small quantities of liquids.

Ring stand with rings are for holding pieces of glassware in place.

Test tubes are for holding small samples or for containing reactions.

Test tube holders are for holding test tubes when tubes should not be touched.

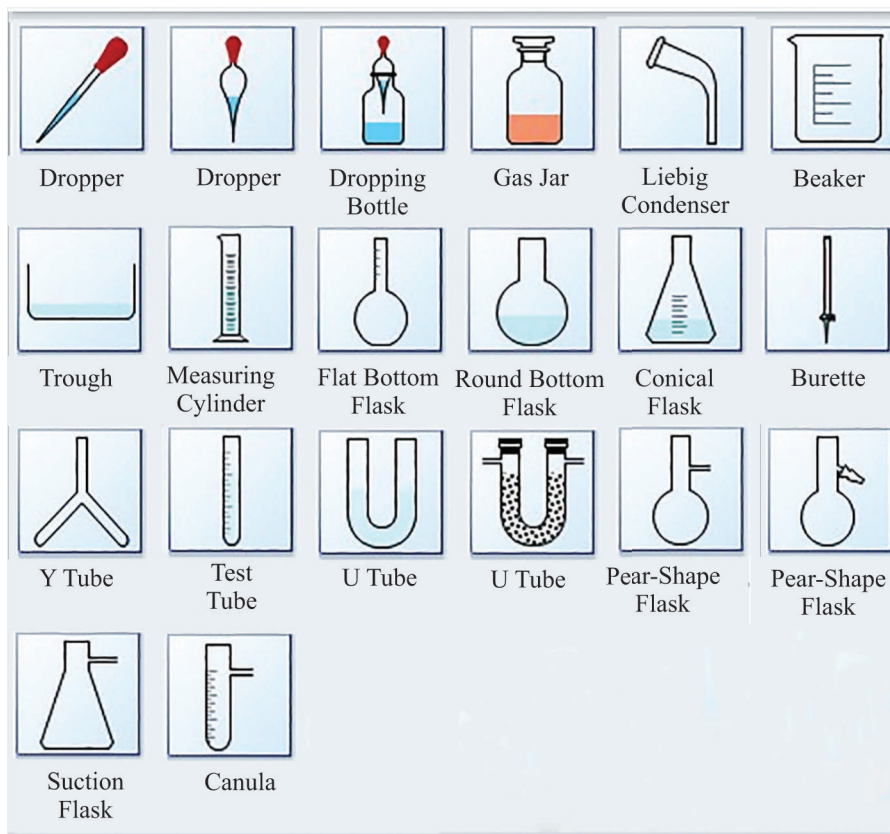


Figure 4. Some of the apparatuses found in chemistry laboratory

Tongs are similar in function to forceps but are useful for larger items.

Volumetric flasks are used to measure precise volumes of liquid or to make precise dilutions.

Wash bottles are used for dispensing small quantities of distilled water.

Watch glasses are for holding small samples or for covering beakers or evaporating dishes.

Wire gauze on a ring supports beakers to be heated by Bunsen burners.

Safety Rules inside the Laboratory

- Clean up your work area before leaving
- Wear safety glasses/goggles
- Wear gloves when using any hazardous or toxic agent
- Turn off gas when not in use
- Disposal – Students are responsible for the proper disposal of used material if any in appropriate containers
- Read labels carefully
- Tie back long hair
- Never eat, drink, or smoke while working in the laboratory
- Keep bags away from burners
- Do not engage in practical jokes or boisterous conduct in the laboratory.
- Never run in the laboratory
- The use of personal audio or video equipment is prohibited in the laboratory
- The performance of unauthorized experiments is strictly forbidden
- Never work in the laboratory without the supervision of an instructor
- Perform the experiments or work precisely as directed by your instructor
- Immediately report any spills, accidents, or injuries to your instructor
- Never leave experiments while in progress
- Make sure no flammable solvents are in the surrounding area when lighting a flame
- Do not leave lit Bunsen burners unattended
- Turn off all heating apparatus, gas valves, and water faucets when not in use

KEY TERMS

- Accuracy
- Atoms
- Analytical chemistry
- Atomic theory
- Basic unit

- Biochemistry
- Chemical change
- Derived unit
- Formula
- Inorganic chemistry
- Natural science
- Organic chemistry
- Physical change
- Precision
- Physical chemistry
- Scientific method
- Scientific notation
- Significant figure
- Symbol

SUMMARY

- The science of chemistry deals with the composition, physical properties, and chemical properties of matter.
- The SI system has seven base units, six of which are used in chemistry.
- Some measurements are expressed directly in terms of base units as well as multiples or submultiples of a base unit. For example, you might express a length in meters as well as in kilometers or millimeters.
- A measured quantity must be expressed with the proper number of significant figures to indicate its precision.
- In reporting calculated quantities special attention must be paid to the concept of significant figures.
- Calculations can be done by the unit-conversion method.
- Techniques of estimating answers are also useful in problem solving.
- The scientific method involves making observations, doing experiments forming hypothesis, gathering data, and formulating laws and theories.
- A chemical symbol is a shorthand notation for the chemical name of an element.
- The first letter of a symbol is always capitalized, but the next letter is not.
- A chemical formula is the symbolic representation of a substance giving the ratios of different kinds of atoms in it.
- The formula of an element consists of one kind of symbol.
- The formula of a compound consists of two or more kinds of symbols.

- Molecules containing only two atoms are called diatomic molecules.
- Molecules containing more than two atoms are called polyatomic molecules.
- The combining power of an element or polyatomic ion is called valence number.
- Binary compounds contain atoms of two different elements only.
- A prefix is a letter or group of letters added to the beginning of a word.
- The formulas of compounds can be written if the symbols and valence numbers of the constituent elements or polyatomic ions are known.
- Qualitatively a symbol represents the identity of the element.
- Quantitatively a symbol represents the number of atoms of the elements.
- Qualitatively a formula represents the kind of elements involved in making a compound.
- A formula stands for one molecule or formula unit of a molecule or a compound.
- A number in front of a symbol or a formula is called a coefficient.
- A subscript is written on the bottom right hand side of a symbol.

Exercises

Part I. Multiple Choice Questions

1. Which of the following numbers has five significant figures?
 - (a) 61,530
 - (b) 0.6153
 - (c) 0.006154
 - (d) 615.40
2. What is the mass of 30.0 mL of a liquid that has a density of 1.60 g/mL?
 - (a) 18.8 g
 - (b) 48.0 g
 - (c) 31.6 g
 - (d) 53.3 g
3. A blood sample analysis showed the presence of total 0.00361 g lead. Express it in micrograms?
 - (a) 0.361 μg
 - (b) $3.61 \times 10^3 \mu\text{g}$
 - (c) $3.61 \times 10^6 \mu\text{g}$
 - (d) $3.61 \times 10^4 \mu\text{g}$

4. The product of 12.76 cm and 0.0030 cm has _____ significant figures.
 - (a) 2
 - (b) 3
 - (c) 4
 - (d) 5
5. A measurement gave a mass of 0.45 kg. This is the same as:
 - (a) 4.5 g
 - (b) 45 g
 - (c) 450 g
 - (d) 4,500 g
6. Which of the following statements is not correct?
 - (a) Density has no units.
 - (b) Every measurement has a unit tied to it.
 - (c) Physical quantities are properties that can be measured.
 - (d) The kelvin degree is larger than the celsius degree.
7. The number 52.415 ± 0.0001 has _____ significant figures.
 - (a) 4
 - (b) 5
 - (c) 6
 - (d) 7
8. The chemical symbol for phosphorus is
 - (a) Po
 - (b) P
 - (c) Pt
 - (d) K
9. Which of the following is the correct name of BaBr_2 ?
 - (a) Boron bromide
 - (b) Barium bromide
 - (c) Barium dibromide
 - (d) Barium (I) bromide
10. Aluminium has a valence number of 3 and sulphur has a valence number of 2. What is the chemical formula for aluminium sulphide?
 - (a) Al_2S
 - (b) AlS_3
 - (c) Al_3S_2
 - (d) Al_2S_3

11. The name of NO_2 is
 - (a) Nitrogen monoxide
 - (b) Nitrogen oxide
 - (c) Nitrogen dioxide
 - (d) Mononitrogen oxide
12. Quantitatively the formula 3N_2 represents
 - (a) 3 atoms of nitrogen
 - (b) 3 molecules of nitrogen
 - (c) 2 molecules of nitrogen
 - (d) 6 molecules of nitrogen
13. The Latin name of sodium is a
 - (a) Argentum
 - (b) Kalium
 - (c) Natrium
 - (d) Cuprum
14. The symbol 'K' stands for
 - (a) Calcium
 - (b) Potassium
 - (c) Cobalt
 - (d) Cadmium
15. The written under symbol to the right is
 - (a) coefficient
 - (b) superscript
 - (c) subscript
 - (d) valence

Part II. Answer the following questions.

16. Convert -40°C to $^\circ\text{F}$.
17. Round off 45.68 meters to one decimal place.
18. How many seconds are there in 8 minutes?
 - (a) Change each of the following measurements to one in which the unit has an appropriate SI prefix
 - (b) $3.76 \times 10^3 \text{ m}$
 - (c) $6.34 \times 10^{-6} \text{ g}$
 - (d) $1.09 \times 10^{-9} \text{ g}$

19. How many significant figures are there in each of the following measured quantities?
- (a) 4.200×10^5 s
 - (b) 0.1050°C
 - (c) 6.02×10^{23}
 - (d) 0.049300 g
 - (e) 0.00075 m
 - (f) 8008 m
20. Express each of the following measured quantities in exponential notation. Assume all the zeros in parts c and d, are significant.
- (a) 0.00090 cm
 - (b) 20.00 s
 - (c) $9,000$ s
 - (d) $2,800$ m
21. Perform the indicated operations and give answers with the proper number of significant figures.
- (a) $48.2 \text{ m} + 3.82 \text{ m} + 48.4394 \text{ m}$
 - (b) $451 \text{ g} - 15.46 \text{ g}$
 - (c) $15.44 \text{ mL} - 9.1 \text{ mL} + 105 \text{ mL}$
 - (d) $73.0 \times 1.340 \times 0.41$
 - (e)
$$\frac{22.61 \times 0.0587}{135 \times 28}$$
22. A 25.0 mL sample of liquid bromine weighs 78.0 g. Calculate the density of the bromine
23. A block of lead, with dimensions $2.0 \text{ dm} \times 8.0 \text{ cm} \times 35 \text{ mm}$, has a mass of 6.356 kg. Calculate the density of lead in g cm^{-3} .
24. Write chemical formulas for the following compounds.
- (a) Sodium phosphate
 - (b) Magnesium iodide
 - (c) Sulphur trioxide
 - (d) Copper (II) oxide
25. Name the following compounds
- (a) MgH_2
 - (b) Al_2O_3
 - (c) BaSO_4
 - (d) NO



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CHAPTER

2

MATTER AND ITS PROPERTIES

Chapter Contents

- 2.1 States of matter and their changes
- 2.2 Properties and Changes of Matter
- 2.3 Standard separation techniques of mixture
- 2.4 Classifications of Mixtures
 - Key Terms
 - Summary
 - Exercises



Chapter outcomes

Upon completion of this chapter, students will be able to:

- define matter;
- discuss the states of matter;
- define physical properties;
- list some physical properties matter;
- make use of separation techniques to purify impure substances;
- identify substances based on their physical properties;
- conduct experiments to identify properties of substances and make group report;
- compare and contrast homogeneous and heterogeneous mixtures.

ACTIVITY 1

Discuss in groups and present the points of your discussion to the class.

Collect some materials from your environment and bring them to the classroom.

1. Identify and name the materials you collected.
2. What criteria do you use in identifying the materials?
3. Are these materials have same properties?

Everything in our environment, whether it is solid, liquid or gas, natural or manmade is a form of matter. Matter is anything that occupies space and has mass. Water, air, sugar, milk, soil, animals and plants are some examples of matter. A form of matter possessing constant physical properties under specific conditions is called a substance. For example, water is a substance because it has constant properties like color, odor, taste, etc.

We describe iron as a silvery-colored metal that melts at 1535°C (2795°F). Once we have collected enough descriptive information about many different kinds of matter, patterns emerge that suggest ways of classifying it. There are two principal ways of classifying matter: by its physical state as a solid, liquid, or gas and by its chemical constitution as an element, compound, or mixture.

2.1 STATES OF MATTER AND THEIR CHANGES

Commonly, a given kind of matter exists in different physical forms under different conditions. Water, for example, exists as ice (solid water), as liquid water, and as steam. See figure 1.

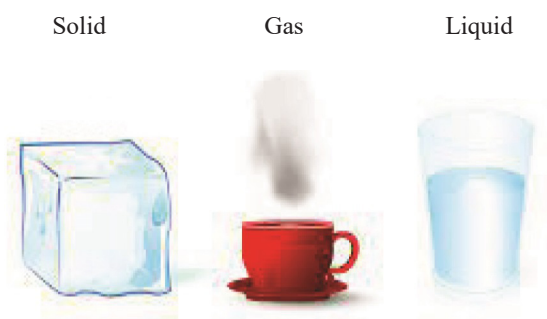


Figure 1. Solid water (ice), liquid water and gaseous water (steam)

The main identifying characteristic of solids is their rigidity: they tend to maintain their shapes when subjected to outside forces. Liquids and gases, however, are fluids; that is, they flow easily and change their shapes in response to slight outside forces.

What distinguishes a gas from a liquid is the characteristic of compressibility (and its opposite, expansibility). A gas is easily compressible, whereas a liquid is not. You can put more and more air into a tire, which increases only slightly in volume. In fact, a given quantity of gas can fill a container of almost any size. A small quantity would expand to fill the container; a larger quantity could be compressed to fill the same space. By contrast, if you were to try to force more liquid water into a closed glass bottle that was already full of water, it would burst.

These two characteristics, rigidity (or fluidity) and compressibility (or expansibility), can be used to frame definitions of the three common states of matter:

Solid is the form of matter characterized by rigidity; a solid is relatively incompressible and has fixed shape and volume.

Liquid is the form of matter that is a relatively incompressible fluid; a liquid has a fixed volume but no fixed shape.

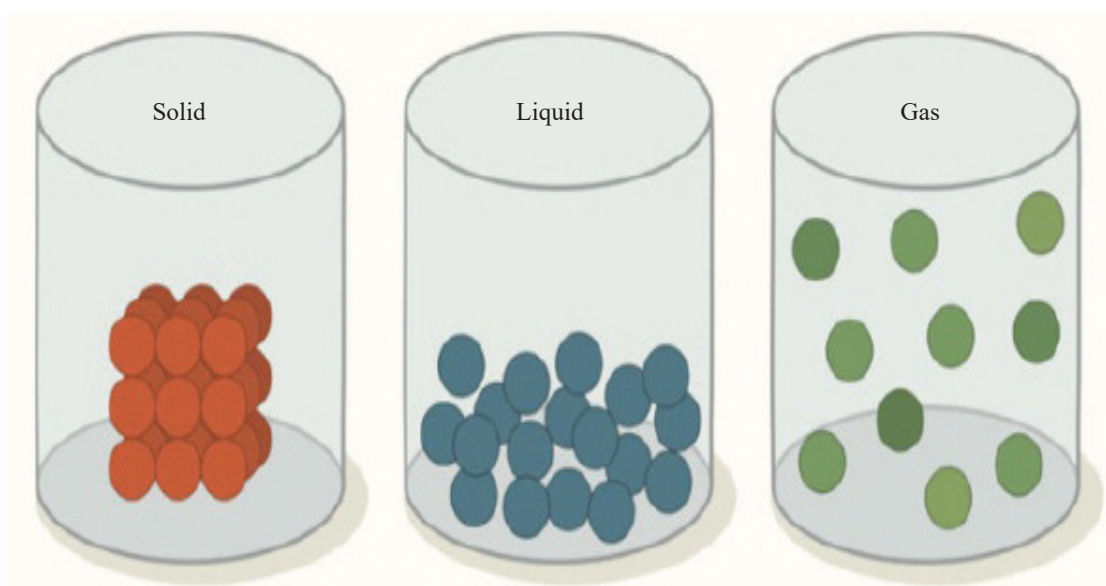


Figure 2. The three physical states of matter

Gas is the form of matter that is an easily compressible fluid; a given quantity of gas will fit into a container of almost any size and shape. The term vapor is often used to refer to the gaseous state of any kind of matter that normally exists as a liquid or a solid. These three forms of matter—solid, liquid, gas—comprise the common states of matter. Figure 2 illustrates the three physical states of matter.

The characteristics of the three states of matter is summarized in the following Table

Table 1 Summary of the characteristics of the three states of matter

Solid	Liquid	Gas
Not compressible (particles very close to each other)	Slightly compressible (particles slightly far from each other compared to solid)	Highly compressible (particles far from each other)
High density	Relatively low density compared to solid	Low density
Definite volume	Definite volume	Fill container completely
Retain its own shape	Assume the shape of its container	Assume the shape of its container
Motion limited to vibrational movement	Slow diffusion – particles can slip past each other	Rapid diffusion
Low expansion on heating	Slightly expanded on heating compared to solid	High expansion on heating

Changing state of matter

You would have observed changing states of matter when ice cubes melt from solid into liquid water or when water boils into vapor, but have you wondered why substances change form? Changing states of matter occur when matter loses or absorbs energy. When a substance absorbs energy the atoms and molecules move more rapidly and this increased kinetic energy pushes particles far enough, which they change form. This energy is usually heat or thermal energy.

A change of state is a physical change in a matter. They are reversible changes and do not involve any changes in the chemical makeup of the matter. Common changes of the state include melting, freezing, sublimation, deposition, condensation, and vaporization.

Changes Between Liquids and Solids

How would you make ice cubes in a tray? First, you would fill the tray with water from a tap. Then you would place the tray in the freezer compartment of a refrigerator. The freezer is very cold. What happens next?

Freezing

Heat transfer occurs between the warmer tray and the colder air in the freezer. The warm water loses heat to the cold air in the freezer. This heat transfer occurs until no energy is available for the particles to slide past each other. This forces them to

remain in fixed positions, locked in place by the force of attraction between them. This way liquid water is changed into solid ice. The process of liquid water changing to solid ice is termed as freezing. The temperature at which it occurs is known as the freezing point.

Melting

If you took out the ice cubes from the freezer and placed them in a warm room, the ice would absorb energy from the warmer air around them. This absorbed energy would facilitate them to overcome the force of attraction holding them together, enabling them to slip out of the fixed position that they held as ice. The process in which a solids change to a liquid is called melting. The melting point is the temperature at which a solids change to a liquid.

Changes Between Liquids and Gases

If you fill a pot with cold tap water and heat it on a hot stovetop, the water heats up. Heat energy travels from the stovetop to the pot, and the water absorbs the energy from the pot. What happens to the water next?

Vaporization

If the water is hot enough, it starts to boil. Bubbles of water vapor are formed in the boiling water. This happens as particles of liquid water gain enough energy to completely overcome the force of attraction between them and change to the gaseous state. The bubbles rise through the water and escape from the pot as steam. The process in which a liquid boils and changes to a gas is called vaporization. The temperature at which a liquid boils at specific conditions is its boiling point.

Condensation

When you take a hot shower in a closed bathroom, the mirror is likely to fog up. You may wonder why does this happen? Some hot water from the shower evaporates and when it comes in contact with cooler surfaces such as the mirror, it cools and loses energy. The cooler water particles no longer have the energy to overcome the forces of attraction between them. They come together and form droplets of liquid water. This process in which a gas changes to liquid at specific condition is known as condensation.

Changes Between Solids and Gases

Solids that change to gas pass through the liquid state first. However, sometimes solids change directly to gases and skip the liquid state. The reverse can also occur. Sometimes gases change directly to solids.

Sublimation

The process in which solids directly change to gases is known as sublimation. This occurs when solids absorb enough energy to completely overcome the forces of attraction between them. Dry ice is an example of solids that undergo sublimation. The reverse process of sublimation is called **deposition**. Figure 3 summarizes the change of state of matter from one form to another form.

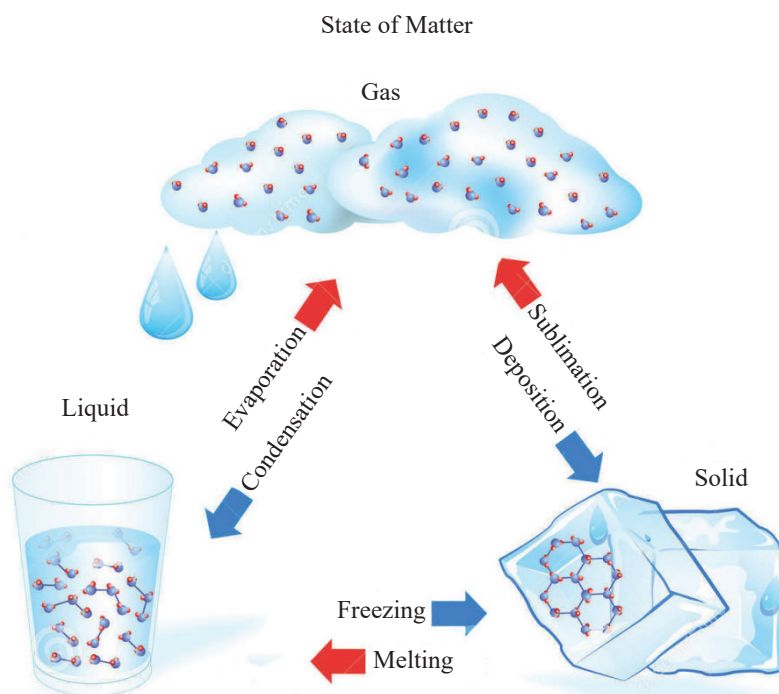


Figure 3. Change of the states of matter from one form into another form

2.2 PROPERTIES AND CHANGES OF MATTER

ACTIVITY 2

Discuss in groups and present your conclusion to the class.

1. How can you identify and differentiate (a) common salt from sugar (b) alcohol from water?
2. Take a sheet of paper and describe it in terms of color and physical state.
3. If you light a match and bring it close to the paper, the paper will start burning. Is the ability of paper to burn a physical property or a chemical property?
4. What are the color, odor and taste of pure water?

Properties are the characteristics that give a substance its identity and enable us to distinguish one kind of substance from another. No two substances have exactly the same set of properties.

ACTIVITY 3

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

Given the following physical properties of substances: odor, color, taste, melting point, boiling point and density.

1. Which of these physical properties have constant values under specific condition, such as temperature?
2. Which physical properties can be recognized directly by our sense organs?
3. Which of these properties are measured using instruments?
4. What will happen to ice kept in a cup in the classroom?

To identify a substance, chemists observe two distinct types of properties: **physical and chemical properties**.

Physical properties are the characteristics a substance shows without changing into a new substance. These properties are perceived by our sense organs or measured without changing the identity of the substance. Some physical properties are color, melting point, boiling point, density, physical state, and electrical conductivity.

Chemical properties are characteristics that describe the ability of a substance to form new substance. Examples of chemical properties include burning of a substance and rusting of iron.

Physical Properties Detected by Sense Organs

The properties of the substance you observed in the above activity are properties that can be perceived by the sense organs. Some examples are color, odor and taste.

Color: The color of a substance results from its interaction with light. Substances can be identified by their colors. For example, chalk is white, water is colorless, gold is yellow and so on.

Odor: Refers to the property of a substance perceived by the sense of smell. Terms commonly used to describe the odor of a substance are pungent, fragrant, spicy, fruity and odorless. For example, water is odourless, flowers are fragrant, orange smells fruity.

Caution! Care has to be taken in smelling substances as they may be harmful

Taste: Refers to physical properties that can be perceived by the taste buds of the tongue. The taste of a substance is usually described by terms like sweet, bitter, sour, salty, and tasteless. For example, honey is sweet, lemon is sour and table salt tastes salty.

Caution! Tasting can be used to identify substances only if the substance to be tasted is not harmful

Measurable Physical Properties

Measurable physical properties are the properties of a substance that can be measured using an appropriate apparatus. These physical properties have constant values under specific conditions. Examples are melting point, boiling point, density and electrical conductivity.

Melting Point: The temperature at which a solid substance changes to its liquid state. For example, ice is the solid form of water. Ice melts to liquid (water) at 0°C. Therefore, the melting point of ice is 0°C.

Boiling Point: The temperature at which the vapor pressure of the liquid equals the surrounding atmospheric pressure. At sea level water boils at 100°C.

Density: The mass per unit volume of a substance. It is expressed mathematically as:

$$\text{Density} = \frac{\text{Mass of the substance}}{\text{Volume of the substance}} \text{ or } d = \frac{m}{v}$$

Note that to determine the density of a substance we should know its mass and the volume it occupies.

The units of density are kilogram per cubic meter (kg/m^3), kilogram per litre (kg/L), gram per cubic centimeter (g/cm^3) or gram per milliliter (g/mL).

ACTIVITY 4

Discuss the following activity in groups and present your opinion to the class.

1. Which apparatus is used to measure the melting point and boiling point of a substance?
2. What do you call the temperature at which a liquid changes to its solid form?
3. Is the temperature at which the solid form of a substance melts same as the temperature at which its liquid form changes to the solid state?

Electrical Conductivity: Electrical conductivity is the ability of a substance to conduct electricity. This is a physical property mostly characteristic of metallic substances such as copper, aluminum, iron, silver, zinc etc.

Changes around Us

The substances in our environment constantly undergo changes. The decay of plants and animals, the burning of gasoline, the evaporation of lakes and seas, and the melting of snow are some of the changes that occur in our environment. Most of the changes in substances are classified into physical and chemical changes.

ACTIVITY 5

Discuss the following phenomena in groups and present your opinion to the class. Place a piece of chocolate or candy in your mouth and wait for a few minutes. Notice the changes carefully.

- (i) What happens to the chocolate or candy?
- (ii) Is there any change that occurs before and after you put the chocolate or candy in your mouth? Explain.

Physical change

ACTIVITY 6

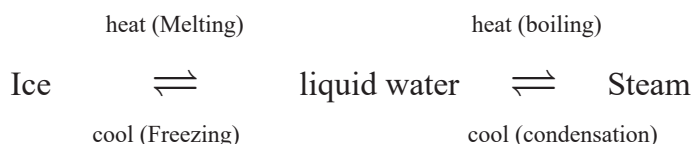
Perform the following tasks in groups and present your conclusions to the class.

1. Cut a piece of paper in four square pieces. Cut each square piece further into four square pieces. Lay these pieces on the floor or a table so that the pieces acquire the shape of the original piece of paper.
 - (i) Can you join the pieces back to the original piece
 - (ii) Is there any change in the property of the paper?
2. Collect the chalk dust lying on the floor near the chalkboard in your classroom or crush a small piece of chalk into dust. Add a little water to the dust to make a paste. Roll it into the shape of a piece of chalk. Let it dry. Did you recover chalk from the dust?

Changes that do not alter the composition of the substances are called **physical changes**. In physical changes, the original substance can be recovered or the change can easily be reversed. Changes like melting, freezing, dissolving and evaporating are physical changes. This is because in all of these changes new substances with new properties are not formed. In other words, the same original substances are observed before and after the changes.

Change of State

We know that matter exists in three physical states or forms: solid, liquid and gaseous states. By varying the temperature, substances can be changed from one physical state to another. Such type of a change is a physical change. Let's take water as an example. Water exists in solid (ice), liquid (water) and gaseous states. When ice is heated, it melts to liquid water, and on further heating the water changes to steam. Liquid water looks different from ice or steam. But the composition and identity of the original substance does not change because it is still water whatever its state is. Consider the following changes of states of water.



Solution (Dissolving Substances)

The process of dissolving a substance in water to form a solution is a physical change. For example, when common salt is dissolved in water, a clear salt solution is formed. If this mixture is heated, the water is evaporated, and the original salt remains as a residue. Thus, in the process of dissolution, no new substances are produced.

Mechanical Changes

Mechanical changes brought about by breaking, hammering, powdering, tearing and cutting of substances are physical changes. Mechanical changes involve changes in the size and shape of the material. Example: powdering of chalk.

Characteristics of a Physical Change

- It is a change in the physical property of the substance.
- No new substance is formed (the identity of the substance does not change).
- The change is easy to reverse (It is a reversible process).
- No change in the mass of the substance is involved.
- No change in composition occurs.
- Energy changes (heat changes) are not necessarily involved.

Chemical Changes

ACTIVITY 7

Discuss the following phenomena in groups and present your opinion to the class.

In your home, you may have observed the formation of rust on the surface of knives or nails that are made of iron.

1. Why is rust formed on the surfaces of knives or nails?
2. Using any process, do you think that this knife or nail can be reversed into the original condition? Explain.

If a piece of paper is burned in air, two observable products are formed: an ash and a smoke. These substances have different properties and composition from the original material. After burning, the composition of the sample is changed because it no longer exists as a paper. Such type of change that results in the formation of new substances with new properties and composition is known as a **chemical change**.

A chemical change is a chemical property of a substance. The change is generally difficult to reverse. Some common examples of chemical changes are digestion of food, growth of plants, burning of wood, rusting of iron, decaying of food and exploding of gun powder.

Characteristics of a Chemical Change

- The change in which the molecular composition is completely altered and a new product is formed is called a chemical change.
- Chemical changes create a new product.
- Chemical changes are irreversible and permanent.
- It reveals that chemical change cannot be reversed by changing or altering the experimental changes.
- The mass of the substance is altered during a chemical change. Either the mass is added or removed.

Experiment 1

Title: Melting sulphur.

Objective: To determine whether the melting of sulphur is a physical or a chemical change.

Materials required: Test tube, watch glass, Bunsen burner, tong, powdered sulphur.

Procedure

1. Put powdered sulphur to half a test tube
2. Heat the test tube gently until all the sulphur melts as shown in Figure 4.
3. Observe the color of the molten sulphur

4. Pour the molten sulphur on a watch glass and allow it to cool.
5. Let the watch glass stay for some days, and observe the change again.

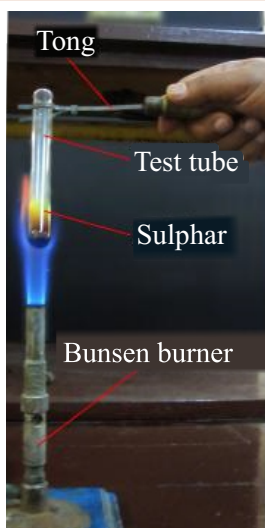


Figure 4. Melting of Sulphur

Observations and Analysis

- (i) Name all the colors of sulphur you observed before and after the experiment.
- (ii) Is the change physical or chemical?

Experiment 2

Title: Rusting of iron

Objective: To investigate the type of change that occurs during rusting of iron.

Materials Required: Iron nails, test tube, Rack and plain water

Procedure

1. Pour some amount of tap water into the test tube.
2. Put 3 or 4 clean and shiny iron nails into the test tube as shown in Figure 5.
3. Let the test tube stay for a few days in open air.
4. Observe the change that takes place.

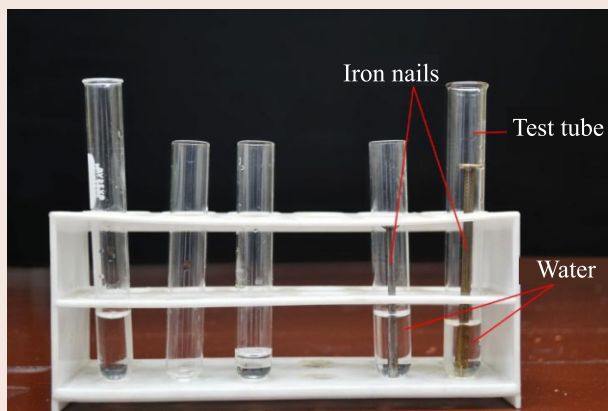


Figure 5. Rusting of iron

Observation and Analysis

- (i) What happened to the surface of the iron nail?
- (ii) What color do you observe on the iron nail?
- (iii) Is the change physical or chemical? Why?

Exercise

Identify the following as physical (P) or chemical (C) changes.

1. NaCl (table salt) dissolves in water.
2. Milk sours.
3. Ag (Silver) tarnishes.
4. Sugar dissolves in water.
5. An apple is cut.
6. Wood rots.
7. Heat changes H_2O to steam.
8. Pancakes cook.
9. Baking soda reacts to vinegar.
10. Grass grows.
11. Fe (Iron) rusts.
12. A tire is inflated.
13. Alcohol evaporates.
14. Paper towel absorbs water.
15. Ice melts.

2.3 STANDARD SEPARATION TECHNIQUES OF MIXTURE

There are various ways of classification of substances in chemistry. At this level, we classify substances based on their composition and properties. Substances are classified as **pure substances** and **mixtures**. A pure substance is a form of matter that has constant composition and uniform properties throughout the sample. It may contain either a single component or two or more components.

Mixtures

ACTIVITY 8

Discuss the following ideas in groups. Present your opinion to the class.

- (i) What is the difference between pure water and a solution of sodium chloride in water.

(ii) Do you think air is a pure substance or a mixture?

Mixtures are substances that consist of two or more pure substances, in which the substances retain their identity. Mixtures are not pure substances because they have variable composition and no unique set of properties. They can be separated by physical means.

For example, when sodium chloride (common salt) is mixed with water, it forms a salt solution.

This salt solution is a mixture because of the following main reasons:

- (i) The components, salt and water, are simply physically mixed. i.e they are not chemically combined.
- (ii) This salt solution retains many of the properties of its constituents, sodium chloride and water. In other words, the salt solution shows the properties of the salt and the water. For example, the solution is colorless like water, and also it tastes salty like the original sodium chloride salt.
- (iii) The salt solution can be separated into salt and water by physical processes like evaporation or distillation.

Therefore, all the above characteristics confirm that a salt solution is a mixture. Some examples of mixtures are milk, sea-water, soil, ink, wood and petroleum.

Separation of Mixtures and its Application

Most of the substances around us exist in the form of mixtures. However, these mixtures can be separated into pure substances using various separation techniques. The methods used to separate mixtures are mainly physical processes. This is because the substances in a mixture are physically combined.

Some of the methods used to separate mixtures are filtration, evaporation, magnetic separation, decantation and distillation. Note that the methods for the separation of mixtures into their components depend on the differences in the density, melting point, boiling point, solubility, etc. of the components.

The most important separation techniques for a given mixture depend on: the type of mixture, and which substance in a mixture we are more interested in. For example, in a mixture of salt solution, if we need to separate only the salt, evaporation can be used. But to get both salt and water separately the distillation process is applied.

We will discuss some of the methods that are used to separate the components of **solid-solid**, **solid-liquid** and **liquid-liquid mixtures**.

Separation of Mixture of two Solids (solid – solid mixtures)

There different methods of separating solid – solid mixtures. Some of the are discussed as follow.

1. Mechanical Picking or Handpicking:

In this method, components are separated on the basis of physical characteristics like size and shape. Example- Removing stones from grains. This method can be used to separate slightly larger pieces of dirt, stones and husk from the grains because the quantity of such impurities is usually not very large.

2. By using a suitable solvent:

Here, a mixture of two solids can be separated from each other by using a solvent which will make one of the substances soluble in the solvent, thus separating the two substances. E.g. A mixture of sugar and sand can be separated by using water as a solvent as sugar is soluble in water and sand is insoluble in water. This difference in the solubility of sugar and sand in water is used to separate them.

3. Separation by sublimation:

The changing of a solid directly into vapours on heating, and of vapours into solid on cooling is called sublimation. The process of sublimation is used to separate those substances from a mixture that sublime on heating. Examples: Ammonium chloride, iodine, camphor, naphthalene etc. Sublimation is defined as heat or energy required to change the state from solid to gas.

4. Magnetic Separation:

If a mixture contains iron as one of the constituents, it can be separated by using a magnet. For example: A mixture of iron filings and Sulphur powder can be separated by a magnet. Iron filings will get attracted to the magnet and sulphur won't. This technique is useful for few minerals which contain magnetic properties and also for minerals with paramagnetic (materials that are less affected) properties. Not all metals are magnetic; gold, silver and aluminium are some examples.

Experiment 3

Title: Separation of a mixture using a magnet.

Objective: To separate a mixture of iron filings and sulfur using a magnet.

Materials Required: Magnet, iron filings, powdered sulphur, beaker, sheet of paper, spatula.

Procedure

1. Take two spatulas of each of iron filings and powdered sulphur, and mix them thoroughly, in a beaker
2. Place some of this mixture on a sheet of paper as shown in .
3. Bring a magnet close to the mixture as shown in the figure

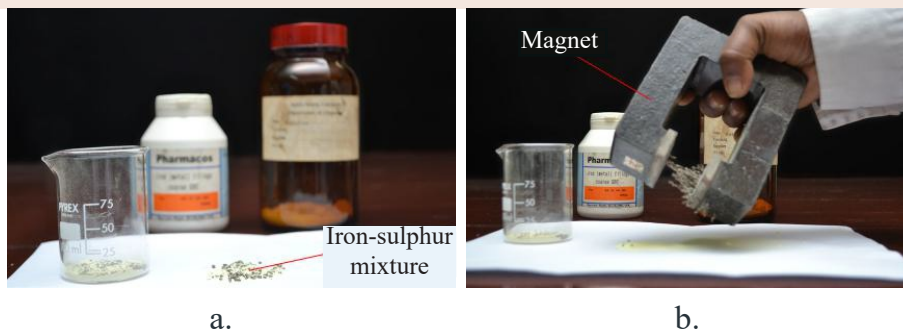


Figure 6. Magnetic separation

Observation and Analysis

- (i) Which component of the mixture is attached to the magnet?
- (ii) What can you conclude from the experiment?

Gravity Separation

A mixture containing two components having different densities but both being insoluble in a solvent can be separated by using this method. Example- Mixture of chalk powder and sand in water. As sand is denser, they settle at the bottom of the water container, while chalk being less dense floats in it. The ores of metals are heavier than gangue particles. This process uses this property for the concentration of ore. Hematite ore is concentrated using the gravity separation process. Hematite is a common ore of iron oxide.

Separation of Mixture of Solid and Liquid

A mixture of solid and liquid can be found in a homogeneous solution or heterogeneous mixture. The liquid in which the solid substance is mixed or gets dissolved is known as a **solvent**. Whereas, the dissolving solid constituent is known as **solute**. These solute and solvents can be separated in their pure form by the usage of different techniques for the separation of mixtures.

Some common methods to separate solid and liquid from a mixture are mentioned below:

- Evaporation,
- Filtration,
- Sedimentation, and
- Crystallization.

Evaporation

ACTIVITY 9

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

Take a beaker full of water outside your classroom and keep it in sunlight. Wait for a day to observe a change.

- (a) What happened to the level of the water in the beaker? Increased, decreased or remains the same? Explain.
- (b) What can you conclude from this phenomenon?

Evaporation is the method of separating a solid (solute) from a homogeneous solution. In this method, the solution is heated until the solvent gets evaporated in the form of vapor and the solute is left behind as residue. For example, sea water is a solution of salts. When it is heated on an evaporating dish, the level of the liquid slowly decreases. This is because some of the water changes into vapor. This vapor will then escape into the atmosphere, leaving behind the salts as shown in Figure 7. This process of escaping of water in the form of vapor is called evaporation.

Experiment 4

Title: Evaporation

Objective: To separate salt from a salt solution.

Materials Required: Burner, evaporating dish, wire gauze, salt, beaker, watch glass, beam balance, measuring cylinder.

Procedure

1. Dissolve about 10 g of common salt in 30 mL of tap water in a beaker.
2. Pour the salt solution in an evaporating dish as shown in Figure 7.
3. Boil the solution until all the liquid evaporates and observe the results.

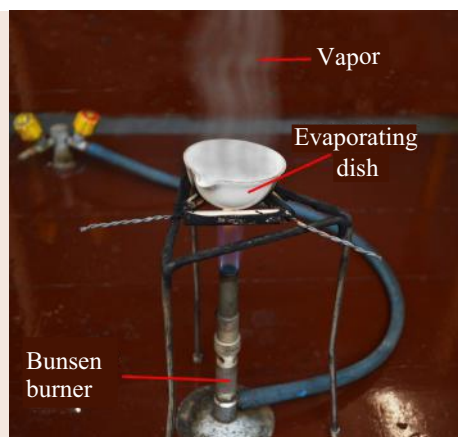


Figure 7. Evaporation of a solution

Observations and Analysis

1. What did you observe in the evaporating dish?
2. What would happen to the level of the liquid if the evaporating dish is covered with a watch glass? Is evaporation possible?

Filtration

Filtration is a method used to separate the components of a mixture containing an insoluble solid and a liquid. When the mixture of solid and liquid is poured into a filtering material the liquid passes through it and the solid is retained. In the laboratory, a filtration process is performed by pouring the mixture into a funnel fitted with a filter paper.

A filter paper has a lot of fine holes that allow only the liquid to pass through, but not the solid particles. The liquid which passes through the filter paper is called the *filtrate*, and the solid which remains on the filter paper is known as the *residue*. Figure 8. illustrates a filtration process.

Filtration can be used to separate mixtures like soil and water, sand and salt solution, powdered chalk and water, etc. In practical application, filtration is a key step in the purification of the tap water you drink.

Sedimentation or Decantation

This method of separation is used when one component of a mixture is a liquid and the other one is an insoluble solid denser than the liquid component. For example, a mixture of sand and water can be separated using decantation.

During the decantation process, first the mixture is allowed to stand in a beaker for some time. The insoluble solid settles down to the bottom of the beaker. This is called **sedimentation**. The liquid above the solid can be poured carefully into another beaker as shown in Figure 9. Such a method of separation is called decantation.

ACTIVITY 10

Form a group and discuss the following concepts. Present your opinion to the class. Consider a mixture of water, sand, and oil.

- (i) How many phases are present?
- (ii) How can you separate this mixture into the individual components?

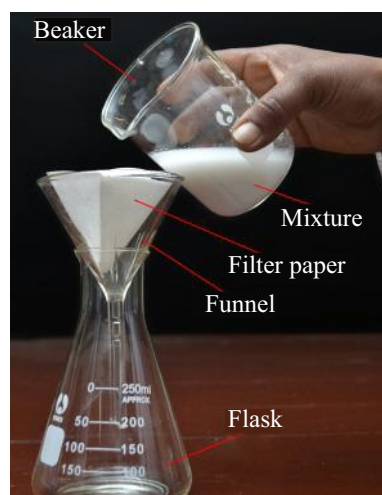


Figure 8. Filtration

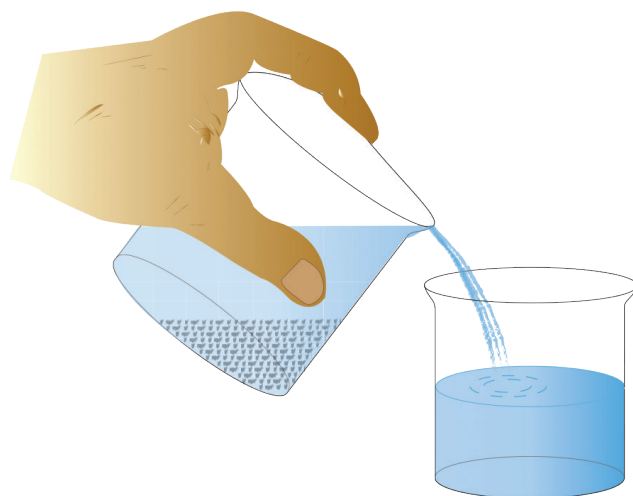


Figure 9. Decantation to separate solid from a solid-liquid mixture

The decantation process is also used to separate two immiscible liquids like oil and water. Immiscible liquids do not mix and thus they form two layers. A mixture of oil and water can be separated using a separatory funnel. When the mixture is poured into a separatory funnel, the oil and water separate into two distinct layers as shown in Figure 10. Since water is denser than oil, it is collected at the bottom of the funnel.

Crystallization

Crystallization is the process of separating a solid from a liquid solution. This process is carried out by cooling highly concentrated solutions. Crystallization is generally used to obtain the pure form or crystal of any substance. In the process, crystals are extracted from a supersaturated solution, and the liquid left behind is termed as *mother liquor*.

Separation of Liquid – liquid mixtures

ACTIVITY 11

Discuss the following concept in groups and present your opinion to the class.

You are given a homogeneous mixture like sugar solution and you are asked to separate and collect the sugar crystal and water separately.

- (i) Is it possible to separate them by evaporation? If not, why?
- (ii) Can you suggest any other separation method?

Distillation

Distillation is a fast way to separate mixtures of two or more pure liquids. Distillation is a purification method that involves vaporizing the constituents of a liquid mixture, then condensing and isolating them. A mixture is heated in simple distillation, and the most volatile component vaporizes at the lowest temperature. The vapor condenses back into liquid after passing through a cooled tube (a condenser). The collected condensate is known as **distillate**.

For example, to get pure water from sea water, simple distillation can be used. When a sample of sea water is heated in a distillation flask, the water vaporizes and boils off at a much lower temperature than the salt. The vapor is condensed to a pure liquid water and collected in the receiver. The salt remains as a residue in the distillation flask.

A mixture of two miscible liquids can also be separated by **simple distillation**. Liquids which mix with each other to form a solution are called **miscible liquids**. The mixture of alcohol and water, benzene and oil are some examples of miscible liquids.

Consider a mixture of alcohol and water. Ethanol, which is an alcohol boils at 78°C , and water boils at 100°C . When the mixture is heated, the alcohol, which has the lower boiling point vaporizes more rapidly than the water. The vapor of alcohol passes through the condenser and then collected as a distillate in the receiver.

Experiment 5

Title: Simple distillation.

Objective: To separate water from salt solution.

Materials Required: Distillation flask, condenser, Wire gauze, Bunsen burner, conical flask, beaker.

Procedure

1. Dissolve about 40 g of common salt in 150 mL of pure water.
2. Set up the distillation apparatus as shown in Figure 10
3. Add 100 mL of the salt solution into the distillation flask.
4. Put a porous material or sand (boiling chips if there are any) in the flask.
5. Heat the distillation flask gently and observe the results.

Observations and Analysis

1. What is collected in the receiver (conical flask) and what remains in the distillation flask?

2. Give the names of the apparatus used for the evaporation and condensation processes in the experiment?
3. Why is the condenser connected to tap water in a simple distillation set up?

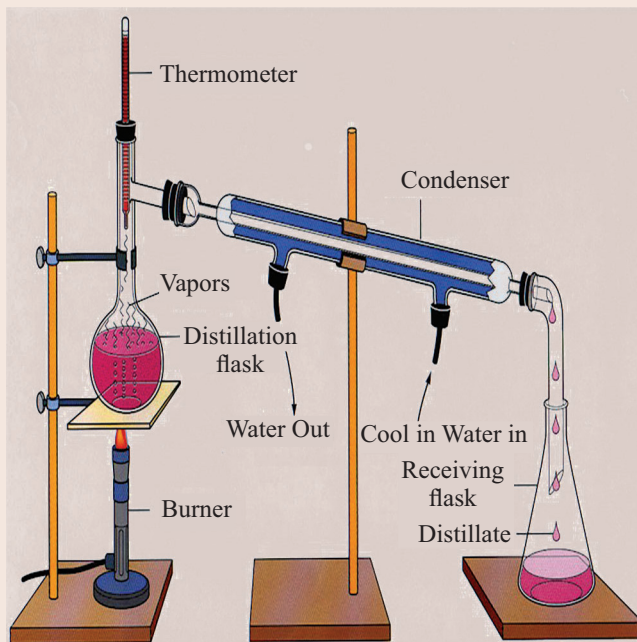


Figure 10. Simple Distillation

Exercise

How would you separate the following mixtures? Describe the separation techniques in each case.

1. Sand from water
2. Alcohol from water
3. Components in muddy water
4. Two colorless liquids
5. Sand from salt
6. Mixture of sugar and sand
7. Mixture of oil, water and sand

2.4 CLASSIFICATIONS OF MIXTURES

Mixtures are classified into two types: homogeneous mixtures (solutions), and heterogeneous mixtures.

Homogeneous Mixture (Solution)

A homogeneous mixture (also known as solution) is a mixture in which the components cannot be seen by our naked eyes or using a microscope or a magnifying glass. A homogeneous mixture has a uniform composition and properties throughout. It contains only one phase. A **phase** is a region with a uniform set of properties. A solution is a homogeneous mixture. For example, a mixture of sugar and water (sugar solution) is a homogeneous mixture because all the parts of the solution have the same sugar water composition. The sugars are totally dispersed into the water and cannot be seen separately. Also there is no visible boundary of separation between the sugar and water components in a solution.

In a homogeneous mixture, the proportions of components may vary from sample to sample like in a body of sea water, but each sample is uniform in appearance. Alloys are homogeneous mixtures of metals. For example, brass is an alloy of two metals: copper and zinc. The following table shows examples of homogeneous mixtures.

Table 2 Examples of homogeneous mixtures

Mixture	Examples
Solution of solid in liquid	Sea water, sugar in water, salt solution
Solution of two miscible liquids	Alcohol and water, benzene and kerosene
Solution of gas in liquid	Soft drinks, mineral water, soda water
Mixture of gas in gas	Air
Mixture of solid in solid	Brass (mixture of copper and zinc), bronze (mixture of copper and tin)

Heterogeneous Mixtures

A mixture that does not have a uniform composition throughout is known as a **heterogeneous mixture**. The components of a heterogeneous mixture can be

ACTIVITY 12

Discuss the following phenomenon in groups and present your opinion to the class.

1. Add a teaspoon full of salt to a cup of warm water, what did you observe? Can you see the components of the mixture separately?
2. Now the mixture is stirred until the salt crystals disappear. What is the difference between these two mixtures? (Before and after stirring).

identified by our naked eyes or with the help of a microscope or a magnifying glass. They contain more than one phase.

For example, a mixture of sand and iron filings is a heterogeneous mixture. This is because the sand particles remain visible and physically separated. Blood is also a heterogeneous mixture of plasma, red blood cells, white blood cells, etc. that can be seen using a microscope. Table 3 shows some heterogeneous mixture with examples.

Table 3 Examples of heterogeneous mixtures

Mixture	Examples
Insoluble solid in liquid	Mixture of sand and water, river water, muddy water
Two immiscible liquids	Benzene and water, oil and water
Solid in gas	Dusty air, smoke

Table 4 Shows the difference between homogeneous and heterogeneous mixtures.

Table 4 Differences between homogeneous and heterogeneous mixtures

Homogeneous mixture	Heterogeneous mixture
It has a uniform composition throughout	It does not have a uniform composition throughout.
It has no visible boundaries of separation between the components	It has visible boundaries of separation between the components
Its components cannot be seen by our naked eyes or by microscope	Its components can be identified by our naked eyes
It consists of only one phase	It contains more than one phase

Exercise

Which of the following are homogeneous mixtures, and which of them are heterogeneous mixtures?

- Soil
- Mixture of oil and water
- Mixture of salt and sugar
- Sugar dissolved in Tea
- Mixture of alcohol and water

KEY TERMS

- Chemical property
- Decantation
- Distillation
- Evaporation
- Filtration
- Gas
- Heterogeneous mixtures
- Homogeneous mixtures
- Liquid
- Magnetic separation
- Mixtures
- Physical property
- Solid

SUMMARY

- A substance is a form of matter possessing constant properties under specific conditions.
- Properties of substances are subdivided into physical and chemical properties.
- Physical properties of a substance describe the characteristics of the substance that are related to physical changes.
- Substances can be identified based on their physical properties such as state, color, odor, taste, hardness, density, melting and boiling points.
- A mixture is a substance which consists of two or more pure substances that are mixed together physically.
- Mixtures are classified into homogeneous and heterogeneous mixtures
- Substances undergo two types of changes: physical and chemical changes.
- Physical changes are changes that do not result in the formation of new substances with new properties.
- Chemical changes are the changes in the composition of substances.
- Mixtures can be separated into their components using different methods such as magnetic separation, filtration, decantation, evaporation and distillation.

Exercises**Part I. Write true for the correct statements and false for the wrong statements.**

1. A physical change is a change in the composition of a substance.
2. A homogeneous mixture always contains only one phase.
3. Two different substances cannot have similar physical properties.
4. The properties of a substance that can be observed without altering its identity are physical properties.
5. The ability of iron to conduct electricity is a chemical property of iron.
6. The temperature at which a solid melts is the same as the temperature at which its liquid solidifies (freezes).
7. Physical properties are used as a basis for identifying substances.
8. Rusting of iron is a physical change.
9. The mixture of oil and water can be separated by crystallization.
10. Decantation is a technique of separating the mixture of two solids.

Part II. Choose the best answer from the given alternatives.

11. Which one of the following does not involve a change of state?
 - (a) Freezing
 - (b) Melting
 - (c) Expansion
 - (d) Boiling
12. Which one of the following processes does not help in separation of mixtures?
 - (a) Filtration
 - (b) Distillation
 - (c) Melting
 - (d) Sieving
13. The apparatus used to separate two immiscible liquids is
 - (a) Separatory funnel
 - (b) Condenser
 - (c) Filter funnel
 - (d) Filter paper
14. The two stages that take place during distillation are
 - (a) filtration followed by evaporation
 - (b) condensation followed by evaporation
 - (c) evaporation followed by condensation
 - (d) filtration followed by condensation

15. The process of settling of suspended particles is known as
 - (a) decantation
 - (b) distillation
 - (c) sedimentation
 - (d) magnetic separation
16. The substance that can not be further decomposed by ordinary chemical means is
 - (a) water
 - (b) gold
 - (c) sugar
 - (d) air

Part III. Give short answers for each of the following questions.

17. What is the difference between a physical and a chemical change?
18. What are the physical properties that can be detected by sense organs?
19. Which physical properties can be used to distinguish water from alcohol?
20. How would you distinguish between homogenous and heterogeneous mixtures? Give three examples for each case.
21. Identify whether the following statements describe a physical or a chemical property of a substance.
 - (a) Charcoal burns in air to form carbon dioxide.
 - (b) Ethyl alcohol boils at 78°C .
 - (c) Sulphur is a yellow solid at room temperature.
 - (d) Mercury is a silvery-white liquid metal.
 - (e) Silver is the best conductor of electricity.
 - (f) Iron rusts in the presence of moisture and air to form an oxide.
 - (g) Benzene is an inflammable liquid.
22. Which of the following are chemical changes and which of them are physical change?
 - (a) Hammering aluminium into thin sheets
 - (b) Dissolving table salt in water
 - (c) Melting gold to make jewelry
 - (d) Burning of wood
 - (e) Grinding of maize
 - (f) Tearing of paper
23. What apparatuses are used to carry out the following processes?
 - (a) filtration
 - (b) evaporation
 - (c) distillation



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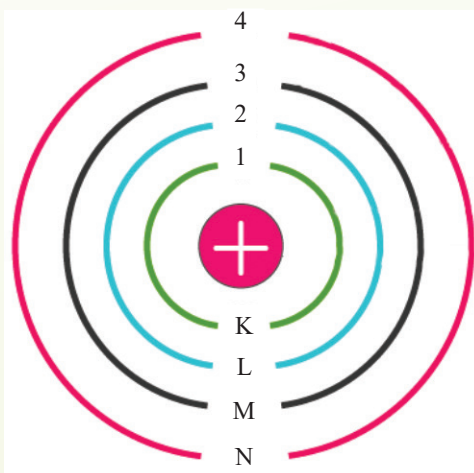
CHAPTER

3

ATOMIC STRUCTURE

Chapter Contents

- 3.1 Historical Development of the Atomic Nature of Substances
- 3.2 The Atomic Theory
- 3.3 Discovery of the Nucleus and Electrons
- 3.4 Atomic Number, Mass Number and Isotopes
- 3.5 Relative Atomic Mass
- 3.6 Quantum numbers
- 3.7 Energy Levels and Electronic Configuration
- 3.8 Rules and principles for filling in electrons.
 - Key Terms
 - Summary
 - Exercises



Chapter outcomes

Upon completion of this chapter, students will be able to:

- discuss contributors to Atomic structures;
- explain the arrangement of fundamental particles;
- discuss the concept of Atomic Theories;
- discuss atomic number and mass number and their relationship to isotopes;
- discuss the four quantum numbers;
- construct electronic configurations for atoms;
- explain the rules and principles for filling in electrons.

ACTIVITY 1

Discuss the following ideas in group and present your findings to the class.

1. What would you obtain if you continually subdivided a piece of chalk, paper or any other substances?
2. What do you think are all substance made of?

Introduction

One of the oldest concepts of matter is that all substances are made up of small, invisible, presumably indivisible and indestructible particles. The ancient Greeks thought that these particles were in constant motion and able to fit together into stable combinations. The characteristics or properties of a particular substance resulted from the different size, shapes, and arrangements of these particles. Today, we refer to these building blocks of all substances as atoms.

3.1 HISTORICAL DEVELOPMENT OF THE ATOMIC NATURE OF SUBSTANCES

Ancient philosophers were curious about the structure of substances. The Greeks first put forward the idea atoms. The idea was not based on any experimental facts. To the ancient Greek philosophers, the problem of the structure of the substances presented itself as a mere choice between two alternatives. Substances are either **continuous or discrete**. Some of them argued that matter is **continuous** i.e., it could be divided endlessly into smaller pieces. Others believed that matter is **discrete**; i.e., it cannot be infinitely divided.

Democritus (460 - 370 B.C) thought that substances are made up of tiny indivisible particles. If a piece of copper is cut into smaller and smaller pieces one finally reaches a point where it can no longer be divided. He called these particles **atoms**. The word atom has its origin from the Greek word “**Atomos**” which means indivisible. Thus according to Democritus matter is discrete.

On the other hand *Aristotle (384 – 322 B.C)* argued that matter is continuous. According to Aristotle and his followers, there was no particle of matter so small that it could not be subdivided into still smaller pieces. Table 1 shows the comparison between the discrete and continuous theory of matter.

Table 1 Comparison between the discrete and continuous theory of matter

Discreteness of Matter		Continuity of Matter	
Democritus		Aristotle	
Matter is discrete		Matter is continuous	
There is a limit to which matter is broken		Matter is infinitely divisible	
Believed in the existence of atoms		Rejected the idea of atoms	

3.2 THE ATOMIC THEORY

In the fifth century B.C, the Greek philosopher Democritus expressed the belief that all matter consists of very small, indivisible particles, which he named atoms (meaning uncuttable or indivisible). Although Democritus' idea was not accepted by many of his contemporaries (notably Plato and Aristotle), somehow it endured. Experimental evidence from early scientific investigations provided support for the notion of "atomism" and gradually gave rise to the modern definitions of elements and compounds. It was in 1808 that an English scientist and schoolteacher, John Dalton, formulated a precise definition of the indivisible building blocks of matter that we call atoms.

Dalton's atomic theory

Dalton's work marked the beginning of the modern era of chemistry. The hypotheses about the nature of matter on which Dalton's atomic theory is based can be summarized as:

1. Matter consists of very small particles called atoms.
2. Atoms are indivisible.
3. All atoms of a given element are identical, they have the same mass and properties.
4. Atoms of one element are different from atoms of other elements in mass and other properties.
5. A chemical reaction involves only the separation, combination, or rearrangement of atoms; it does not result in their creation or destruction.
6. Atoms combine in a fixed small whole number ratio to form compounds. For example in carbon dioxide one atom of carbon combines with 2 atoms of oxygen, in the ratio 1:2

Shortcomings of Dalton's atomic theory

Dalton's atomic theory has certain shortcomings, when evaluated in light of further new experimental findings made after the time of Dalton on the nature of atom.

Because of the discovery of the subatomic particles, Dalton's atomic theory was amended. The following two statements in Dalton's atomic theory were wrong:

- (i) **Atoms are indivisible:** This statement is proved to be wrong because atoms can be divided into subatomic particles.
- (ii) **Atoms of the same element have the same mass:** Atoms of the same element may not have the same mass as explained in the modern atomic theory. These facts will be explained in the modern atomic theory.

Even though some of Dalton's statements are not entirely correct, the particle nature of matter and the existence of atoms are still now accepted. That is why Dalton is considered as the father of modern atomic theory.

The modern atomic theory

ACTIVITY 2

Discuss the following ideas based on Dalton's atomic theory.

1. All atoms of oxygen have the same mass. Is this statement true today?
2. Can an atom of oxygen be broken into two halves to make smaller oxygen atoms?

After the time of Dalton, further scientific research brought about new discoveries that led to the development of the modern atomic theory.

The following statements describe the main points of the modern atomic theory.

1. An element is made of atoms. An element is composed of small particles called atoms.
2. Atoms are built up from small particles called protons, electrons and neutrons. Thus atoms are divisible and are not the smallest particle of a substance.
3. All atoms of the same element have the same number of protons (and electrons) but may have different number of neutrons. This is due to the existence of isotopes of elements. However atoms of the same element have identical chemical properties. We will discuss about isotopes in the next section.
4. Atoms of different elements are different because of the different number of protons. Hence they have different chemical properties

Exercise

Give short answer to the following questions.

1. List the main ideas included in Dalton's atomic theory.
2. What are the shortcomings of Dalton's atomic theory?
3. According to Dalton, all atoms of carbon must have the same _____.
4. According to the modern atomic theory, atoms can be broken down into _____, _____ and _____.
5. According to the modern atomic theory, atoms of the same element may have different _____.

3.3 DISCOVERY OF THE NUCLEUS AND ELECTRONS

Although Dalton had postulated that atoms were indivisible particles, experiments conducted around the beginning 21st century showed that atoms themselves consist of particles. These experiments showed that an atom consists of two kinds of particles: a **nucleus**, the atom's central core, which is positively charged and contains most of the atom's mass, and one or more **electrons**. An electron is a very light, negatively charged particle that exists in the region around the atom's positively charged nucleus.

Discovery of the electrons

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

In 1879, the English scientist William Crookes (1832-1919) experimented with gas discharge tubes. When a very high electrical potential (~ 10,000 volts) is applied across a gas taken in a discharge tube of a very low pressure (~ 0.001 torr) some radiations are emitted from cathode. These radiations are called **cathode rays**. Fig 1a, shows emission of cathode rays in a discharge tube. At this stage the glass walls of the discharge tube opposite to the cathode starts glowing with a faint greenish light. It is now known that this greenish glow on the walls is due to the bombardment of the glass wall with the cathode rays.

Cathode rays normally travel in straight lines, but are deflected when a magnet is brought nearby (Figure 1b).

An English physicist Joseph John Thomson (1856-1940) in 1897 studied the behavior of cathode rays in electric and magnetic fields, Thomson established clearly that the rays consist of negatively-charged particles. Moreover, his experiments showed that the particles were identical, regardless of the materials from which the electrodes were made or the type of gas in the tube.

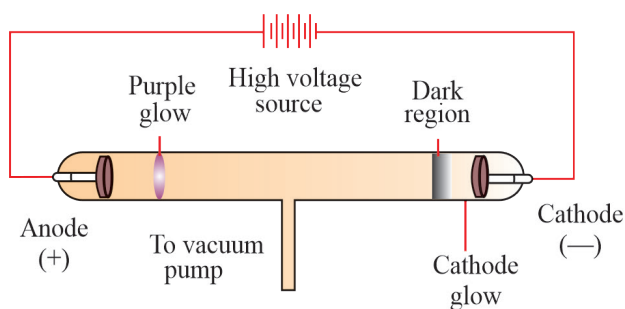
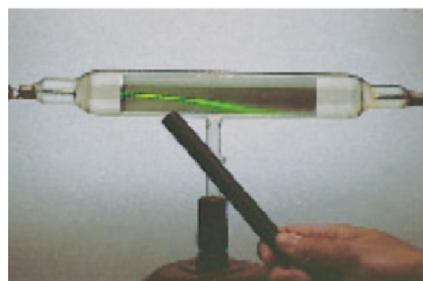


Figure 1. (a) Electric discharge in an evacuated tube



(b) The cathode ray is bent in the presence of a magnet

Thomson concluded that these negatively charged particles were constituents of every kind of atom. We now call these particles electrons, a term that had been coined by the Irish Physicist George Stoney in 1891 to describe the smallest unit of electric charge. Cathode rays are beams of electrons.

Perhaps Thomson's most significant experiment was the one illustrated and described in Figure 2. By measuring the amount of deflection of a cathode ray beam in electric and magnetic fields of known strengths, Thomson was able to calculate the ratio of the mass of an electron to its charge. If we denote the mass of the electron as m_e and its electrical charge as e , the mass-to-charge ratio is m_e/e . This ratio has a value of $-5.686 \times 10^{-12} \text{ kg C}^{-1}$ (kilograms per coulomb).

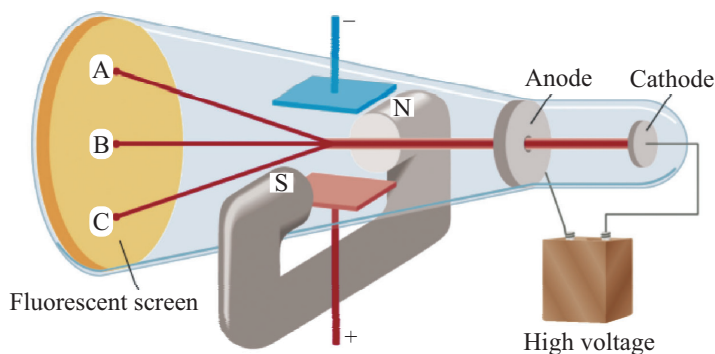


Figure 2. Thomson's apparatus for determining the charge to mass ratio.

In 1909, Robert A. Millikan, an American physicist, determined the charge on the electron by observing the behaviour of electrically-charged oil drops in an electric field. Based on careful experiments, Millikan established the charge on an electron as $e = -1.602 \times 10^{-19} \text{ C}$. From this value and the value for m_e/e , we can calculate the mass of an electron.

$$m_e = m_e/e^- = -5.686 \times 10^{-12} \text{ kg C}^{-1} \times -1.602 \times 10^{-19} \text{ C} \\ = 9.109 \times 10^{-31} \text{ kg}$$

Where m_e mass of electron, m_e/e^- mass to charge ratio of electron e^- charge on electron. In 1923, Millikan was awarded the Nobel Prize in physics for his oil-drop experiment.

Thomson's model of an atom

- According to Thomson,
 - (i) An atom consists of a positively charged sphere and the electrons are embedded in it.
 - (ii) The negative and positive charges are equal in magnitude. So, the atom as a whole is electrically neutral.
- The first model of an atom to be put forward and taken into consideration.
- He proposed a model of the atom be similar to that of a Christmas pudding/watermelon.
- The red edible part of the watermelon is compared with the positive charge in the atom.
- The black seeds in the watermelon are compared with the electrons which are embedded on it. Figure 3 shows the Thomson's model of an atom.

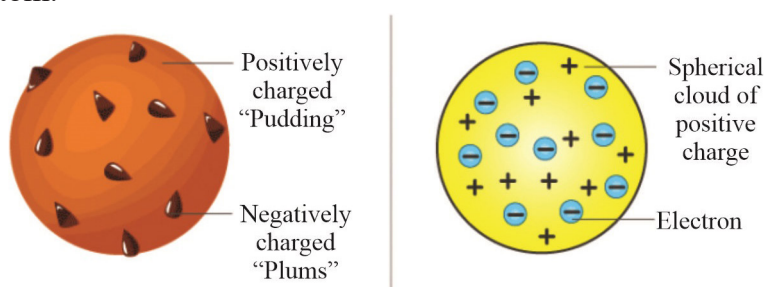


Figure 3. Thomson's Model of an atom

Limitations of Thomson's atomic model:

- It failed to explain the stability of an atom because his model of atom failed to explain how a positive charge holds the negatively charged electrons in an atom. Therefore, This theory also failed to account for the position of the nucleus in an atom.
- Thomson's model failed to explain the scattering of alpha particles by thin metal foils.
- No experimental evidence in its support.

Discovery of nucleus

In 1911, the New Zealand chemist and physicist, Ernest Rutherford, who had studied with J.J. Thomson decided to use α -particles to probe the structure of atoms. Together with his associate Hans Geiger, Rutherford carried out a series of experiments using very thin foils of gold and other metals as targets for α -particles from a radioactive source.

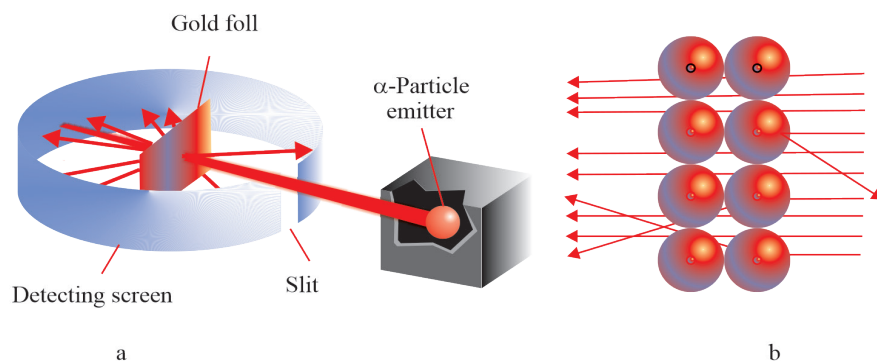


Figure 4. a) Rutherford's experimental design for measuring the scattering of α -particles by a piece of gold foil, b) Magnified view of α -particles passing through and being deflected by the nucleus.

ACTIVITY 3

Form a group and discuss Rutherford's experiment as shown in Figure 4 a and b answer the following questions:

1. Why majority of the α -particles penetrating the foil were undeflected?
2. Why small fraction of the α -particles showed slight deflection?
3. Why all α -particles did not bounced by an angle of 180° ?
4. Based on the finding of Rutherford's experiment how will you conclude that most of the space in the atom is empty?
5. Which observation led to conclusion that all the positive charge in the atom is concentrated in the nucleus?
6. Report your findings to the class.

In 1914, Rutherford suggested that the smallest positive-ray particle was the fundamental unit of positive charge in all nuclei. He called this particle, which has a charge equal in magnitude but opposite in sign to that of an electron, a **proton**. A proton has a mass of $m_p = 1.67262 \times 10^{-27}$ kg, which is about 1840 times the mass of oppositely-charged electrons. Rutherford proposed that protons constitute the positively-charged matter in the nuclei of all atoms.

Rutherford's model of an atom

Rutherford concluded the model of the atom from the α -particle scattering experiment as:

- (i) There is a positively charged centre in an atom called the nucleus. Nearly all the mass of an atom resides in the nucleus.
- (ii) The electrons revolve around the nucleus in well-defined orbits.
- (iii) The size of the nucleus is very small as compared to the size of the atom.

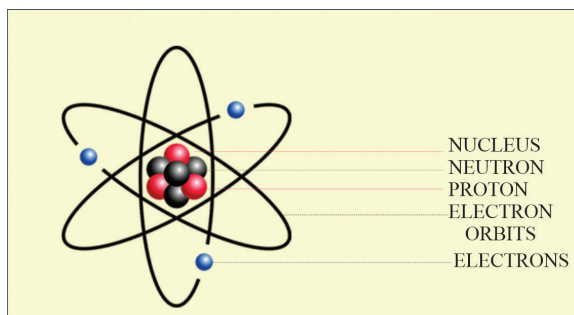


Figure 5. Rutherford's Model of an atom

Drawbacks of Rutherford's model

Rutherford explained that the electrons in an atom revolve around the nucleus in well-defined orbits. Particles in a circular orbit would experience acceleration. Thus, the revolving electron would lose energy and finally fall into the nucleus. But this cannot take place as the atom would be unstable and matter would not exist in the form we know.

Discovery of neutron

In the 1920s and early 1930s, alpha-particles were used as projectiles to bombard a variety of materials. Bombardment of beryllium atoms produced a strange, highly penetrating form of radiation. In 1932, James Chadwick (1891-1972) showed that this radiation was best explained as a beam of neutral particles. These particles, called neutrons, were found to have about the same mass as protons but no electric charge. This discovery finally provided an explanation for the mysterious excess mass. A helium atom has two protons and two neutrons. Because protons and neutrons have roughly the same mass (and electrons have almost no mass) the helium atom should have about four times the mass of the hydrogen atom. The mass of a neutron, $m_n = 1.67493 \times 10^{-27}$ kg, is about 1840 times the mass of electron.

Table 2 summarizes the nature and location of the sub-atomic particles

Table 2 Nature and location of sub-atomic particles

Particle	Location	Actual mass (g)	Relative mass(amu)	Charge
Proton	Nucleus	1.673×10^{-24}	1	+1
Electron	Shell	9.109×10^{-28}	0	-1
Neutron	Nucleus	1.675×10^{-24}	1	0

Bohr's model of an atom

Bohr suggested the following postulates to overcome the objections raised against Rutherford's model:

- Electrons revolve around the nucleus in stable orbits without emission of radiant energy. Each orbit has a definite energy and is called an energy shell or energy level.
- An orbit or energy level is designated as K, L, M, N shells. When the electron is in the lowest energy level, it is said to be in the ground state.
- An electron emits or absorbs energy when it jumps from one orbit or energy level to another.
- When it jumps from a higher energy level to lower energy level, it emits energy while it absorbs energy when it jumps from lower energy level to higher energy level

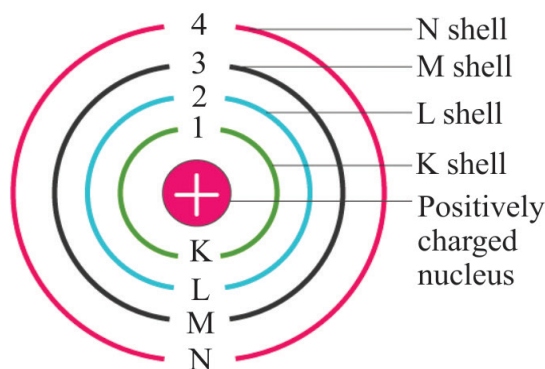


Figure 6. Bohr's Model of an Atom

Bohr's model gives an elaborative explanation on the structure of an atom and overcomes the objections faced by all the other models on the structure of an atom.

3.4 ATOMIC NUMBER, MASS NUMBER AND ISOTOPES

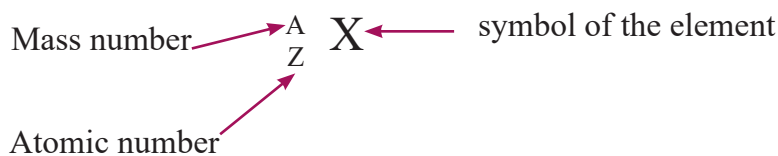
All atoms can be identified by the number of protons and neutrons they contain. The number of protons in the nucleus of each atom of an element is called the **atomic number (Z)**. In a neutral atom the number of protons is equal to the number of electrons, so the atomic number also indicates the number of electrons present in the atom. The chemical identity of an atom can be determined solely by its atomic number. For example, the atomic number of nitrogen is 7; this means that each neutral nitrogen atom has 7 protons and 7 electrons. Or viewed another way, every atom in the universe that contains 7 protons is correctly named “nitrogen.”

The **mass number (A)** is the total number of neutrons and protons present in the nucleus of an atom of an element. Except for the most common form of hydrogen, which has one proton and no neutrons, all atomic nuclei contain both protons and neutrons. In general, the mass number is given by

$$\begin{aligned} \text{Mass number} &= \text{number of protons} + \text{number of neutrons} \\ &= \text{atomic number} + \text{number of neutrons} \end{aligned}$$

The number of neutrons in an atom is equal to the difference between the mass number and the atomic number, or $(A - Z)$. For example, if the mass number of a particular boron atom is 12 and the atomic number is 5 (indicating 5 protons in the nucleus), then the number of neutrons is $12 - 5 = 7$. Note that all three quantities (atomic number, number of neutrons, and mass number) must be positive integers, or whole numbers.

The accepted way to denote the atomic number and mass number of an atom of element X is as follows:



For example, carbon, sodium and sulphur can be described as given below using their respective symbols and this notation.



Exercise

- The mass number of chlorine is 35 and its atomic number is 17. Give the number of its protons, electrons and neutrons
- Complete the following table

Table 3

Notation of the Element	Number of Protons	Number of Electrons	Number of Neutron	Z	A
${}^{24}_{12}\text{Mg}$	-	-	-	-	-
-	-	13	14	-	-
-	15	-	-	-	31
-	-	-	6	6	-

- Use a periodic table to tell the atomic number of the following elements.
 - Lithium
 - Boron
 - Fluorine
 - Calcium

Isotopes**ACTIVITY 4**

Form a group and discuss the following:

- Are neutrons present in all atoms?
- Can two atoms have the same number of electrons but different number of neutrons?
- Do the chemical properties of an atom depend on the number of electrons, protons or neutrons?
- Why the atomic masses for most elements are not whole numbers?
- Does an atom of zinc (Zn) have about the same mass as an atom of sulphur (S), about twice the mass, or about half the mass?

Atoms of the same element with the same number of protons but different number of neutrons are called **Isotopes**. Isotopes have the same atomic number but different mass numbers. For example, there are three isotopes of carbon. The first isotope contains 6 neutrons. The second contains 7 and the third contains 8 neutrons.

In the designation of isotopes, the mass number is written with a hyphen after the name of the element. For example, the isotopes of carbon with mass numbers 12, 13 and 14 can be written as Carbon-12, Carbon-13 and Carbon-14 respectively. Alternately they can be written as ${}^{12}\text{C}$, ${}^{13}\text{C}$ and ${}^{14}\text{C}$.

The three isotopes of hydrogen have the following common names:

Hydrogen-1	Protium
Hydrogen-2	Deuterium
Hydrogen-3	Tritium

The isotopes of an element are not found in nature in equal distribution. Among the isotopes of a given element, one of them will be found in greater quantity and others in smaller quantity. Among the three isotopes of carbon, for example, ^{12}C is the most abundant. Among the three isotopes of hydrogen, protium is the most common.

Note that since it is the electrons which are responsible for chemical properties, isotopes of an element have the same chemical properties. However, they differ in physical properties such as density.

3.5 RELATIVE ATOMIC MASS

The atomic mass of chlorine is 35.5 and that of copper is 63.5. Does it mean that atoms of these two elements have half a proton or half a neutron? Explain.

As discussed earlier, isotopes have the same number of protons but different number of neutrons. Boron for example, has two isotopes, namely, ^{10}B and ^{11}B . The mass of an atom is the total mass of neutrons and protons. Hence, ^{10}B has a mass of 10 amu and ^{11}B has a mass of 11 amu. However, when we determine the atomic mass of boron, the masses of the two isotopes should be considered.

Relative atomic mass is the average mass of the isotopes of a given element. For example, the two isotopes of chlorine exist in nature in different percentages: 75% of the atoms are ^{35}Cl and 25% are ^{37}Cl .

Thus the relative atomic mass of chlorine is calculated as the average mass of these two isotopes.

$$\begin{aligned} & [(75 \times 35)] + [(25 \times 37)] / 100 \\ & = (0.75 \times 35) + (0.25 \times 37) = 35.5 \text{ of chlorine} \end{aligned}$$

Example

There are two isotopes of lithium found on earth. Isotope ^6Li (6.01512 amu) accounts for 7.42% of the total, and isotope ^7Li (7.01600 amu) accounts for the remaining 92.58%. What is the relative atomic mass of lithium?

Solution:

$$7.42 \times \frac{6.01512}{100} + 92.58 \times \frac{7.01600}{100} = 6.942$$

Exercise

- Use the following table to answer the questions given below:

Table 4

Element	Number of electrons	Number of protons	Number of neutrons
A	5	5	6
B	19	19	20
C	5	5	5
D	35	35	46
E	35	35	44

- How many different elements indicated by the letters are listed in the Table
 - Which of them are isotopes of the same element?
- Hydrogen has 3 isotopes. These are hydrogen with a mass number 1, 2 and 3.
 - Write the isotopic notation for these isotopes.
 - How many protons, neutrons and electrons are present in these isotopes?
 - The atomic masses of the two stable isotopes of boron, ^{10}B (20%) and ^{11}B (80%) are 10 and 11 respectively. Calculate the average atomic mass of Boron. The percentages in parenthesis denote the relative abundance.
 - Carbon exists as the isotopes carbon-12, with a fractional abundance of 0.9890 and a mass of exactly 12 amu, and carbon-13, with a fractional abundance of 0.110 and a mass of 13.00335 amu. Calculate the average atomic mass of carbon.

3.6 QUANTUM NUMBERS

An atomic orbital is specified first by three quantum numbers that are associated respectively, with the orbital's size (energy), shape, orientation and, later, independent of these three quantum numbers, the electrons spin. The first three sets of quantum numbers have a hierarchical relationship: the size-related number limits the shape related number, the shape-related number in turn limits the orientation-related number.

Three among the four quantum numbers characterize the orbitals in that atom. That is, they describe the orbital or the space the electron is supposed to occupy. The fourth quantum number is used to describe the spin of the electron that occupy the orbital.

Principal quantum number (n)

This quantum number is the one on which the energy of an electron in an atom principally depends. It can have any positive value 1, 2, 3, and so on. The smaller n is, the lower the energy. In the case of the hydrogen atom or single-electron atomic

ions, such as Li^{2+} and He^+ , n is the only quantum number determining the energy. For other atoms, the energy also depends to a slight extent on the l quantum number. The size of an orbital also depends on n . The larger the value of n is, the larger the orbital. Orbitals of the same quantum state n are said to belong to the same shell. Shells are sometimes designated by the following letters:

Letter	K	L	M	N ...
n	1	2	3	4 ...

Angular momentum quantum number (l) (also called azimuthal quantum number)

This quantum number distinguishes orbitals of given n having different shapes; it can have any integer value from 0 to n . Within each shell of quantum number n , there are n different kinds of orbitals, each with a distinctive shape denoted by an l quantum number. For example, if an electron has a principal quantum number of 3, the possible values for l are 0, 1, and 2. Thus, within the M shell ($n = 3$), there are three kinds of orbitals, each having a different shape for the region where the electron is most likely to be found. Although the energy of an orbital is principally determined by the n quantum number, the energy also depends somewhat on the l quantum number, except for the H atom. For a given value of n , the energy of an orbital increases with l . Orbitals of the same value of n but different value of l are said to belong to different subshells of a given shell. The different subshells are usually denoted by letters as follows:

Letter	s	p	d	f	g	...
l	0	1	2	3	4	...

To denote a subshell within a particular shell, we write the value of the n quantum number for the shell, followed by the letter designation for the subshell. For example, $2p$ denotes a subshell with quantum numbers $n = 2$ and $l = 1$.

Magnetic Quantum Number (m_l)

This quantum number distinguishes orbitals of given n and l that is, of given energy and shape but having a different orientation in space; the allowed values are the integers from $-l$ to $+l$. For $l = 0$ (s -subshell), the allowed m_l quantum number is 0 only; there is only one orbital in the s subshell. For $l = 1$ (p -subshell), $m_l = -1, 0$, and $+1$; there are three different orbitals in the p subshell. The orbitals have the same shape but different orientations in space. In addition, all orbitals of a given subshell have the same energy. Note that there are $2l + 1$ orbitals in each subshell of quantum number l . The following table shows the permissible values of quantum numbers for atomic orbitals.

Table 5 Permissible Values of Quantum Numbers for Atomic Orbitals

n	l	m_l	Sub shell Notation	Number of orbitals in the sub shell
1	0	0	1s	1
2	0	0	2s	1
2	1	-1, 0, +1	2p	3
3	0	0	3s	1
3	1	-1, 0, +1	3p	3
3	2	-2, -1, 0, +1, +2	3d	5
4	0	0	4s	1
4	1	-1, 0, +1	4p	3
4	2	-2, -1, 0, +1, +2	4d	5
4	3	-3, -2, -1, 0, +1, +2, +3	4f	7

Spin quantum number (m_s)

This quantum number refers to the two possible orientations of the spin axis of an electron; possible values are $+1/2$ and $-1/2$. An electron acts as though it were spinning on its axis like the earth. Such an electron spin would give rise to a circulating electric charge that would generate a magnetic field. In this way, an electron behaves like a small bar magnet, with a north and a south pole.

The quantum numbers specify the energy states of the atom.

- The atom's energy levels or shells are given by the n value.
- The atom's sublevels or subshells are given by the n and l values. Each level contains sublevels that designate the shape of the orbital.
- The atom's orbitals are specified by the n , l and m_l values. Thus, the three quantum numbers that describe an orbital express its size (energy), shape and spatial orientation. Each sublevel is designated by a letter:

$l = 0$, is an s sublevel

$l = 1$, is a p sublevel

$l = 2$, is a d sublevel

$l = 3$, is a f sublevel

The letters s , p , d , and f are derived from the names of spectroscopic lines:

s , sharp;
 p , principal;
 d , diffuse;
 f , fundamental.

Sublevels are named by joining the n value and the letter designation. For example, the sublevel (subshell) with $n = 2, l = 0$ is called the $2s$ sublevel; the only orbital in this sublevel has $n = 2, l = 0$ and $m_l = 0$. A sublevel with $n = 3, l = 1$, is a $3p$ sublevel. It has three possible orbitals: one with $n = 3, l = 1$ and $m_l = -1$; another with $n = 3, l = 1$ and $m_l = 0$ and the third $n = 3, l = 1$, and $m_l = +1$.

For a given principal quantum number, n , the total number of orbitals is determined as: Number of orbitals = n^2 in a shell. Similarly, the number of orbitals in each subshell is determined as:

Number of orbitals in a subshell = $2l+1$.

General rules for assigning electrons to atomic orbitals.

Based on the preceding examples we can formulate some general rules for determining the maximum number of electrons that can be assigned to the various subshells and orbitals for a given value of n :

1. Each shell or principal level of quantum number n contains n subshells. For example, if $n = 2$, then there are two subshells (two values of l) of angular momentum quantum numbers 0 and 1.
2. Each subshell of quantum number l contains $(2l+1)$ orbitals. For example, if, $l = 1$, then there are three p orbitals.
3. No more than two electrons can be placed in each orbital. Therefore, the maximum number of electrons is simply twice the number of orbitals that are employed.
4. A quick way to determine the maximum number of electrons that an atom can have in a principal level n is to use the formula $2n^2$.

Example

1. What values of the angular momentum quantum number (l) and magnetic quantum number (m_l) are allowed for a principal quantum number (n) of 3? How many orbitals are allowed for $n = 3$?
2. Give the name, magnetic quantum numbers, and numbers of orbitals for each sublevel with the following quantum numbers:
 - (a) $n = 3, l = 2$
 - (b) $n = 5, l = 1$
 - (c) $n = 2, l = 0$
 - (d) $n = 4, l = 3$

3. What is wrong with each of the following quantum number designations and/or sub-level names?

Table 6

	n	l	m_l	Name
a	1	2	0	1p
b	4	3	+1	4d
c	3	1	-2	3p

Solution:

1. Determining l values:

For $n = 3$, $l = 0, 1, 2$

Determining m_l for each l value:

For $l = 0$, $m_l = 0$

For $l = 1$, $m_l = -1, 0, +1$

For $l = 2$, $m_l = -2, -1, 0, +1, +2$

Number of orbitals = n^2 , n is 3 and $n^2 = 3^2 = 9$ orbitals

3s : 1 orbital

3p : 3 orbitals

3d : 5 orbitals

Total = 9 orbitals

- 2.

Table 7

	n	l	Sub level	Possible m_l values	Number of orbitals
a	3	2	3d	-2, -1, 0, +1, +2	5
b	2	0	2s	0	1
c	5	1	5p	-1, 0, +1	3
d	4	3	4f	-3, -2, -1, 0, +1, +2, +3	7

3. A sublevel of $n = 1$ can have only $l = 0$, not $l = 1$. The only possible subshell is 1s.

- (a) A sublevel with $l = 3$ is an f sublevel, not a d sublevel. The sublevel name should be $4f$.
- (b) A sublevel with $l = 1$ can have only m_l of $-1, 0, +1$, not -2 .

Exercise

- Give the sublevel notation for each of the following sets of quantum numbers.
 - $n = 3, l = 2$
 - $n = 2, l = 0$
 - $n = 4, l = 1$
 - $n = 4, l = 3$
- Indicate whether each of the following is a permissible set of quantum numbers. If the set is not permissible, state why it is not.
 - $n = 3, l = 1, m_l = +2$
 - $n = 4, l = 3, m_l = -3$
 - $n = 3, l = 2, m_l = -2$
 - $n = 0, l = 0, m_l = 0$
 - $n = 3, l = 3, m_l = -3$
- Consider the electronic configuration of an atom:
 - What are the n, l and m_l quantum numbers corresponding to the $3s$ orbital?
 - List all the possible quantum number values for an orbital in $5f$ sub shell.
 - In which specific subshell will an electron be present if the quantum numbers $n = 3, l = 1$, and $m_l = -1$?
- Which of the quantum numbers relates to the electron only? Which relate (s) to the orbital?

3.7 ENERGY LEVELS AND ELECTRONIC CONFIGURATION

It is already stated that electrons revolve in shells around the nucleus. Electrons are not allowed to occupy all spaces around the nucleus. Only certain orbits are permitted for electrons to revolve around the nucleus. An orbit is the path that an electron takes around the nucleus of an atom. As electrons occupy farther orbits from the nucleus, their energy increases. Any electron on a higher orbit has a greater energy than an electron in a lower orbit. Therefore, the orbits are known as **energy levels**. Each shell represents an energy level. There are a number of shells in an atom.

Atomic shells are represented by letters. They are labeled as K, L, M, N, etc. They are also designated by the numbers 1, 2, 3, etc. The 1st shell is the K shell, the 2nd shell is the L shell, the 3rd shell is the M shell and so on.

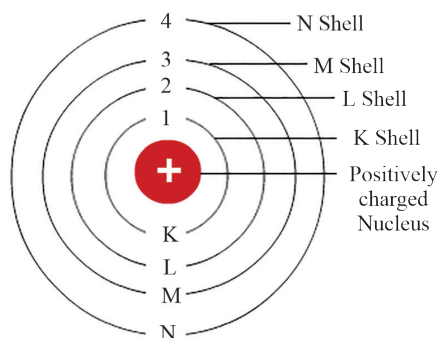


Figure 7. Atomic shells

Electronic configuration is the arrangement of electrons in the shells of atoms. Only a limited number of electrons can be placed in each shell. The maximum number of electrons that can be accommodated in a shell is based on the principal quantum number (n). It is represented by the formula $2n^2$, where ' n ' is the shell number. The shells, values of n , and the total number of electrons that can be accommodated are tabulated below:

Table 8 Maximum number of electrons in atomic shells

Atomic shell		Maximum number of electrons	
First shell (K)	$n = 1$	$n^2 = 2$	
Second shell (L)	$n = 2$	$n^2 = 8$	
Third shell (M)	$n = 3$	$n^2 = 18$	
Fourth shell (N)	$n = 4$	$n^2 = 32$	

The electronic configurations of the first 10 elements are given in Table 9.

Table 9 Electronic configuration of the first 10 elements

Element	Atomic number	Electronic configuration	
		K	L
H	1	1	
He	2	2	
Li	3	2	1
Be	4	2	2
B	5	2	3
C	6	2	4
N	7	2	5
O	8	2	6
F	9	2	7
Ne	10	2	8

ACTIVITY 5

Form a group and perform the following activity.

Complete the electronic configurations for the elements having atomic numbers 11-18 in the table given below:

Table 10

Element	Atomic Number	Electronic Configuration
Na	11	
Mg	12	
Al	13	
Si	14	
P	15	
S	16	
Cl	17	
Ar	18	

Note that the maximum number of electrons in the outer shell of an atom is 8, except helium which has 2 electrons. The maximum number of electrons that can be accommodated in a given shell is already given in Table 2. Regardless of this fact, the outer shell cannot accommodate more than 8 electrons. Therefore, elements with 8 electrons in their valence shells (including helium with 2 valence electrons) are said to have complete shell of electrons. The M-shell, for instance, can accommodate a maximum of 18 electrons. But when it is an outer shell as in 18Ar, the M-shell is complete or full with only 8 electrons. Elements that have complete shell of electrons, such as He, Ne and Ar are called the **noble gases** or **inert gases**.

Table 11 The noble gas configuration

Element	K	L	M	N	O
He	2				
Ne	2	8			
Ar	2	8	8		
Kr	22	8	18	8	
Xe	2	8	18	18	8

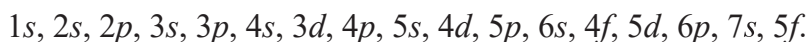
3.8 RULES AND PRINCIPLES FOR FILLING IN ELECTRONS.

Several questions arise when you look carefully at the electron configuration of an atom. To answer these questions, it is necessary to know the basic principles that govern the distribution of electrons among atomic orbitals. The electron configuration

for any atom follows the following three principles. Building up principle (Aufbau principle, Hund's principle and Paul's exclusion principle).

1. Building-up principle (Aufbau principle)

Aufbau principle is a scheme used to reproduce the electron configurations of the ground states of atoms by successively filling subshells with electrons in a specific order (the building-up order). Following this principle, you obtain the electron configuration of an atom by successively filling subshells in the following order:



This order reproduces the experimentally determined electron configurations (with some exceptions, which we will discuss later). You need not memorize this order. As you will see, you can very easily obtain it from the periodic table. Figure 8 shows the order of filling orbitals.

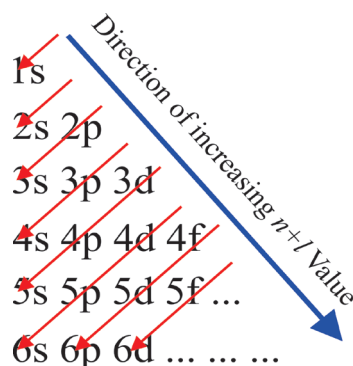


Figure 8. Order of filling Orbitals

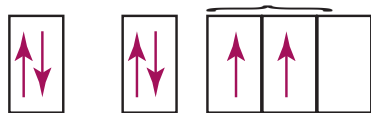
2. Hund's Principle.

In about 1927, Friedrich Hund discovered an empirical rule determining the lowest energy arrangement of electrons in a subshell. Hund's rule states that:

the lowest-energy arrangement of electrons in a subshell is obtained by putting electrons into separate orbitals of the subshell with the same spin before pairing electrons.

Let us see how this would apply to the carbon atom, whose ground-state configuration is $1s^2 2s^2 2p^2$. The first four electrons go into the $1s$ and $2s$ orbitals.

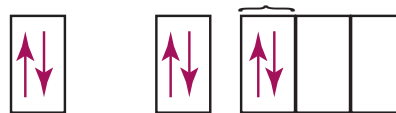
no electron-electron repulsion
equals lower energy



1s 2s 2p
correct

or

electron-electron repulsion
equals higher energy



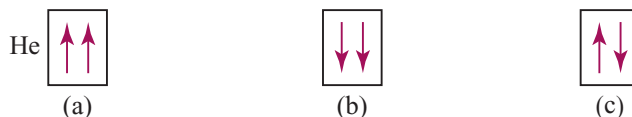
1s 2s 2p
incorrect

3. Pauli's Exclusion Principle.

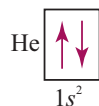
For many-electron atoms we use the Pauli exclusion principle (after the Austrian physicist Wolfgang Pauli) to determine electron configurations. This principle states that:

no two electrons in an atom can have the same set of four quantum numbers.

If two electrons in an atom should have the same n , and m , values (that is, these two electrons are in the same atomic orbital), then they must have different values of m_s . In other words, only two electrons may occupy the same atomic orbital, and these electrons must have opposite spins. Consider the helium atom, which has two electrons. The three possible ways of placing two electrons in the $1s$ orbital are as follows:



Diagrams (a) and (b) are ruled out by the Pauli exclusion principle. In (a), both electrons have the same upward spin and would have the quantum numbers $(1, 0, 0, +1/2)$; in (b), both electrons have downward spins and would have the quantum numbers $(1, 0, 0, -1/2)$. Only the configuration in (c) is physically acceptable, because one electron has the quantum numbers $(1, 0, 0, +1/2)$ and the other has $(1, 0, 0, -1/2)$. Thus, the helium atom has the following configuration:



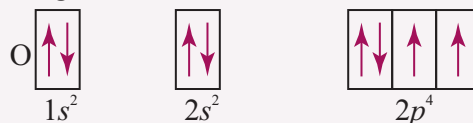
Example

An oxygen atom has a total of eight electrons. Write the four quantum numbers for each of the eight electrons in the ground state.

Strategy We start with $n = 1$ and proceed to fill orbitals in the order. For each value of n we determine the possible values of l . For each value of l , we assign the possible values of m_l . We can place electrons in the orbitals according to the Pauli exclusion principle and Hund's rule.

Solution

We start with $n = 1$, so, $l = 0$, a subshell corresponding to the $1s$ orbital. This orbital can accommodate a total of two electrons. Next, $n = 2$, and l may be either 0 or 1. The $l = 0$ subshell contains one $2s$ orbital, which can accommodate two electrons. The remaining four electrons are placed in the $l = 1$ subshell, which contains three $2p$ orbitals. The orbital diagram is:



ACTIVITY 6

Form a group and discuss the given questions. After discussion share your ideas with rest of the class.

1. Why are there never more than two electrons in an atomic orbital?
2. When there are two electrons in an orbital, why do these always have opposing spins?
3. Why are orbitals occupied singly first before the pairing of electrons occurs?
4. Why do the electrons in singly-occupied orbitals have parallel spins?

ACTIVITY 7

Form a group and complete the electron configuration for the period 3 elements of periodic table.

Table 12

Atomic number	Element	Outer shell Orbital diagram		Full electron configuration
		3s	3p	
11	Na	<input type="text"/>	<input type="text"/> <input type="text"/> <input type="text"/>	$1s^2 2s^2 2p^6 3s^1$
12	Mg	<input type="text"/>	<input type="text"/> <input type="text"/> <input type="text"/>	$1s^2 2s^2 2p^6 3s^2$
13	Al			
14	Si			
15	P			
16	S			
17	Cl			
18	Ar			

Electronic configuration using subshell notation

The subshell (sublevel) notation uses numbers to designate the principal energy levels or principal quantum number and the letters *s*, *p*, *d* and *f* to identify the sublevels or subshells. A superscript number following a letter indicates the number of electrons in the designated subshell (*e*). The designation can be written as nl^e . Where *n* represents the principal quantum number and *l* represents the subshell. The electron configurations for hydrogen (H; $Z = 1, 1s^1$), helium (He; $Z = 2, 1s^2$) and lithium (Li; $Z = 3, 1s^2, 2s^1$).

Example

Write down the electron configuration of each of the following atoms using subshell notation:

- (a) C ($Z = 6$)
- (b) N ($Z = 7$)
- (c) Ar ($Z = 10$)
- (d) P ($Z = 15$)
- (e) Si ($Z = 14$)
- (f) Fe ($Z = 26$)

Solution

Remember that in neutral atom atomic number (Z) is equal to the number of electrons in an atom. The principles of filling orbitals must also be remembered here.

- (a) $1s^2 2s^2 2p^2$
- (b) $1s^2 2s^2 2p^3$
- (c) $1s^2 2s^2 2p^6$
- (d) $1s^2 2s^2 2p^6 3s^2 3p^3$
- (e) $1s^2 2s^2 2p^6 3s^2 3p^2$
- (f) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

Electronic configuration using noble gas notation

The electronic configuration of sodium ($Z = 11$) is $1s^2 2s^2 2p^6 3s^1$. The first ten electrons of the sodium atom are the inner-shell electrons and the configuration of just those ten electrons is exactly the same as the configuration of the element neon ($Z = 10$). This provides the basis for a shorthand notation for electron configurations called the noble gas configuration. A noble gas configuration of an atom consists of the elemental symbol of the last noble gas prior to that atom, followed by the configuration of the remaining electrons. So for sodium, we make the substitution of [Ne] for the $1s^2 2s^2 2p^6$ part of the configuration. Sodium's noble gas configuration becomes [Ne] $3s^1$. Table 13 shows the electronic configuration of elements in the third period of the periodic table.

Table 13 The electronic configuration of elements in the third period of the periodic table.

Element	Symbol	Atomic number	Noble gas electron configuration
Sodium	Na	11	[Ne] $3s^1$
Magnesium	Mg	12	[Ne] $3s^2$

Aluminum	Al	13	$[\text{Ne}]3s^23p^1$
Silicon	Si	14	$[\text{Ne}]3s^23p^2$
Phosphorus	P	15	$[\text{Ne}]3s^23p^3$
Sulfur	S	16	$[\text{Ne}]3s^23p^4$
Chlorine	Cl	17	$[\text{Ne}]3s^23p^5$
Argon	Ar	18	$[\text{Ne}]3s^23p^6$

Exercise

Write down the electronic configuration of the following elements, using noble gas notation.

- Be ($Z = 4$)
- Fluorine ($Z = 9$)
- Potassium ($Z = 19$)
- Vanadium ($Z = 23$)
- Nickel ($Z = 28$)

KEY TERMS

- Atom
- Atomic Model
- Aufbau principle
- Average atomic mass
- Bohr's Model
- Cathode ray
- Dalton's atomic theory
- Dalton's Model
- Electron configuration
- Electron configuration
- Electron
- Isotopes
- Modern atomic theory
- Main energy level
- Neutron
- Nucleus
- Proton
- Quantum number
- Rutherford's Model
- Sub-energy level
- Thomson's Model

SUMMARY

- The idea that matter consists of discrete indivisible particles called “atoms” was introduced by Democritus (460-370 BC).
- Modern chemistry began with Dalton’s atomic theory, which states that all matter is composed of tiny, indivisible particles called atoms.
- According to the modern atomic theory, an atom consists of a very dense central nucleus containing protons and neutrons, with electrons moving around the nucleus at a relatively large distance from it.
- Protons are positively charged, neutrons have no charge, and electrons are negatively charged.
- A proton and a neutron have approximately the same mass; but the mass of an electron is negligible.
- Rutherford’s atomic model is that of a very small positively charged nucleus and extra-nuclear electrons.
- The nucleus consists of protons and neutrons and contains practically all the mass of an atom.
- Atomic masses and relative abundances of the isotopes of an element can be established by mass spectrometry.
- The atomic number of an element is the number of protons in the nucleus of an atom of the element.
- In an atom, the number of electrons equals the atomic number hence, an atom is electrically neutral.
- The mass number is the sum of the number of protons and the number of neutrons in the nucleus of an atom.
- An atom is represented by the notation, in which A_ZX is the symbol of an element Z is the atomic number, and A is the mass number.
- Isotopes are atoms of the same element with the same number of protons but different numbers of neutrons.
- The atomic mass of an element is the average mass of its isotopes.
- Electrons occupy certain energy levels outside the nucleus of an atom; these energy levels are also known as atomic shells.
- The maximum number of electrons that can occupy each of the atomic shells is specified.
- Electronic configuration is the arrangement of electrons in atomic energy levels.
- The n , l and m_l quantum numbers define an orbital, but a fourth quantum number is also required to characterize an electron in an orbital - the spin quantum number, m_s . This quantum number may have either of two values: $+\frac{1}{2}$ or $-\frac{1}{2}$.

- Aufbau Principle is a scheme used to reproduce the electron configurations of the ground states of atoms by successively filling subshells with electrons in a specific order (the building-up order).
- Hund's rule states that the lowest-energy arrangement of electrons in a subshell is obtained by putting electrons into separate orbitals of the subshell with the same spin before pairing electrons.
- Pauli's Exclusion principle states that no two electrons in an atom can have the same four quantum numbers.

Exercises

Part I. Choose the best answer from the given alternatives.

1. The idea that matter is continuous was accepted by the public until the time of
 - (a) Democritus
 - (b) Aristotle
 - (c) Dalton
 - (d) None
2. Which of the following is NOT true about Dalton's atomic theory?
 - (a) Matter consists of tiny particles called atoms.
 - (b) All atoms of an element may not be identical in mass.
 - (c) The atoms of different elements differ in mass and other properties.
 - (d) A chemical reaction involves separation, combination, or rearrangement of atoms.
3. Which of the following has the smallest mass?
 - (a) Proton
 - (a) Neutron
 - (b) Electron
 - (c) Nucleus
4. The sum of the number of protons and neutrons in an atom is known as
 - (a) Atomic number
 - (b) Atomic mass
 - (c) Mass number
 - (d) Number of electron
5. Which of the following atomic energy levels is nearest to the nucleus of an atom?
 - (a) K-shell
 - (b) L-shell
 - (c) M-shell
 - (d) N-shell

6. The atomic number of silicon is 14. How many electrons are contained in its second energy level?
- (a) 2
 - (b) 8
 - (c) 4
 - (d) 6
7. What is the number of valence electrons of ${}_{13}^{27}\text{Al}$?
- (a) 4
 - (b) 2
 - (c) 3
 - (d) 1
8. Suppose two atoms, ${}_{52}^{127}\text{X}$ and ${}_{53}^{127}\text{Y}$ are given. Which of the following is true about the two atoms?
- (a) X and Y are atoms of different elements.
 - (b) X and Y are atoms of the same element.
 - (c) X and Y are isotopes of the same element.
 - (d) X and Y have the same number of neutrons.
9. Which of the following are usually found in the nucleus of an atom?
- (a) Protons and neutrons only
 - (b) Protons, neutrons and electrons
 - (c) Neutrons only
 - (d) Electrons and neutrons only
10. An atom has an atomic number of 31 and a mass numbers of 70. How many electrons will it have in its valence shell?
- (a) 5
 - (b) 3
 - (c) 4
 - (d) 2
11. Which one of the following is not a valid electronic configuration?
- (a) 2, 8, 8, 2
 - (b) 2, 6
 - (c) 2, 8, 9, 1
 - (d) 2, 8, 4
12. Which of the given elements will have the electronic configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$?
- (a) Neon atoms
 - (b) Magnesium ion

- (c) Chlorine atoms
- (d) Calcium atoms

Part II. Fill in the blank spaces

13. The concept of the atom was first introduced by _____.
14. _____ and neutrons have approximately the same mass.
15. A certain isotope of P is represented by the symbol ${}_{15}^{31}\text{P}$. Thus, it has _____ protons _____ neutrons and _____ electrons.
16. The maximum number of electrons that can occupy the third shell of an atom is _____. Molecules of _____ consist of two or more different type of atoms.
17. Complete the following based on the information given.
 - (a) Cl_2 contains _____ chlorine atoms.
 - (b) H_2O contains _____ hydrogen and _____ oxygen.
 - (c) 3H_2 contains _____ hydrogen atoms.

Part III. Give short answer for the following questions

18. Identify the following subatomic particles:
 - (a) The number of these in the nucleus is equal to the atomic number.
 - (b) The particle that is gained or lost when ions are formed.
 - (c) The particle that is not found in the nucleus.
 - (d) The particle that has no electrical charge.
 - (e) The particle that has a much lower mass than the others
19. Complete the following table with appropriate answer

Element	Atomic number (Z)	Mass number(A)	Number of protons	Number of neutrons	Number of electrons
Neon	10	20			
Nitrogen	7	14			
Gallium	31	70			
Nickel	28	59			
Iron	26	56			

20. Carbon has atomic number 6. It comprises three isotopes, the first with 6 neutrons, the second with 7 neutrons, the third with 8 neutrons.
- Calculate the mass numbers of the three isotopes and represent them in the form of ${}^x_y\text{C}$.
 - Explain what is meant by “isotope”.
 - In naturally occurring copper isotopes, ${}^{63}_{29}\text{Cu}$ contributes 69.09% and ${}^{65}_{29}\text{Cu}$, 30.91%. Calculate the relative atomic mass of copper. (Accurate mass determined; ${}^{63}_{29}\text{Cu} = 62.9298$ amu, ${}^{65}_{29}\text{Cu} = 64.9278$ amu)

Chapter Outcomes

Upon completion of this chapter, students will:

- discuss the history and development of the Periodic Table;
- identify that elements are placed on the Periodic Table due to similar properties;
- identify the main blocks, groups and the periods of the Periodic Table;
- discuss the chemical and physical properties of the groups;
- discuss the periodic trends.

START UP ACTIVITY

Discuss the following in your group and present to the class.

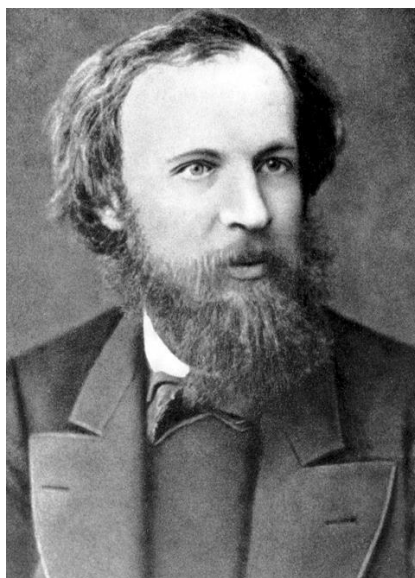
Consider days, months and seasons of a year.

1. How many days are there in a week according to the tradition in your locality?
2. How many days are there in a month?
3. Do you know seasons like summer, spring, autumn and winter. What similarity do you observe among months in a season?

4.1 HISTORY AND ORIGIN OF THE PERIODIC LAW**ACTIVITY 1**

1. What do we mean by the term periodic?
2. How do you describe the periodicity of days?
3. What other periodicities do you know?

In 1869 the Russian chemist Dmitri Mendeleev (1834–1907) and the German chemist J. Lothar Meyer (1830–1895), working independently, made similar discoveries. They found that when they arranged the elements in order of atomic mass, they could place them in horizontal rows, one row under the other, so that the elements in each vertical column have similar properties. Finally they conclude that “when the elements are arranged in order of increasing atomic mass, certain sets of properties recur periodically.” This was the first periodic law produced. A tabular arrangement of elements in rows and columns, highlighting the regular repetition of properties of the elements, is called a Periodic Table.



Dmitri Mendeleev (1834–1907)

Mendeleev had the idea of arranging ‘families’ of elements that have similar properties in vertical columns or groups. Mendeleev’s Periodic Table is given in Figure 1.

		I										
		H			III	IV	V	VI	VII			
0		1.01	II							VIII		
He 4.00	Li 6.94	Be 9.01	B 10.8	● C 12.0	N 14.0	O 16.0	F 19.0					
Ne 20.2	Na 23.0	Mg 24.3	Al 27.0	Si 28.1	P 31.0	● S 32.1	Cl 35.5					
Ar 40.0	K 39.1	Ca 40.1	Sc 45.0	Ti 47.9	V 50.9	Cr 52.0	Mn 54.9	● Fe 55.9	Co 58.9	Ni 58.7		
	● Cu 63.5	Zn 65.4	Ga 69.7	Ge 72.6	As 74.9	Se 79.0	Br 79.9					
Kr 83.8	Rb 85.5	Sr 87.6	Y 88.9	Zr 91.2	Nb 92.9	Mo 95.9	Tc (99)	Ru 101	Rh 103	Pd 106		
	● Ag 108	Cd 112	In 115	● Sn 119	Sb 122	Te 128	I 127					
Xe 131	Ce 133	Ba 137	● La 139	Hf 179	Ta 181	W 184	Re 180	Os 194	Ir 192	Pt 195		
	● Au 197	● Hg 201	Ti 204	● Pb 207	Bi 209	Po (210)	At (210)					
Rn (222)	Fr (223)	Ra (226)	● Ac (227)	● Th 232	● Pa (231)	● U 238						

Dobereiner's triads
 Known to Mendeleev

● Lanthanide series
● Actinide series
● Known to Ancients

Figure 1. Mendeleev's Periodic Table

Shortcomings of Mendeleev's Periodic Table

- (i) Wrong order of the atomic masses of some elements. Certain elements are arranged according to their increasing atomic masses, similarity in the chemical properties of the elements in a group is violated. For example, Ar with atomic mass of 39.9 comes first and K with atomic mass of 39.1 comes after it so that similarity of elements in a group is realized.
- (ii) The position of isotopes in the Periodic Table isotopes could not be given separate places in Mendeleev's Periodic Table. This observation led to the conclusion that atomic mass cannot be the basis of the classification of elements

4.2 STRUCTURE OF THE MODERN PERIODIC TABLE

In the years after Mendeleev, chemists made many discoveries that brought about changes in the Periodic Table. Consequently in 1913, an English physicist, Henry Mosley performed experiments that led to the discovery of a new property of

elements. This fundamental property of the elements is known as atomic number. No two elements can have the same atomic number. Therefore, atomic number of an element is a better basis for the classification of elements than atomic mass.

Modern periodic law states that “the properties of the elements are periodic functions of their atomic numbers”. In other words, if elements are arranged in the order of increasing atomic numbers, elements having similar properties will fall in the same column.

In the Modern Periodic Table, the elements are classified in periods and groups.

Periods: The horizontal rows of elements in the Periodic Table are called periods. There are seven periods in the modern Periodic Table. Each period contains a certain number of elements.

Period numbers are represented by Arabic numerals, 1 to 7.

Period 1 contains 2 elements, H and He
Period 2 contains 8 elements, Li to Ne
Period 3 contains 8 elements, Na to Ar
Period 4 contains 18 elements, K to Kr
Period 5 contains 18 elements, Rb to Xe
Period 6 contains 32 elements, Cs to Rn
Period 7 is incomplete, it starts with Fr.

In all the elements in a given period, electrons are filled in the same valence shell. For example, the first period consists of hydrogen and helium, in which the K-shell (1st shell) is occupied, in the second period from lithium to neon, the L shell (2nd shell) is occupied, in the third period from sodium to argon the M-shell (3rd shell) is occupied by the valence electrons.

Each period ends with elements whose valence shell is completely filled. Hence, the last element in a period contains the maximum number of electrons the valence shell can accommodate. Figure 3 shows the modern Periodic Table.

The number of elements in a period corresponds with the maximum number of electrons that can be placed in the valence shells of an atom. For example, the first period contains only 2 elements because the first shell can accommodate only 2 electrons. Period 2 contains 8 elements because the 2nd shell can have a maximum of 8 electrons, and so on.

Legend

Alkali Metals	Nonmetals
Alkaline Earth Metals	Noble Gases
Transition Metals	Other Metals
Rare Earth Metals	Metalloids

Relative Atomic Mass → **1.0079** Symbol → **H**
 Atomic Number → **1**

Period	Group	IA	IIA	IIIB	IVB	VB	VIB	VII B	VIII B	IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA		
1		1.0079 H 1															4.003 He 2		
2		6.941 Li 3	9.01 Be 4									10.81 B 5	12.01 C 6	14.01 N 7	15.999 O 8	18.998 F 9	20.18 Ne 10		
3		22.990 Na 11	24.31 Mg 12									26.98 Al 13	28.09 Si 14	30.97 P 15	32.06 S 16	35.45 Cl 17	39.95 Ar 18		
4		39.098 K 19	40.08 Ca 20	44.96 Sc 21	47.90 Ti 22	50.94 V 23	51.996 Cr 24	54.94 Mn 25	55.85 Fe 26	58.93 Co 27	58.70 Ni 28	63.55 Cu 29	65.37 Zn 30	69.72 Ga 31	72.59 Ge 32	74.92 As 33	78.96 Se 34	79.90 Br 35	83.80 Kr 36
5		85.458 Rb 37	87.62 Sr 38	88.91 Y 39	91.22 Zr 40	92.91 Nb 41	95.94 Mo 42	(98) Tc 43	(101.07) Ru 44	102.91 Rh 45	106.40 Pd 46	107.87 Ag 47	112.47 Cd 48	114.82 In 49	118.69 Sn 50	121.75 Sb 51	127.60 Te 52	126.90 I 53	131.30 Xe 54
6		132.91 Cs 55	137.33 Ba 56	136.91 La 57	178.49 Hf 72	180.95 Ta 73	183.85 W 74	186.21 Re 75	190.20 Os 76	192.22 Ir 77	195.09 Pt 78	196.97 Au 79	200.59 Hg 80	204.37 Tl 81	207.19 Pb 82	208.98 Bi 83	(209) Po 84	(210) At 85	(222) Rn 86
7		(223) Fr 87	226.03 Ra 88	227.03* Ac 89	(261) Rf 104	(262) Ha 105	(266) Sg 106	(262) Bh 107	(277) Hs 108	(268) Mt 109	(281) Uun 110	(272) Uun 111	(285) Uub 112	(284) Uut 113	(289) Uuq 114	(288) Uup 115			

* Lanthanides	40.12 Ce 58	140.91 Pr 59	144.24 Nd 60	(145) Pm 61	(145) Sm 62	151.96 Eu 63	157.25 Gd 64	158.93 Tb 65	162.50 Dy 66	164.93 Ho 67	167.26 Er 68	168.93 Tm 69	168.93 Yb 70	174.97 Lu 71
* Actinides	232.04 Th 90	231.04 Pa 91	238.03 U 92	237.05 Np 93	(244) Pu 94	(243) Am 95	(247) Cm 96	(247) Bk 97	(251) Cf 98	(252) Es 99	(257) Fm 100	(260) Md 101	(259) No 102	(262) Lr 103

Figure 2. The Modern Periodic Table

Groups: The vertical column of the elements in the Periodic Table are known as Groups or Families. In the modern Periodic Table, there are 18 vertical columns or groups. Each group is usually represented by the Roman numerals, I to VIII followed by the letters A or B as shown in Figure 4.

The “A” groups are designated with IA to VIIIA and referred to as the main groups or representative elements.

Some of the main groups of elements have common names or family names as given in Table 1.

Table 1 Common names of group ‘A’ element

Groups	Common names
IA	Alkali metals
IIA	Alkaline earth metals

IIIA	Boron family
IVA	Carbon family
VA	Nitrogen family
VIA	Oxygen family (Chalcogens)
VIIA	Halogens
VIIIA	Noble gases (Inert gases)

Legend:

- Alkali metals
- Alkaline-earth metals
- Transition metals
- Other metals
- Other nonmetals
- Halogens
- Noble gases
- Rare-earth elements (21, 39, 57–71) and lanthanoid elements (57–71 only)
- Actinoid elements

period	group	1*	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1	H																	He
2	1	Li	Be											B	C	N	O	F	Ne
3	1	Na	Mg											Al	Si	P	S	Cl	Ar
4	1	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	1	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	1	Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
7	1	Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og
lanthanoid series	6	58	59	60	61	62	63	64	65	66	67	68	69	70	71				
		Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu				
actinoid series	7	90	91	92	93	94	95	96	97	98	99	100	101	102	103				
		Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr				

Figure 3. Classification of the elements

Electronic configuration and arrangement of elements

Electronic configuration of the elements helps us to explain the repetition of properties. The position of elements in the Periodic Table can also be predicted from their electronic configurations.

The period number of an element corresponds to the number of shells in its atom. For example, if an element has 3 shells (K, L and M) then it belongs to the third period.

The number of shells is equal to the period number to which the element belongs. For example, if an element has two shells, it is found in period 2, an element with three shells belongs to the third period, and so on. See Table 2.

Table 2 Relationship between atomic number, electron configuration and number of shells of 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

Relationship between the number of shells and period number

- The group number of an element is equal to the number of its valence electrons. For example, if the valence electron of an element is 1, it belongs to groups IA, if it has 2 valence electrons, the element is found in group IIA, and so on.
- All elements in the same group have the same number of valence electrons. For example, Group IA elements, have 1 valence electron, Group IIA elements have 2 valence electrons, etc.

All the elements in a group have similar properties due to the same outer electronic configuration.

Electrons found in the outer most shell of an atom are called **valence electrons**.

Valence electrons are electrons that are involved in chemical combination. For example, all Group IA elements have one valence electron as shown in Table 3 and thus have similar chemical properties.

Table 3 Electronic configuration of Group IA elements

Group IA Elements	Atomic Number	Electronic configuration					No. of valence electrons
H	1	1					1
Li	3	2	1				1
Na	11	2	8	1			1
K	19	2	8	8	1		1
Rb	37	2	8	18	8	1	1

Classification of the elements

Elements are classified as metals, non-metals and metalloids.

- Metallic elements are found to the left and center of the Periodic Table.
- Non-metallic elements are found to the right of the Periodic Table.

All Group IA and IIA elements except hydrogen are metals.

All transition elements are also metals.

There are a few number of elements that show both metallic and non-metallic properties. These are called **semi-metals** or **metalloids**.

On the right hand side of the Periodic Table, there is a stair step line that separates metals and non-metals.

Elements that are found near the border line are metalloids.

Silicon, Boron and Germanium are some examples of metalloids.

Figure 4 shows the position of Metals, non-metals and metalloids in the Periodic Table.

																		Metal			Metalloid			Non-metal																							
H																							He																								
Li	Be													B	C	N	O	F	Ne																												
Na	Mg													Al	Si	P	S	Cl	Ar																												
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr																														
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe																														
Cs	Ba	La-Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn																														
Fr	Ra	Ac-Lr																																													
<table border="1" style="width: 100%; text-align: center;"> <tr> <td>La</td><td>Ce</td><td>Pr</td><td>Nd</td><td>Pm</td><td>Sm</td><td>Eu</td><td>Gd</td><td>Tb</td><td>Dy</td><td>Ho</td><td>Er</td><td>Tm</td><td>Yb</td><td>Lu</td> </tr> <tr> <td>Ac</td><td>Th</td><td>Pa</td><td>U</td><td>Np</td><td>Pu</td><td>Am</td><td>Cm</td><td>Bk</td><td>Cf</td><td>Es</td><td>Fm</td><td>Md</td><td>No</td><td>Lr</td> </tr> </table>																		La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr
La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu																																	
Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr																																	

Figure 4. Metals, non-metals and metalloids in the Periodic Table.

On the basis of electronic configuration or the type of atomic orbital which receives the last electron in atom, elements can be classified into three distinct categories. These are:

- Representative elements
- Transition elements
- Rare earth elements (inner transition elements)

Main group elements

These are elements having valence electrons entering the s or p orbital. The group in which they occur are usually designated as the “A” groups. Thus groups IA through group VIIIA elements are referred to as the representative elements. These elements are also called the s -block and the p -block elements. s -block elements are elements in which the last electron enters the s -orbital. p -block elements are elements in which the last electron occupies the p -orbital.

The transition elements

These are elements in which the last electron enters the d -sub level. They are also called d -block elements and are designated as the “B” groups. The transition elements are characterized by:

- Metallic nature such as metallic luster, thermal and electrical conductivity
- Variable valences
- Ability to form colored compounds
- Ability to form complex compounds

The rare earth metals

f -block elements or rare earth elements are those metals whereby the last electron goes to f -orbital. They are also called inner transition elements. There are two rows of these elements placed separately outside the main table. The first series of elements ($4f$) is known as the **lanthanides** and the second series ($5f$) is called **actinides**.

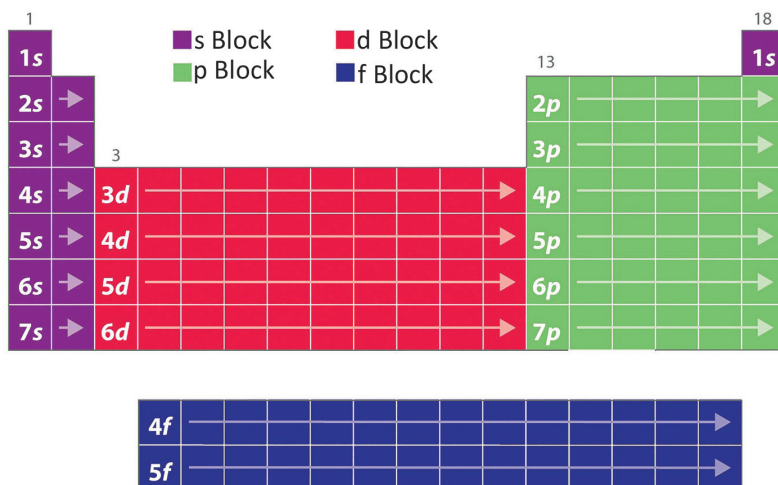


Figure 5. Classification of elements according to the type of subshells being filled.

Exercise

- Using Periodic Table, name the metals, non-metals and metalloids in group IVA.
- Name at least five common examples of transition metals.
- Element *Z* has an electronic configuration of 2,8,4. To which group and period does it belong?
- Determine the period and group for the elements with atomic numbers:
 - 40
 - 12
 - 9
 - 18
- An atom of a certain element has 14 electrons. Without consulting a Periodic Table, answer the following questions:
 - What is the ground-state electronic configuration of the element?
 - How should the element be classified?
- Group these electron configurations in pairs that would represent similar chemical properties of their atoms:
 - $1s^2 2s^2 2p^6 3s^2$
 - $1s^2 2s^2 2p^3$
 - $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
 - $1s^2 2s^2$
 - $1s^2 2s^2 2p^6$
 - $1s^2 2s^2 2p^6 3s^2 3p^3$

4.3 TRENDS IN PERIODIC PROPERTIES**ACTIVITY 2**

Form a group and discuss the following questions.

- How does size of elements vary across a period?
- Why are elements of Group 1 and 2 called metals, while those of Group 17 non-metals?
- Why elements of Group 18 are least reactive?
- Define shielding effect and effective nuclear charge.

As we have seen, the electron configurations of the elements show a periodic variation with increasing atomic number. Consequently, there are also periodic variations in physical and chemical behavior. In this section we will examine some physical properties of elements that are in the same group or period and additional properties that influence the chemical behavior of the elements.

A. Atomic size (Atomic radius)

ACTIVITY 3

Do you following activity individually and then share your answers with each other.

1. Plot a graph of atomic number of the second period elements vs ionic radius.
2. Plot another graph of atomic number vs ionic radius of the main group elements.

Use these graphs to explain the variation of ionic radius along a series and within a group of the Periodic Table.

The radius of an atom is an important factor in determining the ease with which the atom gain or lose electrons to form molecules or compounds. Atomic radius may be defined as a half distance between two chemically bonded atoms in a molecule Figure 6(a). It can also considered as the distance between the nucleus and the valence electrons of an atom Figure 6(b).

What is Atomic radius of an element?

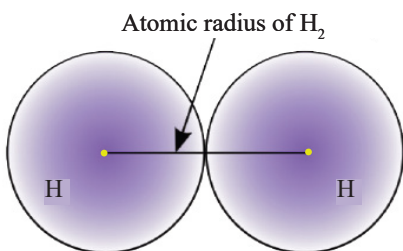
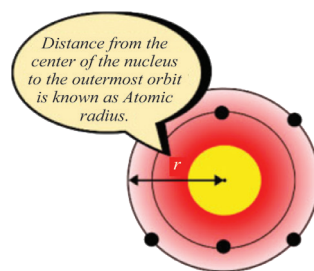


Figure 6. (a) Atomic radius of H_2 molecule



(b) Atomic radius between the nucleus and the valence electron

In going down a group, the number of shells increases and atomic radius also increases. Table 4 shows the increase in atomic radius of group IA elements down the group.

Table 4 Atomic radii and number of shells of Group IA elements

Group IA elements	Li	Na	K	Rb	Cs
Number of shells	2	3	4	5	6
Atomic radius Å	1.52	1.86	2.27	2.46	2.65

While moving across a period in the Periodic Table, there is a decrease in atomic size of the elements as electrons are added to the same energy level. Table 5 describes the trend of atomic radius across period 3 elements from left to right .

Table 5 Atomic radii of Period 3rd elements

Elements	Na	Mg	Al	Si	P	S	Cl	Ar
Atomic radius Å	1.86	1.60	1.43	1.17	1.10	1.04	0.99	0.88

Generally, in the Periodic Table, elements on the left side have larger atomic size, while elements on the right side have smaller atomic size.

B. Ionization Energy

ACTIVITY 4

Form a group and discuss the following questions:

Consider the following statement: “The ionization energy for the potassium atom is negative, because when K loses an electron to become K^+ , it achieves a noble gas electron configuration.” Indicate everything that is correct in this statement. Indicate everything that is incorrect. Correct the incorrect information and explain.

Present your findings to the class

The ionization energy (IE) is the amount of energy required to remove the outermost electronic in an isolated gaseous atom or ion.



The process of changing a neutral to an electrically charged species is called ionization. The minimum energy required to remove the first valence electron from an isolated neutral gaseous atom is called the first ionization energy. The second ionization energy removes the second electron and so on. Table 6 describes the first ionization energy of the alkali metals.

Table 6 The first ionization energies of the alkali metals

Group IA element	Li	Na	K	Rb	Cs
Ionization energy kJ/mol)	520	496	419	403	375

Factors affecting ionization energy

- **Atomic size** – with increasing atomic size, the attraction between the nucleus and the electrons is reduced. Thus, less energy is required to ionize the atom.
- **Effective nuclear charge** – the force of attraction of the nucleus for the electron is directly proportional to the nuclear charge. A larger effective nuclear charge makes removal of the electron more difficult.

- **Type of electron** – in a given energy level, *s*-electrons are closer to the nucleus than *p*-electrons. In turn, *p*-electrons are closer to the nucleus than the *d*-electrons and *d*-electrons are closer than *f*-electrons. Thus, their corresponding ionization energies decrease in that order.
- **Shielding effect or screening effect** of the inner electrons on the valence electrons. Electrons found in the lower energy levels will hinder the valence electrons from the attraction to the nucleus. This also affects the magnitude of ionization energy.
- **Effective nuclear Charge** – higher ionization energy means that it takes more energy to remove one electron from an atom ionization energy increases across a period because, going across a period effective nuclear charge increases.

As one goes down a group ionization energy decrease because of the increase in size of the atom. Metals have low ionization energies while non – metals have high ionization energies. Across any period, there is a general increase from left to right in the first ionization energy, corresponding to the related decrease in atomic radii.

C. Electron Affinity

Electron affinity is the energy liberated when an extra electron is added in isolated neutral gaseous atom to form a free ion. It is a measure of attraction between the nucleus and the extra electrons added to the atom. Electron affinity depends on the size of the atom and effective nuclear charge. The smaller the atomic size, the larger the effective nuclear charge and the higher the electron affinity.

What is the trend in electron affinity as one across a period from left to right?

D. Electronegativity

ACTIVITY 5

Form a group and discuss the following.

1. By using electron configuration, explain in why the electron affinity of F is negative value where as the electron affinity of Ne is a positive value.
2. Which group of elements have electron affinities with the largest negative values? Explain why?
3. Silicon has an electron affinity of -134kJ/mol . The electron affinity of phosphorus is -72kJ/mol .

Give a reason for this difference. Share your ideas with the rest of the class.

The term electronegativity was proposed by the American Chemist Linus Pauling in 1931. Electronegativity is the power of an atom to attract electrons towards itself when bonded to other atoms.

Atoms of the elements in the upper right of the Periodic Table (small, non-metal atoms) attract bonding electrons most strongly. Therefore, they have the greatest electronegativities.

Atoms of the elements to the left side of the table (large, metal atoms) have a weaker hold on electrons. They have the smallest electronegativities. On an electronegativity scale devised by Linus Pauling, the most non-metallic and hence most electronegative element, fluorine, is assigned a value of 4.0. Typical active metals have electronegativities of about 1.0 or less.

Within a group, electronegativity decreases from top to bottom. Chlorine is less electronegative than fluorine and sulphur is less electronegative than oxygen. The following figure shows the decrease in the electronegativity of group VIIIA elements down the group. See Figure 7.

Within a period of the Periodic Table, elements become more electronegative from left to right. In the second period, the trend is regular. Table 7 describes the increase in electronegativity of period 2 elements across a period.

Table 7 Electronegativities of period 2 elements

Period 2 elements	Li	Be	B	C	N	O	F
Electronegativity	1.0	1.5	2.0	2.5	3.0	3.5	4.0

E. Metallic Character

Metallic character refers to the chemical properties associated with elements classified as metals. These properties arise from the elements ability to lose electrons. As one moves across a period from left to right in the Periodic Table, the metallic character decreases, as atoms are more likely to gain electrons to fill their valance shell rather than to lose them to remove the shell. Down a group, the metallic character increases, due to the lesser attraction from the nucleus to the valence electrons.

F. Non-metallic Character

Non-metals tend to gain electrons in chemical reactions and have a high attraction for electrons within a compound. The most reactive non-metals reside in the upper

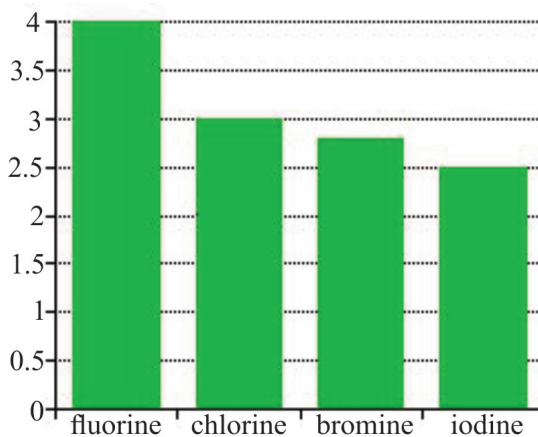


Figure 7. Electronegativity of group VIIIA elements down the group.

right portion of the Periodic Table. Since the noble gases are a special group because of their lack of reactivity, the element fluorine is the most reactive non-metal. It is not found in nature as a free element. Fluorine gas reacts explosively with many other elements and compounds and is considered to be one of the most dangerous known substances.

Non-metallic character increases from left to right across a period and decreases down a group.

G. Ionic and covalent bonding in compounds

A molecule or compound is made when two or more atoms form a chemical bond that links them together. There are two types of bonds:

- ionic bonds and
- covalent bonds

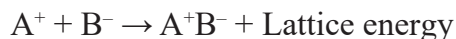
In an ionic bond, the atoms are bound together by the electrostatic forces of attraction between ions of opposite charge. Ionic bonds usually occur between metal and non-metal ions. For example, sodium (Na), a metal, and chloride (Cl), a non-metal, form an ionic bond to make NaCl. In a covalent bond, the atoms are bonded by sharing electrons.

Covalent bonds usually occur between non-metals. For example, in water (H₂O) each hydrogen (H) and oxygen (O) share a pair of electrons to make a molecule of two hydrogen atoms single bonded to a single oxygen atom.

In general, ionic bonds occur between elements that are far apart on the Periodic Table. Covalent bonds occur between elements that are close together on the Periodic Table. Ionic compounds tend to be brittle in their solid form and have very high melting points.

H. Lattice energy

Ionic compounds are more stable because of their electrostatic force between the two opposite ions. After the formation of ions, they combine together to form ionic compound. The energy released in this process is known as lattice energy or lattice enthalpy. That means, energy released when a cation and an anion combine together to form one mole of an ionic compound is known as lattice energy or lattice enthalpy. Thus, we can write



The strength of ionic bond increases with the increase of lattice energy. Lattice energy depends on two factors: size or radius of ions and charge of ions.

Figure 8 shows the formation of lattice.

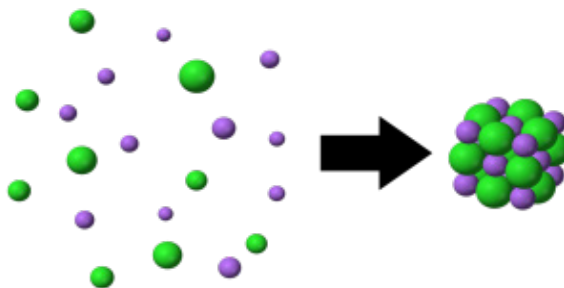


Figure 8. The formation of lattice

As the radius of ions increases, the lattice energy decreases. The following graph of the lattice energy of lithium halide (Figure 9). As the size of halide increases down the group, the lattice energy decreases. This is because with the increase of size of ions, the distance between their nuclei increases. Thus the attraction between them decreases and finally the less lattice energy released during the process.

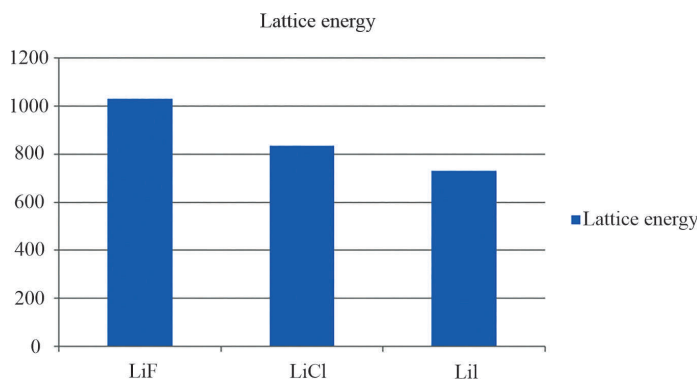


Figure 9. Lattice energy of LiF, LiCl and LiI

Lattice energy increases with the increase of charge on the ions because of their more attractive force between them. Thus +2 or -2 ions will release more lattice energy than the +1 or -1 ions. Figure 10 shows the comparison between the lattice energy of NaCl and MgO.

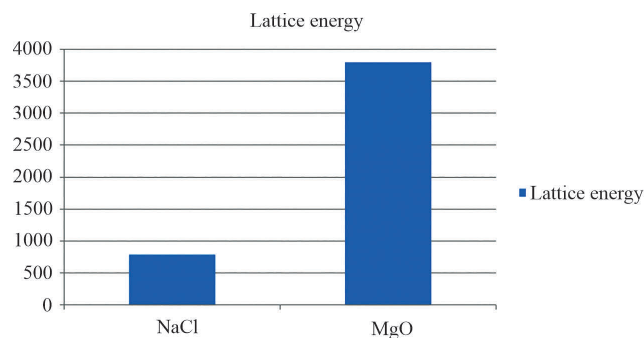


Figure 10. Lattice energy of NaCl and MgO

Here we can see that the lattice energy of MgO is much greater than the lattice energy of NaCl.

Exercise**Part I: Choose the correct answer from the given alternatives.**

- Which of the following properties of the elements remain unchanged down a group?
 - Ionization energy
 - Nuclear charge
 - Electron affinity
 - Valence electrons
- Which of the following elements has the largest atomic size?
 - Be
 - Ca
 - Ba
 - Mg
- Which of the following elements has the lowest electronegativity?
 - F
 - Br
 - I
 - Cl
- Which of the main groups in the Periodic Table has the elements with the most negative value of electron affinity?
 - Halogens
 - Noble gases
 - Alkali metals
 - None

Part II: Give short answers for the following:

- Compare the sizes of lithium (atomic number=3) and boron (atomic number=5). Which is one larger and which one is smaller?
- Identify the most metallic and the most non-metallic elements in period 3.
- Look at period 3 elements. Classify them as metals, metalloids and non-metals.
- Use the third period of the Periodic Table as an example to illustrate change in the first ionization energies of the elements as we move from left to right. Explain the trend.
- Based on their positions in the Periodic Table, predict which atom of the following pairs will have larger first ionization energy:

- (a) Ga, Ge
 - (b) Br, Sb
 - (c) K, Cr,
 - (d) Mg, Sr
 - (e) O, Ne
10. Define the following terms:
- (a) Atomic radius
 - (b) Nuclear charge
 - (c) Ionization energy
 - (d) Electron affinity
 - (e) Electronegativity
 - (f) Lattice energy
11. Fill in the following table regarding the trends in properties of elements across a period and down the group of the Periodic Table.

Table 8

Periodic properties	Across a period	Down the group	Reason
Atomic size			
Ionization energy			
Electronegativity			
Electron affinity			
Metallic character			

4.4 MAIN GROUP ELEMENTS

In chemistry, the main group is the group of elements (sometimes called the representative elements) whose lightest members are represented by helium, lithium, beryllium, boron, carbon, nitrogen, oxygen, and fluorine as arranged in the Periodic Table of the elements.

Helium (He)

Since its outermost electron orbital is full with two electrons, helium falls under inert gas. Also, helium can be found in compressed air tanks, lasers, and as a coolant in nuclear reactors. Helium holds the lowest melting and boiling points among all the other elements. The nuclear fusion of hydrogen in stars generates helium in a significant amount. Figure 11 shows the glowing nature of helium.

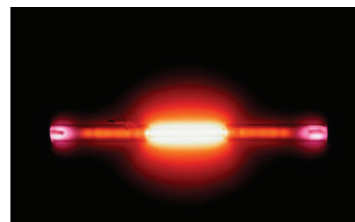


Figure 11. Helium

Table 9 Physical properties of helium

Phase	Gas, when cooled it condenses to liquid helium. Liquid helium is the only element that does not solidify at normal pressure regardless of the temperature
Density	Helium is less dense than any other known gas except hydrogen and is about one seventh as dense as air
Color	Colorless
Odor	It is odorless
Viscosity	Extremely low viscosity (resistance to flow - stickiness)
Taste	Tasteless gas
Compounds	Helium is completely inert. It does not form compounds or react with any other element

Chemical properties are simply the characteristics that define how the element reacts with other substances or changes from one to another substance. The chemical properties are seen only during a chemical reaction. The Chemical Properties of Helium are Tabulated in Table 10.

Table 10 Chemical properties of helium

Chemical symbol	He
Reactivity with water	It is less soluble in water than any other gas
Toxicity	Non Toxic
Flammability	It is non-flammable

ACTIVITY 6

Discuss in group on the following questions and present your idea to the rest of the class.

1. Where is the location of the main group elements in the Periodic Table?
2. What makes the main group elements different from other groups of the elements?
3. Are the main group elements metals or non-metals?

Lithium (Li)

Lithium is the first member of the alkali metal family. The alkali metals are the elements that make up Group 1 (IA) of the Periodic Table. The Periodic Table is a chart that shows how chemical elements are related to one another. The alkali metals include *sodium*, *potassium*, *rubidium*, *cesium*, and *francium*.

Lithium is the least dense of all metals. It has a density about half that of water.

Physical properties of lithium

Lithium is a very soft, silvery metal. It has a melting point of 180.54°C (356.97°F) and a boiling point of about $1,335^{\circ}\text{C}$ ($2,435^{\circ}\text{F}$). Its density is 0.534 grams per cubic centimeter. By comparison, the density of water is 1.000 grams per cubic centimeter. Lithium's hardness on the Mohs scale is 0.6. A hardness of 0.6 means that the material can be scratched with a fingernail.



Figure 12. Lithium

Chemical properties of lithium

Lithium is an active element, but not as active as the other alkali metals. It reacts slowly with water at room temperature and more rapidly at higher temperatures. It also reacts with most acids, giving off hydrogen gas. Lithium does not react with oxygen at room temperature, but above 100°C does so to form lithium oxide (Li_2O). Under the proper conditions, the element also combines with sulfur, hydrogen, nitrogen, and the halogens.

Beryllium (Be)

Beryllium is a chemical element with atomic number 4, represented by the chemical symbol 'Be'. It is a toxic, bivalent element having a steel gray appearance. It is a lightweight metal having very high melting point and good thermal conductivity.

It is non-magnetic by nature and exhibits good resistance to corrosion due to exposure to air and contact with nitric acid.



Figure 13. Beryllium

Beryllium belongs to the alkaline earth metal category, *s*-block and period 2 of the Periodic Table.

Beryllium is light, silver-gray, relatively soft metal that is strong but brittle. Beryllium has the highest melting point (1278°C) of the light metals.

Under normal conditions, a thin layer of the hard oxide BeO forms on beryllium's surface, protecting the metal from further attack by water or air. As a result of this, beryllium does not oxidize in air even at 600°C and it resists corrosion by concentrated nitric acid.

Beryllium has high thermal conductivity and is non-magnetic.

Boron (B)

Boron is the first element in the thirteenth column of the Periodic Table. The chemical symbol of Boron is B. The boron atom has five electrons and five protons. It is classified as a metalloid which means that it exhibits the properties of both metals and non-metals.

Physical Properties of Boron:

- Pure boron is a dark amorphous powder.
- Boron atoms can bond in a number of different types of crystal networks called allotropes.
- Crystalline boron is black in color and is extremely hard. The chemical compound boron nitride is the second hardest substance after diamond, which is an allotrope of carbon.
- It has a low density of 2.37 grams per cubic centimeters.
- It has a very high melting point of 2076°C because of the very strong crystal lattice.
- Boron salts emit a green color when introduced in Bunsen flame.

Chemical Properties of Boron:

- Boron in its crystalline form is very unreactive. Amorphous boron is reactive.
- Boron tends to make covalent bonds rather than ionic bonds.
- At room temperature, it is a poor electrical conductor, but it is a good conductor at high temperatures.
- Boron is not reactive when comes in contact with acids and alkalis at moderate temperatures.
- It reacts with metals to form borides

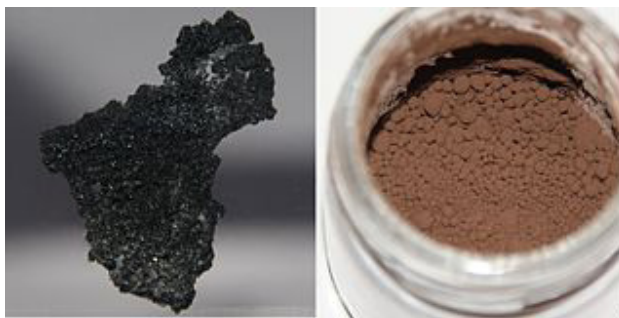


Figure 14. Left: Pure crystalline boron. Right: Amorphous brown boron as a powder in a glass jar

Carbon (C)

Carbon is a chemical element that is non-metallic. Its symbol is 'C'. Its atomic number is 6. It belongs to group 14 in the Periodic Table. Carbon is mainly found in coal deposits; however, the carbon obtained from coal deposits must be processed for its commercial use.

Carbon is one of the most significant elements on Earth. It is also a major constituent of most organic compounds. Since all the living bodies on Earth are made of mostly organic compounds, they (indirectly, carbon and its compounds) become important for our existence. The human body constitutes about 18% of carbon. Apart from this, carbon has its utility in various domains of technology, also such as in pieces of jewellery, paints, synthetic fibres, production of steel and carbon dating, etc.

The atoms of carbon can be bonded in different ways. They form allotropes of carbon. Allotropes are a different form of an element with a difference in physical properties but a similarity in chemical properties. It occurs in the same physical state in two or more crystalline forms. The most popular among the three allotropes of carbon are graphite and diamond. They both have different crystalline structures.

Graphite: It is black in color and opaque. It is a good conductor of electricity. It is a very good lubricant and the softest material known and greasy to touch. Certain forms of graphite are used in thermal insulation and certain others in thermal conduction. It burns in air at 700–800°C to give CO₂. It is insoluble in all ordinary solvents.

Diamond: It occurs naturally in a free state. It is a very stable allotrope of carbon made up of four covalent bonds around one carbon atom. It is a non-conductor of heat and electricity. Diamond is the hardest natural mineral found until now. It burns in the air at 900°C to give carbon dioxide. It is insoluble in all solvents.



Figure 15. Graphite



Figure 16. Diamond

Chemical properties of carbon

Chemical properties determine how carbon will react with other substances or change from one form to the other. The chemical properties of carbon are observed during the chemical reactions. Carbon forms millions of compounds.

Carbon takes part in four main reactions:

1. **Combustion Reaction:** When carbon burns in air, it gives carbon dioxide, heat, and light. Combustion is referred to as the process in which burning of carbon in excess of oxygen results in the production of heat and light.

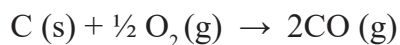


Unsaturated carbon burns with a yellow flame and produces soot while saturated carbon burns with a blue flame. Combustion could be either complete combustion or incomplete combustion.

In the complete combustion of a hydrocarbon, combustion takes place in excess of oxygen and the final products are carbon dioxide and water. Heat and light are generated in the form of energy. Saturated compounds undergo complete combustion.

On the contrary, incomplete combustion takes place when there is insufficient oxygen and there is an excess of hydrocarbon. It produces products which are carbon monoxide or carbon and water. It also has a smoky flame and produces soot. Unsaturated compounds undergo incomplete combustion.

2. **Oxidation Reaction:** Carbon and its compounds are oxidised in the presence of oxygen.



All combustion reactions are oxidation reactions but all oxidation reactions are not combustion reactions.

3. **Addition Reaction:** Carbon has the ability to make long strings or chains of atoms. This is referred to as the catenation.

Example: When ethene which contains a double bond is heated in the presence of hydrogen using a nickel catalyst, it produces ethane.



4. **Substitution Reaction:** The substitution reaction is a reaction in which a functional group in a compound is replaced by another functional group.



Nitrogen (N)

Nitrogen makes up the bulk of earth's atmosphere: 78.1 percent by volume. It is so inert at standard temperature and pressure that it was termed "azote" (meaning "without life") in Antoine Lavoisier's Method of Chemical Nomenclature. Nevertheless, nitrogen is a vital part of food and fertilizer production and a constituent of the DNA of all living things.

Physical properties of nitrogen

Nitrogen gas (chemical symbol N) is generally inert, non-metallic, colorless, odorless and tasteless. Its atomic number is 7, and it has an atomic weight of 14.0067. Nitrogen has a density of 1.251 grams/liter at °C and a specific gravity of 0.96737, making it slightly lighter than air. At a temperature of -210°C (63K) and a pressure of 12.6 kilo pascals, nitrogen reaches its triple point (the point an element can exist in gaseous, liquid and solid forms simultaneously).

At temperatures below nitrogen's boiling point of -195.79°C (77K), gaseous nitrogen condenses into liquid nitrogen, a fluid that resembles water and remains odorless and colorless. Nitrogen solidifies at a melting point of -210°C (63K) into a fluffy solid resembling snow.

Chemical properties of nitrogen

Elemental nitrogen is a diatomic gas (N₂). It forms a number of oxides (NO, N₂O, NO₂, N₂O₄, and N₂O₅), of which only N₂O₅ is a solid; the others are gases. Nitrogen has a tendency to accept three electrons to form the nitride ion, N³⁻ (thus, achieving the electron configuration 1s² 2s² 2p⁶, which is isoelectronic with neon). Most metallic nitrides (Li₃N and Mg₃N₂, for example) are ionic compounds

Oxygen (O)

Oxygen is by far the most abundant element in the earth's crust. Nearly half of all the atoms in the earth are oxygen atoms. Oxygen also makes up about one-fifth of the earth's atmosphere. Nearly 90 percent of the weight of the oceans is due to oxygen. In addition, oxygen is thought to be the third most abundant element in the universe and in the solar system.

Physical properties of oxygen

Oxygen is a colorless, odorless, tasteless gas. It changes from a gas to a liquid at a temperature of -182.96°C (-297.33°F). The liquid formed has a slightly bluish color to it. Liquid oxygen can then be solidified or frozen at a temperature of -218.4°C (-361.2°F). The density of oxygen is 1.429 grams per liter. By comparison, the density of air is about 1.29 grams per liter.

Oxygen exists in three allotropic forms. Allotropes are forms of an element with different physical and chemical properties. The three allotropes of oxygen are normal oxygen, or diatomic oxygen, or dioxygen; monatomic oxygen; and ozone, or triatomic oxygen.

Chemical properties of oxygen

Oxygen's most important chemical property is that it supports combustion. That is, it helps other objects to burn. The combustion (burning) of charcoal is an example. Charcoal is nearly pure carbon (C).

Oxygen also combines with elements at room temperature. Rusting is an example. Rusting is a process by which a metal combines with oxygen. When iron rusts, it combines with oxygen.

Oxygen also reacts with many compounds. Decay is an example. Decay is the process by which once-living material combines with oxygen. The products of decay are mainly carbon dioxide (CO_2) and water (H_2O).

Oxygen itself does not burn. A lighted match in a container of pure oxygen burns much brighter, but the oxygen does not catch fire.

Fluorine (F)

Fluorine is the lightest member of the halogen family, elements in Group 17 (VIIA) of the Periodic Table. These include chlorine, bromine, iodine, and astatine. Fluorine is the most active chemical element, reacting with virtually every element. It even reacts with the noble gases at high temperatures and pressures.

Fluorine was discovered in 1886 by French chemist Henri Moissan (1852-1907). Moissan collected the gas by passing an electric current through one of its compounds, hydrogen fluoride (HF).

Physical properties of fluorine

Fluorine is a pale yellow gas with a density of 1.695 grams per liter. That makes fluorine about 1.3 times as dense as air. Fluorine changes from a gas to a liquid at

a temperature of -188.13°C (-306.5°F) and from a liquid to a solid at -219.61°C (-363.30°F).

Fluorine has a strong and characteristic odor that can be detected in very small amounts, as low as 20 parts per billion. This property is very helpful to those who work with fluorine. It means that the gas can be detected and avoided in case it leaks into a room.

Chemical properties of fluorine

Fluorine is the most reactive element. It combines easily with every other element except helium, neon, and argon. It reacts with most compounds, often violently. For example, when mixed with water, it reacts explosively. For these reasons, it must be handled with extreme care in the laboratory.

Exercise

Chose the best answer from the given alternatives.

- Which of the following is NOT the member of the main group elements?
 - B
 - Be
 - Ca
 - Cr
- What makes helium falls under inert gases?
 - Because it contains eight electrons its outer most shell.
 - Because its outer most orbital is full.
 - Because it is insoluble in water.
 - Becasue it has only one orbital.
- Which of the following main group elements is metallic?
 - He
 - Be
 - N
 - F
- Which of the following is more inert than others?
 - C
 - F
 - N
 - O

5. Which of the following a good conductor of heat and electricity?
- Graphite
 - Diamond
 - Ozone
 - Flourine

4.5 COMPOUNDS OF PERIOD THREE ELEMENTS

The structure and physical properties of the elements :

- Sodium Na, magnesium Mg and aluminium Al are silvery solids, with a metal lattice structure, high boiling points and are good conductors of heat/electricity due to the delocalised free electrons moving between the immobile metal ions.
- Si has a non-metallic giant covalent structure based on a tetrahedral arrangement of Si-Si bonds and is a poor conductor of heat/electricity.
- Phosphorus P₄, sulfur S₈ and chlorine Cl₂ are simple-small covalent molecules and Ar consists of single atoms. The molecules are only held together by the weakest of the intermolecular forces, namely the instantaneous dipole - induced dipole forces, and consequently have very low melting/boiling points.

The hydrides of period 3 elements.

The hydrides formed by period 3 elements are NaH, MgH₂, AlH₃, SiH₄, PH₃, H₂S, HCl. These hydrides are formed by action of hydrogen on the elements. Table 11 shows the physical properties of period 3 hydrides.

Table 11 Physical properties of period 3 hydrides.

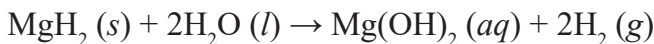
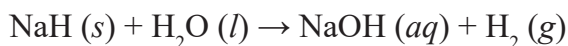
Hydrides	NaH	MgH ₂	AlH ₃	SiH ₄	PH ₃	H ₂ S	HCl
Bonding	Ionic	Ionic	Ionic	Covalent	Covalent	Covalent	Covalent
M.pt/°C	804				-134	-86	-144
B.pt/°C	1413				-88	-60	-85

NaH and MgH₂ have high melting and boiling points due to the strength of ionic bonds. H₂S and HCl have relatively higher melting and boiling points than SiH₄ and PH₃ because their molecules are polar molecules and are held together by hydrogen bonds whereas the molecules of SiH₄ and PH₃ are non-polar and therefore, held together by the weak van der Waals forces.

Chemical properties of period three hydrides

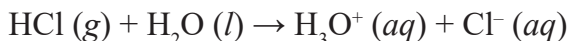
The reaction of hydrides of period 3 elements with water.

Due to the high polarity of the bonds in the hydrides NaH and MgH₂, they react with water readily producing hydrogen and the metal hydroxides.



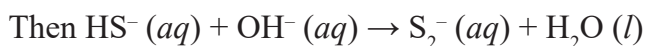
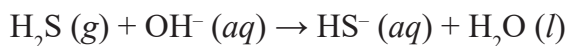
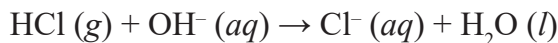
AlH₃, SiH₄, and PH₃ have no reaction with water due to lack of polarity in their bonds.

H₂S and HCl dissociate in water producing acids.



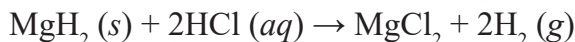
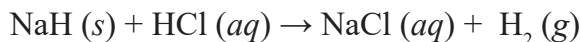
Reactions of hydrides of period 3 elements with hydroxides (NaOH)

HCl and H₂S react with hydroxides to form salts.



Reaction of period 3 element with HCl.

NaH and MgH₂ react to form hydrogen and metal salts.



Oxides of period three elements

They are formed by reacting the elements with oxygen. The following table describes the formula and physical properties of period three oxides

Table 12 The oxides of period 3 elements

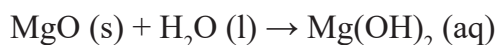
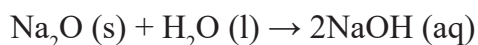
Element	Na	Mg	Al	Si	P	S	Cl
Oxide	Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₂ O ₅	SO ₃	Cl ₂ O ₇
Principal bonding	Ionic	Ionic	Ionic-covalent	Covalent	Covalent	Covalent	Covalent
Character	Basic	Basic	Amphoteric	Acidic	Acidic	Acidic	Acidic
M.pt/°C	1193	3075	2300	1728	563	30	-91

- Across the period there is a gradual change in the character of oxides from strongly basic to strongly acidic. Al₂O₃ is an amphoteric oxide, i.e., possesses both basic and acidic properties.
- Na₂O, MgO and Al₂O₃ have high melting points due to the strength of ionic bonding.
- The melting point of magnesium is higher than that of Na₂O because MgO, stronger electrostatic forces since Mg²⁺ has 2 positive charges, secondary, it has higher molecular mass.
- The melting point of Al₂O₃ is lower than that of MgO because Al₂O₃ has bigger covalent character due to high polarizing power of Al³⁺ ions.
- SiO₂, although covalent, has a high melting point because the crystal of SiO₂ is a giant structure in which silicon and oxygen atoms are bonded together by strong single covalent bonds.
- P₂O₅ (white solid) consists of P₄O₁₀ molecules; these contain electric dipoles due to electron displacements in their bonds and are quite strongly attracted to each other. This is why its melting point is high compared with that of P₄O₁₀ which consists of simple SO₃ molecules.
- Cl₂O₇ has a low melting point because the molecules are held together by the weak van der Waals forces.

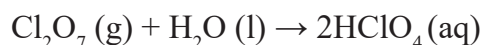
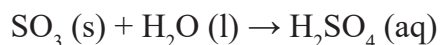
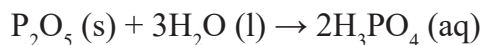
Chemical properties of oxides of period 3 elements

The reaction of oxides of period 3 elements with water.

Na₂O and MgO react with water to form hydroxides.



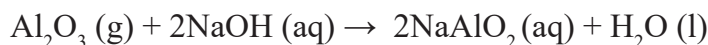
Al_2O_3 and SiO_2 have no reaction with water because they are insoluble oxides. P_2O_5 , SO_3 , Cl_2O_7 are acid anhydrides, i.e., react with water to form phosphoric, sulphuric and perchloric acids respectively.



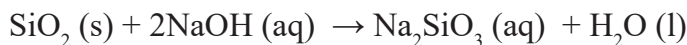
The reactions of period 3 oxides with NaOH.

Na_2O , MgO are basic, therefore, have no reaction with NaOH.

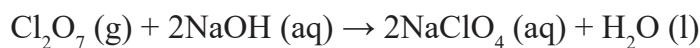
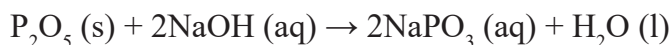
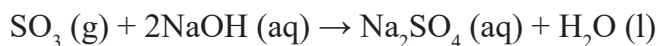
Al_2O_3 being amphoteric reacts with dilute NaOH to produce sodium aluminate and water.



SiO_2 reacts to form sodium silicate and water.

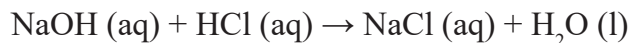


P_2O_5 , SO_3 and Cl_2O_7 react to form salts and water

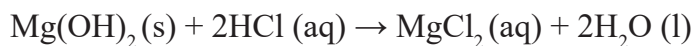


Hydroxides of Period 3 Elements

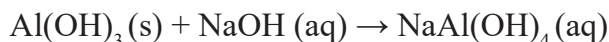
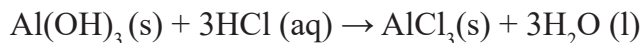
NaOH is a strong base and will react with acids to form a salt and water:



Mg(OH)_2 is also a basic compound which is often used in indigestion remedies by neutralizing the excess acid in the stomach to relieve pain:



$\text{Al}(\text{OH})_3$ is amphoteric and can act both as an acid and base:



Chlorides of period 3 elements

Physical properties of period 3 chlorides.

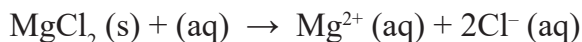
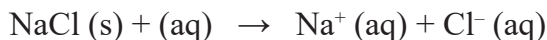
NaCl and MgCl_2 have high melting points due to the strength of the ionic bonding. AlCl_3 has a fairly high melting point because, in the solid-state, it consists of Al_2Cl_6 molecules and not simple AlCl_3 . These molecules are produced through dative bonding between Al and Cl in the Al_2Cl_6 molecules.

SiCl_4 and S_2Cl_2 are liquids and consist of respective simple SiCl_4 and S_2Cl_2 molecules and in the solid-state the molecules are held by weak van der Waals force which explains their very low melting points. PCl_5 a pale yellow solid has a fairly high melting point because the solid undergoes partial ionisation.



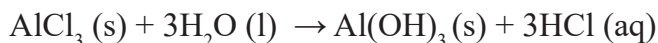
Chemical properties of period three chlorides

NaCl and MgCl_2 have no reaction with water but simply dissociate into ions.

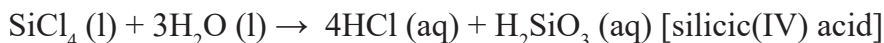
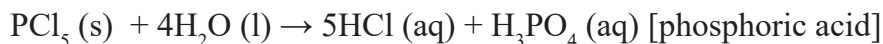


AlCl_3 , SiCl_4 , PCl_5 , Cl_2 are hydrolyzed by water. However, the extent of the hydrolysis of these chlorides varies across the period.

AlCl_3 is slightly hydrolyzed liberating HCl and $\text{Al}(\text{OH})_3$.



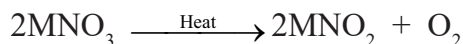
SiCl_4 , PCl_5 , and S_2Cl_2 are completely hydrolyzed liberating hydrogen chloride.



Hydrogen chloride under moist conditions appears as white fumes. For this reason, SiCl_4 , PCl_5 , S_2Cl_2 fume in moist air since their hydrolysis leads to formation of HCl.

Thermal stability of CO_3^{2-} , NO_3^- of Li, Na, K, Mg and Ca

Alkali and alkaline earth metal nitrates decompose on heating. On heating alkali metal (Na, K, Rb and Cs) decompose to form metal nitrites and oxygen.



Alkaline earth metal nitrates decompose on heating to give metal oxide, nitrogen dioxide and oxygen.

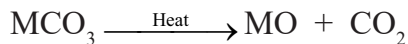


Lithium nitrate (like magnesium nitrate) decomposes to form metal oxide, nitrogen dioxide and oxygen.

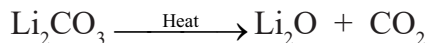


This is because small sized Li^+ ion cannot stabilize nitrate ion.

Alkaline earth metal carbonates have less stability towards heat and decompose to carbon dioxide.



On moving down the group, the stability of alkaline earth metal carbonates increases. Because of diagonal relationship, lithium carbonate decomposes similar to magnesium carbonate.



Other alkali metal carbonates being stable to heat do not decompose.

Experiment 1

Comparison of thermal stability of Na_2CO_3 , Li_2CO_3 and CuCO_3 .

Objective : To compare the reactivity of Na_2CO_3 , Li_2CO_3 and CuCO_3 .

Materials and apparatus : Eye protection, test tubes, delivery tube, spatula, Bunsen burner, clamp and stand.

Chemicals : Lime water (calcium hydroxide solution).

About 2 g each of following solids: copper carbonate, sodium carbonate, anhydrous and lithium carbonate.

Procedure

1. Put a large spatula measure of the carbonate to be tested in a test tube.
2. Fit a delivery tube and then clamp the test tube so that the delivery tube dips into a second test tube containing 2–3 cm³ lime-water.
3. Heat the solid gently at first, then more strongly.
4. Lift the delivery tube from the lime-water before, or as soon as, the heating is stopped. This is to avoid suck-back.
5. Write down all observations. Notice what happens to the limewater and how long it takes to turn milky (if at all). Notice whether any melting occurs in the heated test tube and any colour changes taking place. Write your results in tabular form.
6. Repeat the experiment with the other metal carbonates supplied, and in each case write down your observations.
7. Wash your hands thoroughly at the end of these experiments, since some of the metal carbonates are toxic.

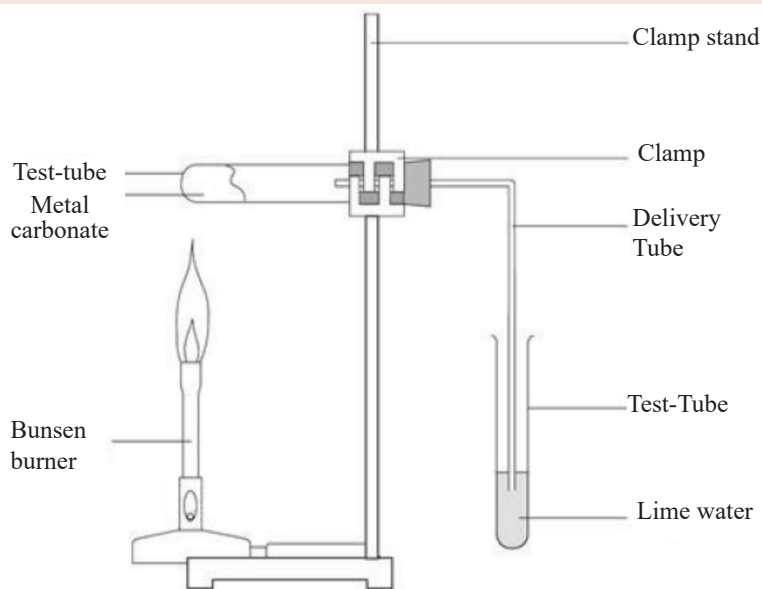


Figure 17. Thermal stability of metal carbonates

Table 13 Table of observation

Carbonate	Color before heating	Color after heating	Gas evolved	Ease of decomposition
Li_2CO_3	White			
Na_2CO_3	White			
CuCO_3	Green			

4.6 PERIOD 4 METALS POTASSIUM AND CALCIUM

Potassium (K, $Z = 19$).

Potassium is a soft, silvery metal that reacts extremely vigorously with water, and tarnishes rapidly in air. Its name is derived from the English word “potash,” for potassium carbonate, a compound found in high concentrations in wood ashes. The symbol “K” is derived from the Latin name for the element, kalium. Potassium is the eighth most abundant element in the earth’s crust (2.1%). The main ores in which potassium is found are sylvite [potassium chloride, KCl], carnallite [$\text{KMgCl}_3 \cdot 6\text{H}_2\text{O}$], and alunite [$\text{KAl}_3(\text{SO}_4)_2(\text{OH})_6$].

Potassium is essential for plant growth, and is heavily used in fertilizers. In the body, potassium plays a vital role in the contraction of muscle tissue; the movement of sodium and potassium ions in nerve cells plays a major role in the transmission of nerve impulses. When heated, potassium salts glow with a purple color, and are used in fireworks. Like sodium, metallic potassium is usually stored under mineral oil or some other hydrocarbon; it can also react with oxygen in dry air to produce potassium superoxide, KO_2 (see below).

Potassium undergoes a reaction with water similar to that of sodium; the products of the reaction are potassium hydroxide and hydrogen gas. This reaction releases a great deal of heat energy, often igniting the hydrogen gas that is produced.

Potassium-40, which accounts for 0.0117% of the world’s potassium, is radioactive, with a half-life of 1.25 billion years. It undergoes electron capture to produce argon-40; a comparison of the ratio of potassium-40 to argon-40 in rocks can be used to determine the age of the rock (potassium-argon dating). Trace amounts of potassium-40 are found in all sources of potassium; in a typical human, about 170,000 atoms of potassium-40 decay every second. The energy released by the decay of potassium-40 is partially responsible for the interior heat of the Earth, along with the decays of thorium and uranium.

There are a number of widely-used compounds of potassium. Potassium chloride, KCl, is used in salt substitutes (mixed with sodium chloride to improve its flavor), and in fertilizers; massive amounts of potassium chloride are used in lethal injections to cause rapid death by cardiac arrest. Potassium carbonate, K_2CO_3 , also known as potash, is used in the manufacture of glass. Potassium hydroxide, KOH, also known as caustic potash, is used in making soaps and detergents. Potassium nitrate, KNO_3 , also known as saltpeter, is a powerful oxidizer, and is one of the ingredients of gunpowder. Potassium chlorate, KClO_3 , is a very powerful oxidizer, and is used in

match heads and fireworks. Potassium superoxide, KO_2 , reacts with carbon dioxide to produce potassium carbonate and oxygen gas; it is used in rebreathers and respiration equipment to generate oxygen, and is also used in mines, submarines, and spacecraft.

Calcium (Ca, $Z = 20$).

Calcium is a silver-colored, relatively soft metal. The name of the element is derived from the Latin word for lime, *calx*. It is found in the Earth's crust at a concentration of 4.1%, making it the 5th most abundant element. The major sources of calcium are calcite and limestone [calcium carbonate, CaCO_3], anhydrite [calcium sulfate, CaSO_4], gypsum [calcium sulfate dihydrate, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$], and dolomite [a mixture of calcium and magnesium carbonate, $\text{CaMg}(\text{CO}_3)_2$]. See Figure 18.



Figure 18. (a) Calcite

(b) Gypsum

(c) Dolomite

Calcium salts form the hard parts of the bodies of most living creatures, from the shells of marine organisms and the coral of coral reefs (in the form of calcium carbonate, CaCO_3) to the bones and teeth of land-dwelling creatures (in the form of hydroxyapatite crystals, $[\text{Ca}_3(\text{PO}_4)_2]_3 \cdot \text{Ca}(\text{OH})_2$).

Since calcium forms such hard minerals, it is useful in building materials, such as plaster, mortar, and cement. Mortar is made from calcium oxide, CaO , also known as lime, or quicklime. When calcium oxide is treated with water it forms calcium hydroxide, $\text{Ca}(\text{OH})_2$, or slaked lime, which absorbs carbon dioxide from the air and gradually forms calcium carbonate, CaCO_3 . Lime, heated by hydrogen burning in oxygen, burns with a brilliant white light, which can be focused into a narrow beam visible over great distances. This kind of lighting was used in lighthouses, in surveying, and in theaters to produce spotlights (leaving the actor in the limelight).

Calcium chloride is a deliquescent (it absorbs enough water from the air that it dissolves in the solution), and is used to remove moisture from the air in damp basements.

“Hard water” contains dissolved minerals having +2 or +3 charges, such as calcium and magnesium; these salts cause some soaps and detergents to precipitate out as “soap scum”; these minerals precipitate out over time to form “scale” in water heaters and pots. The calcium can be removed by water softeners, which exchange the calcium ions for sodium ions, which have +1 charges, and do not readily precipitate out.

4.7 GROUP 7 ELEMENTS

The elements in group 7 are known as the halogens. These are fluorine, chlorine, bromine, iodine and astatine. These elements are non-metals that are poisonous. All halogens have similar reactions as they have seven electrons in their outermost shell. Halogens are diatomic, meaning they form molecules made of pairs of atoms sharing electrons (forming a single covalent bond between the two halogen atoms). When halogen atoms gain an electron during reactions, they form -1 ions called halide ions. Table 14 Shows the physical properties of the halogens.

Table 14 Appearance, Characteristics and Colour in Solution of the Halogens

Halogen	State & Appearance	Characteristics	Colour in solution
Fluorine	Yellow gas	Very reactive poisonous gas	-
Chlorine	Pale yellow-green gas	Reactive, poisonous and dense gas	Pale green
Bromine	Red-brown liquid	Dense red-brown volatile liquid	Orange
Iodine	Purple-black solid	Shimmery, crystalline solid, sublimates to form a purple vapour	Dark brown

The melting and boiling points of the halogens increase as you go down the group. This is due to increasing intermolecular forces as the atoms become larger, so more energy is required to overcome these forces. At room temperature (20°C), the physical state of the halogens changes as you go down the group. Fluorine and chlorine are gases, bromine is a liquid and iodine is crumbly solid. The colours of the halogens also change as you descend the group - they become darker.

Reactivity of the halogen

Reactivity of group 7 elements decreases as you go down the group. As you go down the group 7, the number of shells of electrons increases, the same as with

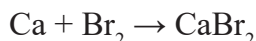
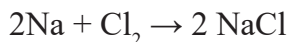
all other groups. However, halogen atoms form negative ions when they gain an electron to obtain a full outer shell. This means that the increased distance from the outer shell to the nucleus as you go down a group makes the halogens less reactive.

Fluorine is the smallest halogen, which means its outermost shell is the closest to the positive nucleus of all the halogen. Therefore, the ability to attract an electron is strongest in fluorine making it the most reactive. As you move down the group, the forces of attraction between the nucleus and the outermost shell decreases. This makes it harder for the atoms to gain electrons as you descend the group. Therefore, the halogens are less reactive the further down the group you go.

A **halogen displacement reaction** occurs when a more reactive halogen displaces a less reactive halogen from an aqueous solution of its halide. If you add chlorine solution to colourless potassium bromide or potassium iodide solution a displacement reaction occurs: The solution becomes orange as bromine is formed or the solution becomes brown as iodine is formed



The halogens react with some metals to form ionic compounds which are metal halide salts. The halide ion carries a -1 charge so the ionic compound formed will have different numbers of halogen atoms, depending on the valency of the metal



The halogens react with non-metals to form simple molecular covalent structures. For example, the halogens react with hydrogen to form hydrogen halides (e.g., hydrogen chloride)

4.8 ELEMENTS OF THE FIRST TRANSITION SERIES

This series contains elements from atomic number 21 (scandium) to atomic number 30 (Zinc). It is called 3d series because last electron goes in 3d orbital. It is present in fourth period of Periodic Table.

Scandium is a rare element that exists in compounds mainly in the +3 oxidation state— for example, in ScCl_3 , Sc_2O_3 , and $\text{Sc}_2(\text{SO}_4)_3$. The chemistry of scandium strongly resembles that of the lanthanides, with most of its compounds being colorless and diamagnetic. This is because the color and magnetism of transition

metal compounds usually arise from the *d*-electrons on the metal ion, and Sc^{3+} has no *d*-electrons. Scandium metal, which can be prepared by electrolysis of molten ScCl_3 , is not widely used because of its rarity, but it is found in some electronic devices, such as high-intensity lamps.

Titanium is widely distributed in the earth's crust (0.6% by mass). Because of its relatively low density and high strength, titanium is an excellent structural material, especially in jet engines, where light weight and stability at high temperatures are required. Nearly 5000 kg of titanium alloys is used in each engine of a Boeing 747 jetliner.

In addition, the resistance of titanium to chemical attack makes it a useful material for pipes, pumps, and reaction vessels in the chemical industry.

The most familiar compound of titanium is no doubt responsible for the white color of this paper. Titanium dioxide, or more correctly, titanium(IV) oxide (TiO_2), is a highly opaque substance used as the white pigment in paper, paint, linoleum, plastics, synthetic fibers, whitewall tires, and cosmetics (sunscreens, for example). Over one million tons is used annually in these and other products. Titanium(IV) oxide is widely dispersed in nature, but the main ores are rutile (impure TiO_2) and ilmenite (FeTiO_3).

Rutile is processed by treatment with chlorine to form volatile TiCl_4 , which is separated from the impurities and burned to form TiO_2 .

Vanadium is widely spread throughout the earth's crust (0.02% by mass). It is used mostly in alloys with other metals such as iron (80% of vanadium is used in steel) and titanium.

Vanadium(V) oxide (V_2O_5) is used as an industrial catalyst in the production of materials such as sulfuric acid.

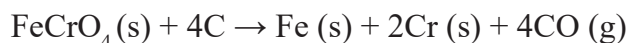
Pure vanadium can be obtained from the electrolytic reduction of fused salts, such as VCl_2 , to produce a metal similar to titanium that is steel gray, hard, and corrosion resistant. Often the pure element is not required for alloying. For example, ferrovanadium, (produced by the reduction of a mixture of V_2O_5 and Fe_2O_3 with aluminum), is added to iron to form vanadium steel, a hard steel used for engine parts and axles.

Vanadium commonly forms compounds in which it has the oxidation state +2, +3, +4 or +5, as shown in Table 4. The principal oxidation state of vanadium is +5, found in compounds such as the orange V_2O_5 (m.p. 65°C) and the colorless VF_5 (m.p. 19.5°C).

Table 15 Oxidation states and species for vanadium in aqueous solution

Oxidation state of vanadium	Species in aqueous solution
+5	VO_2^+ (yellow)
+4	VO^{2+} (blue)
+3	V^{3+} (aq) (blue-green)
+2	V^{2+} (aq) (violet)

Chromium: Although chromium is relatively rare, it is a very important industrial material. The chief ore of chromium is chromite (FeCr_2O_4), which can be reduced by carbon to give ferrochrome,



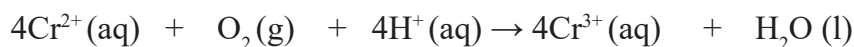
Which can be added directly to iron in the steel making process. Chromium metal, which is often used to plate steel, is hard and brittle and maintains a bright surface by developing a tough invisible oxide coating.

Chromium commonly forms compounds in which it has the oxidation state +2, +3, or +6, as shown in Table 16.

Table 16 Typical Chromium Compounds

Oxidation state of chromium	Examples of compounds (X- halogens)
+2	CrX_2
+3	CrX_3
	Cr_2O_3 (green)
	$\text{Cr}(\text{OH})_3$ (blue – green)
	$\text{K}_2\text{Cr}_2\text{O}_7$ (orange)
	Na_2CrO_4 (yellow)
+6	CrO_3 (red)

The Cr^{2+} (chromous) ion is a powerful reducing agent in aqueous solution.



The chromium(VI) species are excellent oxidizing agents, especially in acidic solution, where chromium(VI) as the dichromate ion ($\text{Cr}_2\text{O}_7^{2-}$) is reduced to the Cr^{3+} ion:



Manganese is relatively abundant (0.1% of the earth's crust). The most common use of manganese is in the production of an especially hard steel used for rock crushers, bank vaults, and armor plate.

One interesting source of manganese is from manganese nodules found on the ocean floor. These roughly spherical “rocks” contain mixtures of manganese and iron oxides as well as smaller amounts of other metals such as cobalt, nickel, and copper. Apparently, the nodules were formed at least partly by the action of marine organisms. Because of the abundance of these nodules, there is much interest in developing economical methods for their recovery and processing.

Manganese can exist in all oxidation states from +2 to +7, although +2 and +7 are the most common.

Manganese(II) forms an extensive series of salts with all the common anions. Table 17 shows some compounds of Manganese in its most common oxidation states.

Table 17 Some compounds of manganese in its most common oxidation states.

Oxidation state of manganese	Examples of compounds
+2	Mn(OH) ₂ (pink), MnS (salmon), MnSO ₄ (reddish), MnCl ₂ (pink)
+4	MnO ₂ (dark brown)
+7	KMnO ₄ (purple)

Iron is the most abundant heavy metal (4.7% of the earth's crust) and the most important to our civilization. It is a white, lustrous, not particularly hard metal that is very reactive toward oxidizing agents. For example, in moist air it is rapidly oxidized by oxygen to form rust, a mixture of iron oxides.

The chemistry of iron mainly involves its +2 and +3 oxidation states. Typical compounds are shown in Table 4.

In aqueous solutions iron(II) salts are generally light green because of the presence of $\text{Fe}(\text{H}_2\text{O})_6^{2+}$.

Although the $\text{Fe}(\text{H}_2\text{O})_6^{3+}$ ion is colorless, aqueous solutions of iron(III) salts are usually yellow to brown in color due to the presence of $\text{Fe}(\text{OH})(\text{H}_2\text{O})_5^{2+}$, which results from the acidity of $\text{Fe}(\text{H}_2\text{O})_6^{3+}$.

Table 18 Typical compounds of iron

Oxidation state of Iron	Examples of compounds
+2	FeO (black) FeS (brownish black) FeSO ₄ · 7H ₂ O (green)
+3	K ₄ Fe(CN) ₆ (yellow) FeCl ₃ (brownish black), Fe ₂ O ₃ (reddish brown), K ₃ Fe(CN) ₆ (red), Fe(SCN) ₃ (red)
+2, +3 (mixtures)	Fe ₃ O ₄ (black), KFe[Fe(CN) ₆] (deep blue, “Prussian blue”)

Although cobalt is relatively rare, it is found in ores such as smaltite (CoAs₂) and cobaltite (CoAs) in large enough concentrations to make its production economically feasible.

Cobalt is a hard, bluish white metal mainly used in alloys such as stainless steel and stellite, an alloy of iron, copper, and tungsten that is used in surgical instruments.

The chemistry of cobalt involves mainly its +2 and +3 oxidation states, although compounds containing cobalt in the 0, 1, or 4 oxidation state are known.

Nickel, which ranks twenty-fourth in elemental abundance in the earth’s crust, is found in ores, where it is combined mainly with arsenic, antimony, and sulfur.

Nickel metal, a silvery white substance with high electrical and thermal conductivities, is quite resistant to corrosion and is often used for plating more active metals.

Nickel is also widely used in the production of alloys such as steel.

Nickel in compounds is almost exclusively in the +2 oxidation state.

Copper, widely distributed in nature in ores containing sulfides, arsenides, chlorides, and carbonates, is valued for its high electrical conductivity and its resistance to corrosion.

It is widely used for plumbing, and 50% of all copper produced annually is used for electrical applications.

Copper is a major constituent in several well-known alloys (see Table 19).

Table 19 Alloys containing copper

Alloy	Composition %by mass
Brass	Cu (20–97), Zn (2–80), Sn (0–14), Pb (0–12), Mn (0–25)
Bronze	Cu (50–98), Sn (0–35), Zn (0–29), Pb (0–50), P (0–3)
Sterling silver	Cu (7.5), Ag (92.5)
Gold (18-karat)	Cu (5–15), Au (75), Ag (10–20)
Gold (14-karat)	Cu (12–28), Au (58), Ag (4–30)

The chemistry of copper principally involves the +2 oxidation state, but many compounds containing copper(I) are also known. Aqueous solutions of copper(II) salts are a characteristic bright blue color due to the presence of the $\text{Cu}(\text{H}_2\text{O})_6^{2+}$ ion. Table 20 lists some typical copper compounds.

Although trace amounts of copper are essential for life, copper in large amounts is quite toxic; copper salts are used to kill bacteria, fungi, and algae. For example, paints containing copper are used on ship hulls to prevent fouling by marine organisms.

Table 20 Typical Compounds of Copper

Oxidation State of Copper	Examples of Copper d
+1	Cu_2O (red) Cu_2S (black) CuCl (white)
+2	CuO (black) $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ (blue) $\text{CuCl}_2 \cdot 2\text{H}_2\text{O}$ (green), $[\text{Cu}(\text{H}_2\text{O})_6](\text{NO}_3)_2$ (blue)

Magnetic properties of transition metals

There are various substances which show magnetic behaviour. We have substances that are attracted by the magnetic field and are called paramagnetic. This phenomenon is called *paramagnetism*. Paramagnetic property is only shown when the substance contains one or more unpaired electrons. When a substance acquires a permanent magnetic moment, it is known as ferromagnetic and the phenomenon is called *ferromagnetism*. On the other hand, we also have substances which are repelled

by magnetic field and are called as diamagnetic substances. A substance shows *diamagnetism* when it contains only paired electrons.

Most of the transition elements show paramagnetic behaviour. The unpaired electrons in $(n-1)$ *d*-orbitals are responsible for the *magnetic properties*. The paramagnetic character of the transition metals increases on moving from left to right as the number of unpaired electron increases from one to five. The middle elements are found to possess the maximum paramagnetic property. The magnetic properties decrease with the decrease in the number of unpaired electrons. The transition metals which contain paired electrons depict diamagnetic behaviour.

KEY TERMS

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

SUMMARY

- Mendeleev's periodic law states that the properties of elements are periodic functions of their atomic masses.
- Group is the vertical column of elements.
- Period is the horizontal row of elements.
- The modern periodic law states that the properties of elements are periodic functions of their atomic numbers.
- The Periodic Table is the arrangement of the elements into periods and groups.
- The Periodic Table consists of 8 main groups and 7 periods.
- The electronic configurations of elements helps to determine the group and period number of an element.
- All elements in the same group have the same number of valence electrons.
- The number of valence electrons determines the group number of an element in main group elements.
- The number of shells of an atom equals the period number to which the element belongs.
- Elements in the same group have similar chemical properties because they have the same valence electronic configuration (valence electrons).
- Elements are classified as metals, non-metals and metalloids.
- Metals are found on the left side while non-metals are found on the right side of the Periodic Table.
- Certain atomic properties vary periodically, when atoms are considered in terms of increasing atomic number. The properties and trends considered in this unit are those of atomic radius, ionic radius, ionization energy and electron affinity. Values of these atomic properties strongly influence physical and chemical properties of the elements.

Exercises**Part I. Write true for correct statement and false for wrong statement.**

1. Mendeleev classified elements in the order of increasing atomic number.
2. Elements in the same family have the same number of shells.
3. The atomic numbers of elements were assigned by Mosley.

4. Elements in a given period have different properties because they have different valence electrons.
5. In the Periodic Table, most elements are metals.
6. An element with atomic number 15 is found in group IVA and period.
7. Atomic size increases as we go down a group and decreases across a period.
8. Halogens are the most electropositive elements.
9. Compounds of most transition elements form colored ions in aqueous solutions.
10. The most common oxidation states of manganese are +2 and +7.

Part II. Multiple Choice Type Questions

11. An element with atomic number 8 is found in _____.
 - (a) Group IIA and period 8
 - (b) Group VIA and period 2
 - (c) Group IIA and period 3
 - (d) Group VIIIA and period 1
12. All elements in the same period have the same _____.
 - (a) atomic number
 - (b) physical properties physical
 - (c) number of shells
 - (d) number of valence electrons
13. How many elements are there in the fourth period of the Periodic Table?
 - (a) 8 (b) 18 (c) 4 (d) 32
14. Across a period, atomic size of the elements _____.
 - (a) decreases
 - (b) increases
 - (c) increases and decreases
 - (d) remains the same
15. The first ionization energy of aluminium is slightly lower than that of magnesium because:
 - (a) magnesium has a higher nuclear charge.
 - (b) the outer electron in aluminium is in a *p*-orbital not an *s*-orbital.
 - (c) in aluminium the electron is being lost from a doubly filled orbital.
 - (d) the radius of the aluminum atom is greater than that of the magnesium atom.
16. Which of the following the member of elements of the first transition series?
 - (a) Na (b) Cr (c) K (d) Ca

17. Which of the following is the main group element?
 - (a) Fe
 - (b) C
 - (c) V
 - (d) Zn
18. The ability to be attracted towards a magnet is referred to as
 - (a) ferromagnetism
 - (b) diamagnetism
 - (c) paramagnetism
 - (d) all of these
19. If the electronic configuration of element 'X' is 2,8,8,1, the element is
 - (a) a non-metal
 - (b) an alkali metal
 - (c) found in group IVA and period
 - (d) an alkaline earth metal
20. The common name of Group VIIA elements is _____.
 - (a) noble gases
 - (b) chalcogens
 - (c) halogens
 - (d) alkaline earth

Part III. Give short answer to the following questions

21. Why is argon (atomic mass 40) placed before potassium (atomic mass 39) in the modern Periodic Table?
22. An element has mass number of 23 and its nucleus has 12 neutrons. To which group of the Periodic Table does it belong? Explain.
23. Why do atomic sizes of the elements increase from top to bottom in a given group?
24. If the atomic number of an element X is 17, deduce the following from the given information.
 - (a) Number of protons
 - (b) Number of electrons
 - (c) Electronic configuration
 - (d) Number of valence electrons
 - (e) Number of shells
 - (f) The period and group number
25. Arrange the following in order of increasing ionization energy: Li, Na, Ne, N, O.



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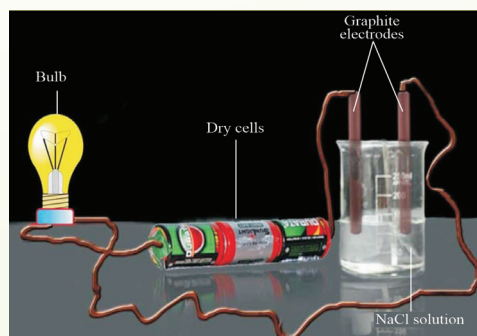
CHAPTER

5

CHEMICAL BONDING

Chapter Contents

- 5.1 Introduction
- 5.2 Types of Chemical Bonding
- 5.3 Hybridization of Atomic Orbitals
- 5.4 Simple molecules and their Shapes
- 5.5 Molecular Geometry
- 5.6 Metallic Bonding
- 5.7 Intermolecular forces
- 5.8 Coordinate Covalent Bond
 - Key Terms
 - Summary
 - Exercises



Chapter Outcomes

Upon completion of this topic, students will:

- distinguish the different types of chemical bonding;
- discuss hybridization of atomic orbitals;
- demonstrate the types of bonding of molecules by using the Lewis Structure;
- discuss electronegativity relative to the concept of bond polarity;
- discuss the molecular geometry;
- distinguish between interatomic bonding and intermolecular bonding and
- explain coordinate covalent (dative) bond.

Start-up Activity

Form a group and discuss the following questions.

1. Why do atoms readily combine to form molecules?
2. Why molecules are more stable than free atoms?
3. What keeps the atoms together in a molecule?
4. Why do elements combine in certain fixed ratio?

5.1 INTRODUCTION

Almost everything a person sees or touches in daily life, like the air we breathe, the food we eat, the clothes we wear, are the result of chemical bonds. The concept of chemical bonding lies at the very core of chemistry; it is what enables about little over one hundred elements to form millions of known chemical substances that make up our physical world.

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.

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

In this chapter we will focus on three types of bonding, namely ionic bonding, covalent bonding and metallic bonding.

Octet Rule

You have studied in chapter 4 that noble gases have very stable electron arrangements such as 2; 2, 8; 2, 8, 8 and their outer shells are fully saturated. The first three are shown in Figure 1 and explains why noble gases are so reluctant to form compounds with other elements.

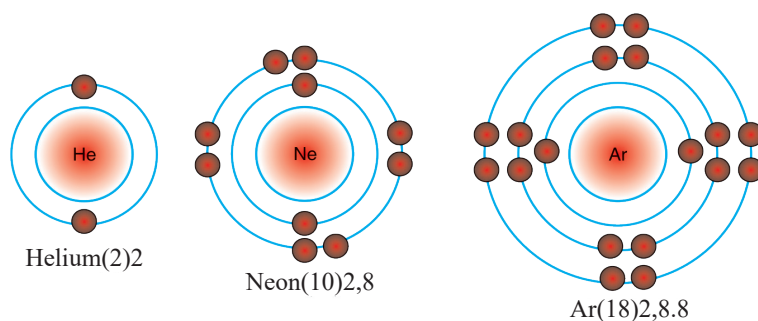


Figure 1. The first three noble gases.

The noble gases have very stable electronic configuration, as reflected by their high ionization energies, low electron affinity and general lack of reactivity. Because all noble gases (except He) have eight valence electrons, many atoms undergoing reaction also attain eight valence electrons. The ns^2np^6 electronic configuration of the valence shell of all noble gas (except Helium) atoms, is commonly called an octet of electrons. The octet rule, a useful generalization that applies to all types of bonding which states that when atoms bond, they lose, gain or share electrons to attain the electronic configuration ns^2np^6 of the nearest noble gas. Nearly every main-group monoatomic ion has a filled outer level of electrons (either two or eight), the same number as in the nearest noble gas.

5.2 TYPES OF CHEMICAL BONDING

There are three main types of chemical bonds – covalent, ionic and metallic bonds. In general, there is a gradual change from more metallic to more non-metallic property when moving from left to right across a period and when moving from bottom to top within a group. Three types of bonding can result from the manner in which the atoms can combine:

- Ionic bond is formed by electron transfer from a metal to a non-metal with different electronegativity values.
- Covalent bond is formed as a result of electron sharing between two non-metals. If the electronegativity values are very similar then it is non-polar covalent bonding but if the electronegativity values are much different, then it is a polar covalent bonding.
- Metallic bond refers to the interaction between the delocalised electrons and the metal nuclei.

Exercise

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?

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.

Table 1

Element	Atomic number	Valence - shell electronic configuration	Number of valence electrons
Helium, He	2		
Neon, Ne	10		
Argon, Ar	18		
Krypton, Kr	36		
Xenon, Xe	54		
Radon, Rn	86		

Ionic Bonding**ACTIVITY 1**

Form a group and discuss the following:

1. Why some atoms are very reluctant to combine with other atoms and exist as single atoms?
2. Is there anything common amongst these atoms?
3. What is special about these atoms with respect to their electronic configuration?
4. What is the common name for this group of elements?
5. What is the reason for their stability?
6. Atoms lose or gain electrons not merely to satisfy the octet rule but to reach a lower energy state in an ionic compound. But it is in reaching this lower energy state that they often tend to follow the octet rule. Explain.

Share your ideas with the rest of the class.

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.

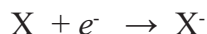
The chemical properties of metals differ from those of non-metals. A metal has 1, 2, or 3 electrons in its outermost shell. Metals tend to lose these electrons and become positively charged ions. For example, if a metal (M) loses one electron, it becomes an ion with a charge of +1.



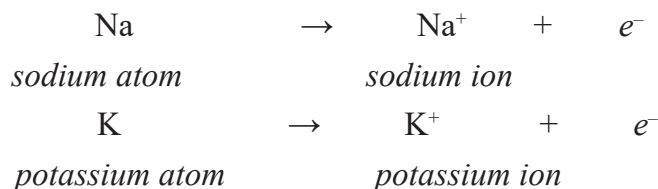
However, if it loses two electrons it becomes an ion with a charge of +2:



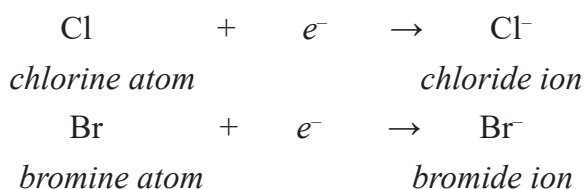
A non-metal may have 4, 5, 6, or 7 electrons in its outermost shell. Non-metals tend to gain electrons to form negatively charged ions. For example, if a non-metal (X) gains one electron, it becomes an ion with a charge of -1.



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.



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.



ACTIVITY 2

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. Sulfur (atomic number = 16)
 - (a) Predict whether each of the elements gain or lose electrons in chemical bond formation.

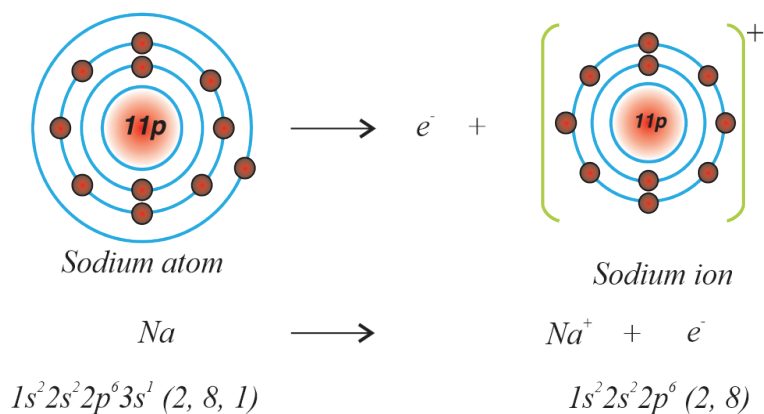
- (b) Write the type of ions they form; and
 (c) Indicate the charges on the ions formed. Present your findings to the class.

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 non-metal 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 electronic 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 electronic 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 non-metal. 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.



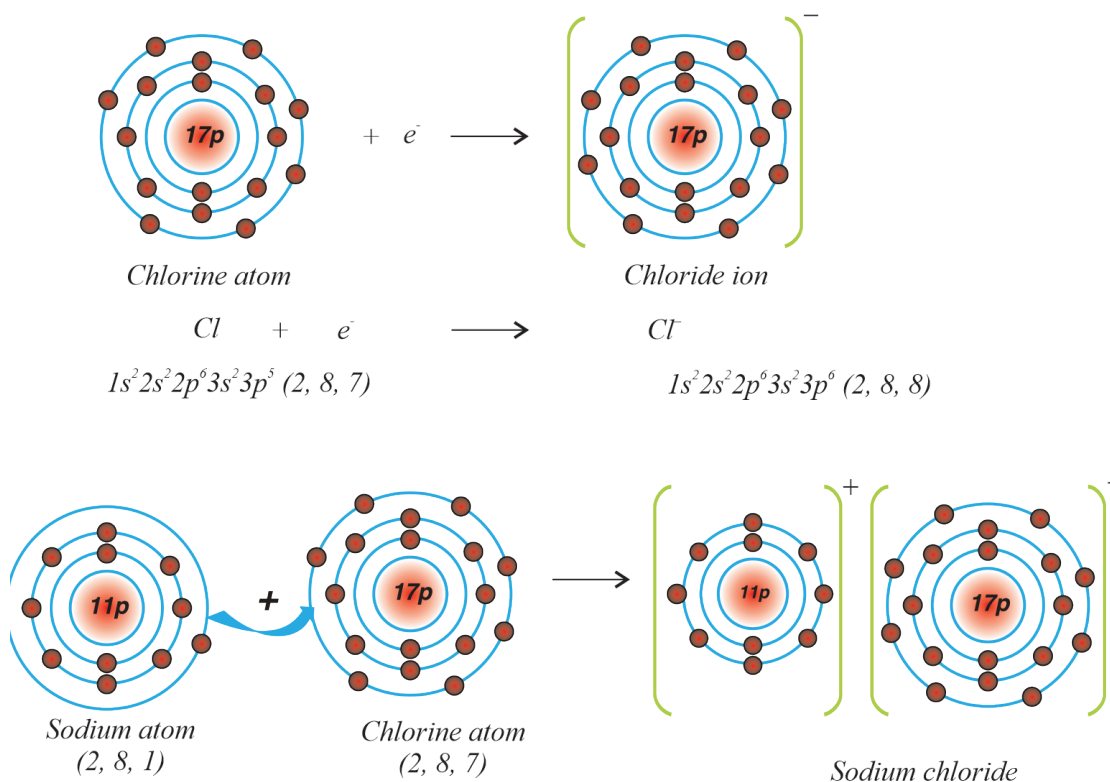
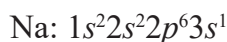


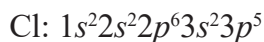
Figure 2. Formation of sodium chloride.

Electron-dot notation is often used to represent the outermost shell electronic 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:



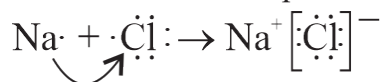
The Lewis symbol for sodium is $\text{Na}\cdot$

Chlorine is a halogen with seven valence electrons:



The Lewis symbol for chlorine is $:\ddot{\text{Cl}}\cdot$

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:



ACTIVITY 3

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:

- Potassium chloride
- Magnesium oxide
- Calcium chloride
- Potassium sulphide
- Aluminum oxide

Share your ideas with rest of the class.

Factors affecting formation of ionic bond

There three main factors that affect the formation of affinity and Lattice energy.

1. Ionization energy

As we have seen in chapter 4, ionization energy is the amount of energy required to remove the outermost electron in the atom when the gas atom is isolated in free space. For the formation of ionic bond there must exist one atom which forms a cation by losing one or more electrons. Hence elements having electrons which are loosely bound have greater tendency to form ionic bond. Plus, metallic elements having lower ionization energy favors the formation of an ionic bond. Alkali and alkaline earth metals have greater tendency to form ionic bond because of ionization energy associated with the higher electron affinity.

2. Electron affinity

In the formation of an ionic compound the atom participating form an anion by gaining an electron(s). Therefore, formation of an ionic bond will be greater in which the elements have higher electron affinity. Halogens have greater tendency to form ionic compounds because of having high electron affinities.

3. Lattice energy

Lattice energy is the amount of energy released when cations and anions are brought together in the crystal lattice to form one mole of the ionic compound. The higher the lattice energy, the greater is the tendency of the formation of an ionic bond.

General properties of ionic compounds

ACTIVITY 4

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

- hard or soft
- brittle or strong
- liquids or solids

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

Experiment 1

Investigation of Solubility of Ionic Compounds.

Objective: To investigate the solubility of NaCl and CuCl_2 .

Apparatus and Chemicals: Test tube, water bath, Bunsen burner, NaCl and CuCl_2 , ethanol, hexane and benzene.

Procedure

- Place about 0.5 g each of NaCl and CuCl_2 in three separate test tubes and add about 2.5 mL of water and shake well.
- If some residue is there in the test tube/s, heat it on Bunsen burner.
- Repeat step 1 using ethanol, hexane and benzene instead of water. These solvents are highly flammable and should be kept away from flames.

(If the salt is soluble at room temperature, do not heat it.)

Observations and analysis:

Prepare an observation table in your notebook for the solubility of NaCl and CuCl_2 in all the three solvents at room temperature and on heating (wherever required) and record the observations.

Table 2

Solvent	NaCl	CuCl ₂
Water		
Ethanol		
Hexane		
Benzene		

Experiment 2

Electrical Conductivity of Ionic Compound.

Objective: To test the electrical conductivity of molten compounds.

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

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

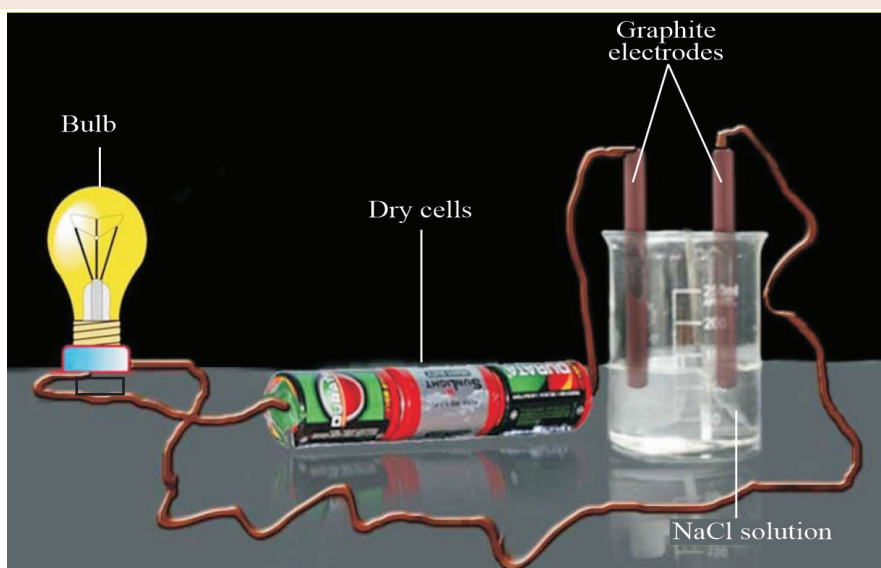


Figure 3. Conductivity of sodium chloride solution

Observations and analysis:

- Does sodium chloride conduct electricity in its solid state?
- What do you observe in the bulb when the sodium chloride is dissolved? What does this indicate?
- What can you conclude from this experiment?
- What are your observations for aqueous copper(II) chloride, benzene and charcoal?

What are your conclusions about:

- Melting point,
- Solubility in polar and non-polar solvents, and
- Electrical conductivity of ionic compounds?

Summary of the general properties of ionic compounds:

- 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. See Figure 4.

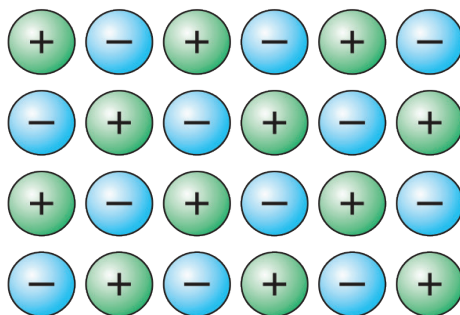


Figure 4. Arrangement of ions in ionic crystal

- 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.
- 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.
- 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.

- Ionic compounds are soluble in polar solvents such as water. They are insoluble in non-polar solvents such as benzene.

Exercise

1. KCl is soluble in water but insoluble in benzene. Why?
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 sulfur
 - (b) sodium and iodine
 - (c) silver and bromine
4. List the properties of ionic compounds.

Covalent Bonding

ACTIVITY 5

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.

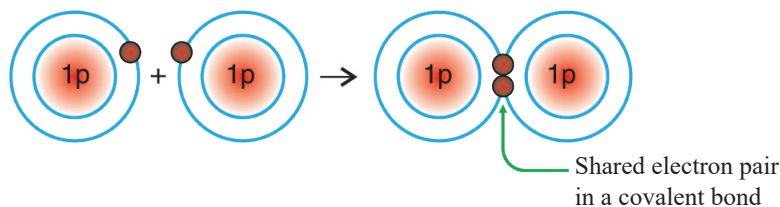


Figure 5. Sharing of electrons between hydrogen atoms in H_2 molecule.

In the hydrogen molecule, each hydrogen atom attains the stable electronic 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 5. 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. See Figure 6.

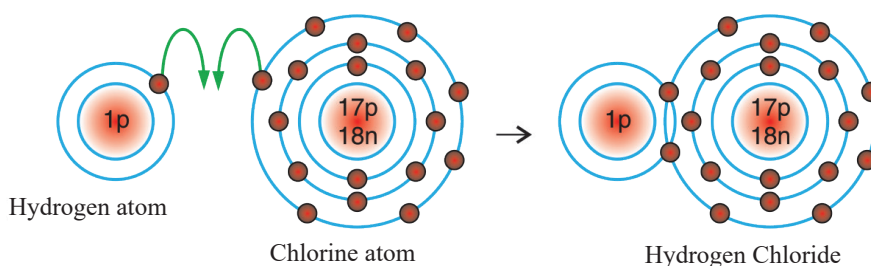


Figure 6. Hydrogen and chlorine share a pair of electrons in HCl.

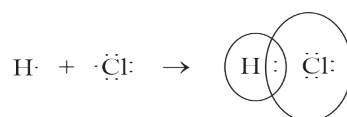
The concept of Lewis formula representation can also be extended to covalent bonds. The Lewis structure for a covalent compound shows the arrangement of atoms in a molecule and all the valence electrons for the atoms involved in the compound. It is conventional to represent the non-bonding (lone pair) 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:

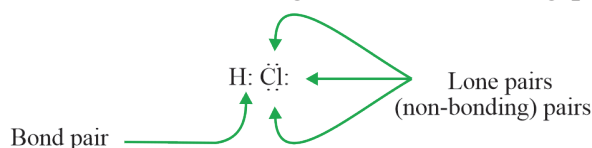


The covalent bond in hydrogen molecule is also written as H – H.

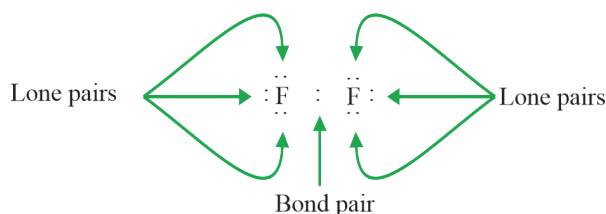
The formation of the covalent bond in hydrogen chloride is shown by the following electron-dot formula. This formula must satisfy the octet rule (for chlorine) and the doublet rule (for hydrogen). As shown in the illustration, these requirements are satisfied. The shared pair belongs to both of the atoms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



The chlorine atom in the molecule has three pairs of electronic, 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.

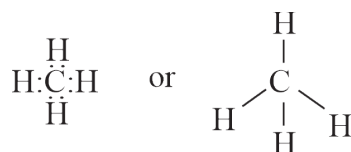


Consider the fluorine molecule, F_2 . The electron configuration of fluorine is 2, 7. Thus each fluorine atom has seven valence electrons. Accordingly, there is only one unpaired electron on fluorine. Therefore, the formation of the fluorine molecule is represented as:

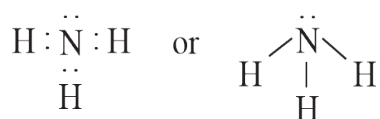


Note that only two valence electrons participate in the formation of fluorine molecule. The others are non-bonding electron (lone pairs). Thus each fluorine atom in fluorine molecule has three lone-pairs of electrons. The resulting molecule is a diatomic molecule. A diatomic molecule consists of two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

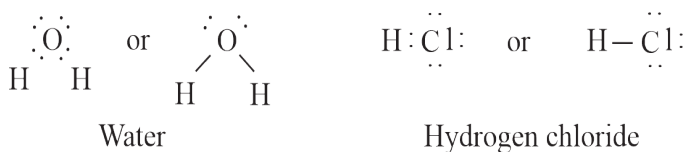
The maximum number of covalent bonds that an atom can form can be predicted from the number of electrons needed to fill its valence shell. For example, each member of Group IVA elements has four electrons in its valence shell, and it needs four more electrons to achieve stable noble-gas electronic configuration. 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.



Factors affecting formation of covalent bond

Electron Affinity - A covalent bond is generally favoured between the two atoms having high electronic affinity.

Ionization energy - Ionization energy of both the atoms participating in covalent bond formation should be high.

Atomic size - Atomic size of atoms forming covalent bond should be smaller. Smaller the atomic radii, more stronger the covalent bond will be.

Electronegativity - The electronegativities of both the atoms should be high. The difference of electronegativity between the two atoms should be minimum.

Exercise

Write a plausible Lewis structure of:

- nitrogen trichloride, NCl_3
- chlorate ion, ClO_3^-
- phosphonium ion, PH_4^+
- phosgene, COCl_2

Types of covalent bonds

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

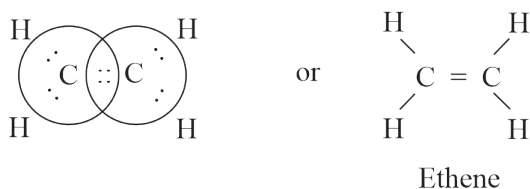
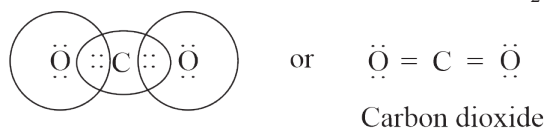
Atoms can form different types of covalent bonds. These are single bonds, double bonds and triple bonds.

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

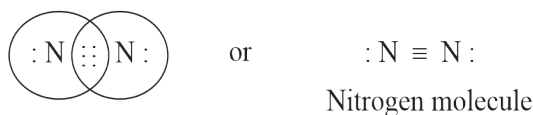
How are the covalent bonds in H_2 , Cl_2 and HCl formed?

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

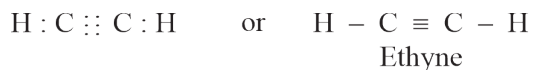
share two pairs of electrons, the covalent bond is called a double bond. For example, double bonds are found in molecules of carbon dioxide (CO_2) and ethene (C_2H_4).



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



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



ACTIVITY 6

Form a group and perform the following tasks.

- 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.
 - N_2
 - H_2
 - O_2
 - C_2H_4
- Based on your models, identify which species contain:
 - only single bonds
 - double bonds
 - triple bonds

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

Exercise

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

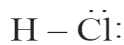
Electronegativity and polarity in covalent molecules

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.



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.



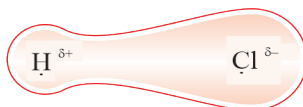
A chemical bond in which shared electrons spend more time in the vicinity of one atom than the other is called a polar covalent bond, or simply a polar bond. Polarity of bonds is caused by differences in the electronegativity of the two atoms forming the bonds. Electronegativity is the ability of an 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 possesses 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.

Atoms of elements with widely different electronegativities tend to form ionic bonds (such as those that exist in NaCl and CaO compounds) with each other because the atom of the less electronegative element gives up its electron(s) to the atom of the more electronegative element. An ionic bond generally joins an atom of a metallic element and an atom of a non-metallic element.

Atoms of elements with comparable electronegativities tend to form polar covalent bonds with each other because the shift in electron density is usually small. Most covalent bonds involve atoms of non-metallic elements. Only atoms of the same element, which have the same electronegativity, can be joined by a pure covalent bond. These trends and characteristics are what we would expect, given our knowledge of ionization energies and electron affinities

There is no sharp distinction between a polar covalent bond and an ionic bond, but the following rules are helpful as a rough guide. An ionic bond forms when the electronegativity difference between the two bonding atoms is 2.0 or more. This rule applies to most but not all ionic compounds. A polar covalent bond forms when the electronegativity difference between the atoms is in the range of 0.3–2.0. If the electronegativity difference is below 0.3, the bond is normally classified as a covalent bond, with little or no polarity. Sometimes chemists use the quantity percent ionic character to describe the nature of a bond.

Electronegativity is a relative concept, meaning that an element’s electronegativity can be measured only in relation to the electronegativity of other elements. The

American chemist Linus Pauling devised a method for calculating relative electronegativities of most elements. These values are shown in Figure 7.

															8A							
1A																						
H	2A											3A	4A	5A	6A	7A						
2.1	Li	Be											B	C	N	O	F					
1.0	1.5											2.0	2.5	3.0	3.5	4.0						
Na	Mg	3B	4B	5B	6B	7B	8B			1B	2B	Al	Si	P	S	Cl						
0.9	1.2											1.5	1.8	2.1	2.5	3.0						
K	Ca	Se	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr					
0.8	1.0	1.3	1.5	1.6	1.6	1.5	1.8	1.9	1.9	1.9	1.6	1.6	1.8	2.0	2.4	2.8	3.0					
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe					
0.8	1.0	1.2	1.4	1.6	1.8	1.9	2.2	2.2	2.2	1.9	1.7	1.7	1.8	1.9	2.1	2.5	2.6					
Cs	Ba	La-Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At						
0.7	0.9	1.0-1.2	1.3	1.5	1.7	1.9	2.2	2.2	2.2	2.4	1.9	1.8	1.9	1.9	2.0	2.2						
Fr	Ra																					
0.7	0.9																					

Figure 7. The electronegativities of common elements.

Examples

Classify the following bonds as ionic, polar covalent, or covalent.

- The bond in HCl,
- The bond in KF, and
- The C-C bond in $\text{H}_3\text{C}-\text{CH}_3$.

Strategy . We follow the 2.0 rule of electronegativity difference and look up the values in Figure 7.

Solution:

- The electronegativity difference between H and Cl is 0.9, which is appreciable but not large enough (by the 2.0 rule) to qualify HCl as an ionic compound. Therefore, the bond between H and Cl is polar covalent.
- The electronegativity difference between K and F is 3.2, which is well above the 2.0 mark; therefore, the bond between K and F is ionic.
- The two C atoms are identical in every respect; they are bonded to each other and each is bonded to three other H atoms. Therefore, the bond between them is purely covalent.

Exercise

- How many electrons are shared in a:
 - single bond,
 - double bond, and
 - triple covalent bond?
- Draw Lewis structures for:
 - H_2
 - Cl_2
 - C_3H_6
- Draw Lewis structures for each of the following molecules:
 - HBr
 - CO_2
 - H_2O

Also indicate the partial charges using δ^+ and δ^-

- Which of the following bonds is covalent, which is polar covalent, and which is ionic?
 - the bond in CsCl ,
 - the bond in H_2S ,
 - the NN bond in H_2NNH_2

General properties of covalent compounds

- Covalent compounds are generally liquids or gases at ordinary temperature. For example : water and ethyl alcohol are liquids. Hydrogen chloride, methane and carbon dioxide are gases. Some covalent compounds are solids (e.g. sugar)
- Compared to ionic compounds, covalent compounds have relatively lower melting points and boiling points.
- They do not conduct electric current when molten or in aqueous solution, because they consist of molecules rather than of ions.
- 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

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.

Procedure

- 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.
 - What do you observe? Does the solid melt or vaporize?
 - Is the melting point high or low?
- 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.

Table 3

Substances	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 tube, Test tube rack.

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

Procedure:

- 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.
- 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.

Table 4

Substances	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

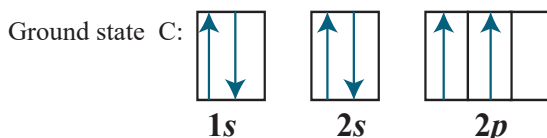
- Which of the following molecules contain a covalent bond?
 - CaO
 - HCl
 - CO
 - SO₂
 - Na₂O
 - PCl₃
 - MgO
 - NaH
 - CH₄
- Which of the following contain a coordinate covalent bond?
 - H₃O⁺
 - NH₃
 - NH₂⁻
 - CaO
- Which of the following molecules are polar ?
 - SO₂
 - CO₂
 - H₂O
 - CS₂
 - BCl₃
 - CH₄CH₃Cl
- Which of the following are non-polar covalent compounds/molecules?
 - O₂
 - HCl
 - CH₄
 - O₃
 - H₂O
 - Cl₂
 - Br₂

5.3 HYBRIDIZATION OF ATOMIC ORBITALS

The concept of atomic orbital overlap should apply also to polyatomic molecules.

The concept of hybridization is the overlap or blending of s , p and d orbitals to explain bond formation.

In applying the valence bond method to the ground state electronic configuration of carbon you can assume that the filled $1s$ orbital will not be involved in the bonding and focus your attention on the valence-shell orbitals.



From its ground state configuration, two unpaired electrons in the $2p$ subshell are observed. One can predict the simplest hydrocarbon molecule to be CH_2 , by overlapping the two unpaired electrons from two H atoms with the two unpaired electrons of carbon. However, you might question the existence of this molecule because it does not follow the octet rule. In fact, experiment shows that CH_2 is not a stable molecule. The simplest stable hydrocarbon is methane, CH_4 . To account for this, you need an orbital diagram that shows four unpaired electrons in the valence shell of carbon, requiring four bonds (and therefore four atoms of hydrogen). To get such a diagram imagine that one of the $2s$ electrons is promoted to the empty $2p$ orbital. To excite the $2s$ electron to a higher energy sublevel, energy must be absorbed. The resulting electronic configuration is that of an excited state having energy greater than the configuration in the ground state.

Valence bond theory proposes that the one $2s$ and all three $2p$ orbitals of the carbon atom ‘mix’ to produce four new orbitals that are equivalent to each other in energy and in shape and pointing in different directions with equal H–C–H bond angles. This blending is called hybridization and the resultant orbital as hybrid orbitals. It is a hypothetical process (not an observed one) that can be carried out as a quantum mechanical calculation. The symbols used for hybrid orbitals identify the kinds and numbers of atomic orbitals used to form the hybrids.

In hybridization scheme, the number of new hybrid orbitals is equal to the total number of atomic orbitals that are combined. Hybridization provides a useful and convenient method of predicting the shapes of molecules. It must be noted that though it is valuable in predicting and describing the shape, it does not explain the reason for the shape.

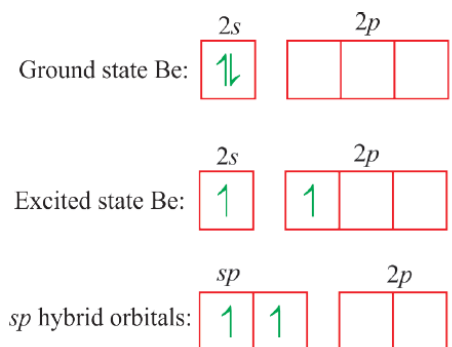
Exercise

- How do carbon and silicon differ with regard to the types of orbitals available for hybridization? Explain
- Are these statements true or false? Correct any that is false.
 - Two σ bonds comprise a double bond.
 - A triple-bond consists of one π bond and two σ bonds.
 - Bonds formed from atomic s orbitals are always σ bonds.
 - A π -bond consists of two pairs of electrons.
 - End-to-end overlap results in a bond with electron density above and below the bond axis.

sp hybrid orbitals

When two electron sets surround the central atom, you observe a linear molecular shape. VB theory explains this by proposing that mixing two non-equivalent orbitals of a central atom one s and one p , gives rise to two equivalent sp hybrid orbitals that lie 180° apart. Combination (mixing) of one s and one p orbital is called sp hybridization and the resultant orbitals are called sp hybrid orbitals.

For instance, the beryllium atom has two electrons and four orbitals in its valence shell. In the triatomic molecule BeCl_2 , which is present in gaseous BeCl_2 at high temperatures, the $2s$ and one of the $2p$ orbitals of the Be atom are hybridized into sp hybrid orbitals. The remaining two $2p$ orbitals are left unhybridized and unoccupied in the orbital diagram.



Since there are no unpaired electrons this atom cannot form any covalent bonds. It follows that in BeCl_2 , the Be atom is not in the ground state but it is in the excited state, with sufficient extra energy to unpair a $2s$ electron.

As shown in Figure 10 the two sp hybrid orbitals are directed along a straight line, 180° apart. It is predicted that the BeCl_2 molecule should be linear, and this prediction

is confirmed by experimental evidence. Most examples of sp hybridization are found in organic molecules, especially those with triple bonds.

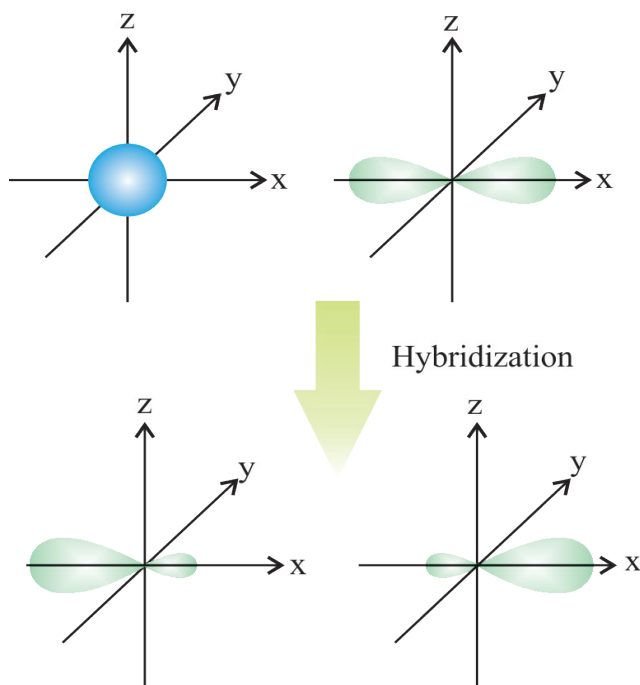
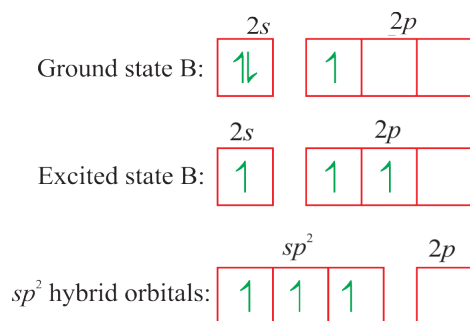


Figure 8. The sp -hybridization scheme.

sp^2 hybrid orbitals

Now turn your attention to boron, a Group IIIA element. The boron atom has four orbitals but only three electrons in its valence shell. In most boron compounds the hybridization scheme combines one $2s$ and two $2p$ orbitals into three sp^2 hybrid orbitals. Using orbital diagrams to represent this hybridization, we have:

sp^2 hybridization for B



The atom in its ground state has only one unpaired electron, so that it can form only one covalent bond, but in the excited state there are three unpaired electrons, hence three bonds can be formed. See figure 11.

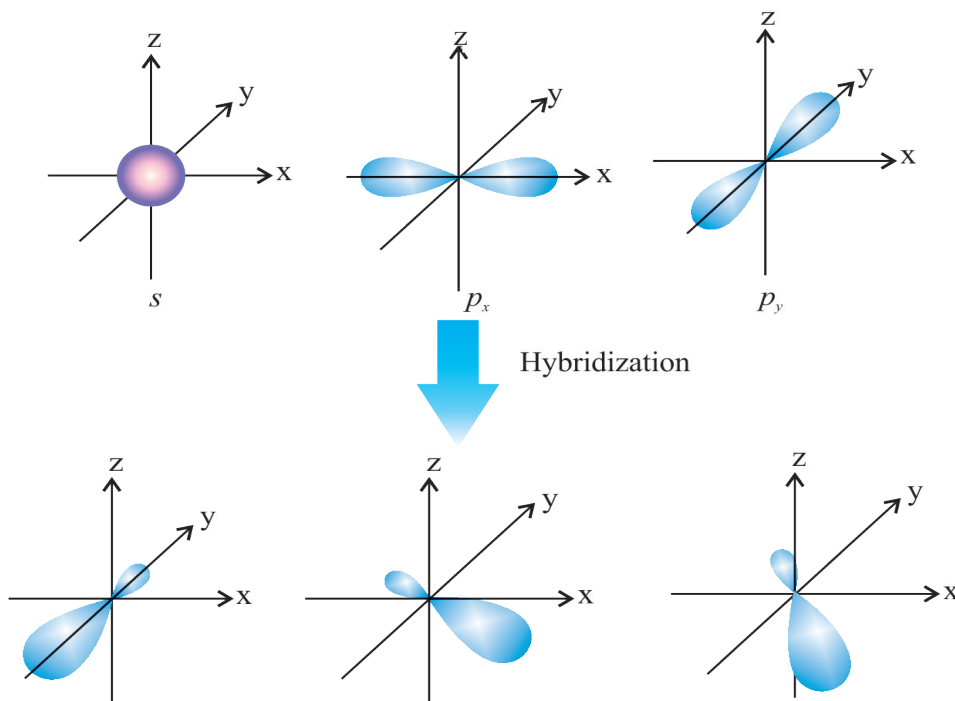


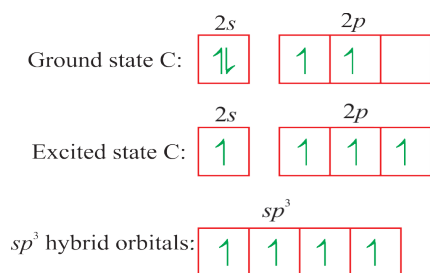
Figure 9. The sp^2 hybridization scheme.

As shown in Figure 9 the sp^2 hybrid orbitals are distributed geometrically within a plane at 120° angles. The valence bond method predicts that BCl_3 is a trigonal planar molecule with 120° Cl–B–Cl bond angle.

This indeed is what is observed experimentally. By far the most common examples of sp^2 hybridization are found in organic molecules with double bonds.

sp³ hybridization

Carbon, the central atom in a molecule of methane CH_4 has only two unpaired electrons in the ground state. The two electrons in the $2p$ level are not paired, i.e., put into the same box, in accordance with Hund's rule. Carbon in its excited state can form four bonds. The one $2s$ and three $2p$ -orbitals of carbon are mixed to produce four new orbitals that are equivalent to each other in energy and in shape, but pointing in different directions, as shown in Figure 10.

sp³ hybridization for carbon

The four sp^3 hybrid orbitals point to the corners of a regular tetrahedron, and make a bond angle of 109.50° in CH_4 (Figure 11).

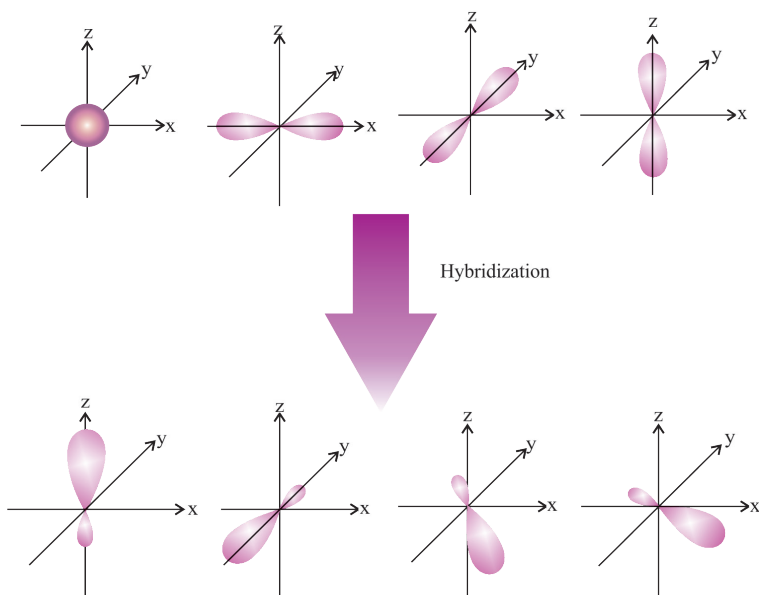


Figure 10. The sp^3 hybridization.

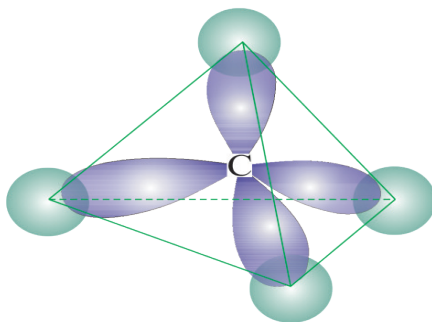


Figure 11. sp^3 Hybrid orbitals and bonding in methane.

We might also expect to use sp^3 hybridization not only for structures of the type AX_4 type (as in CH_4), but also for AX_3E type (as in NH_3) and AX_2E_2 type (as in H_2O). Nitrogen has three unpaired electrons in its ground state, sufficient to form three bonds; so it is not necessary to excite the atom.

sp^3 hybridization of the central atom N in NH_3 :

Accounts for the formation of three N–H bonds and a lone-pair of electrons on the N atom (Figure 12). The predicted H–N–H bond angles of 109.5° are close to the experimentally observed angles of 107° .

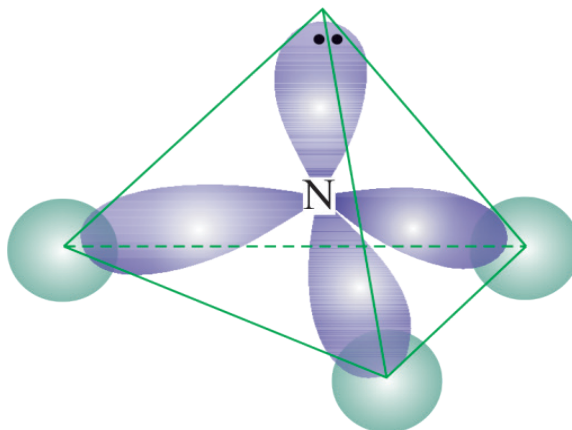
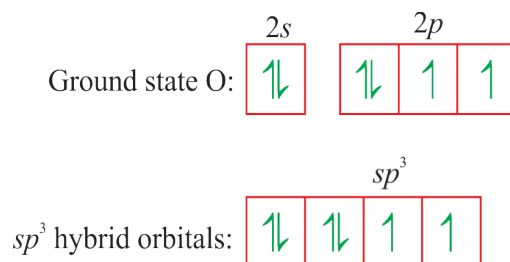


Figure 12. sp^3 hybrid orbitals and bonding in NH_3 .

A similar scheme for H_2O accounts for the formation of two O–H bonds and two lone-pairs of electrons on the oxygen atom.



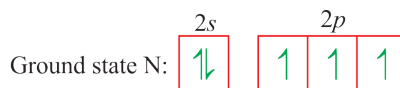
The predicted H–O–H bond angle of 109.5° is also reasonably close to the observed angle of 104.5° .

5.4 SIMPLE MOLECULES AND THEIR SHAPES

The geometrical shapes of some simple molecules can be predicted based on the following general patterns:

Molecular Shapes with Two Electron Sets (Linear Arrangement- AX_2 type)

The repulsion of two electron groups with each other results in the assignment on opposite sides of the central atom in a straight line. The linear arrangement of electron sets results in a linear molecular shape and bond angle of 180° . All AX_2 type molecules or ions are linear in geometry. The examples of the molecules of this type are CO_2 , $BeCl_2$, CS_2 , HCN , etc.

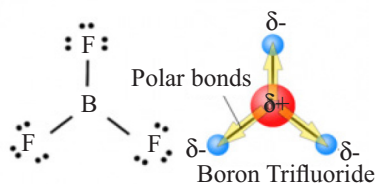


Molecular shapes with three electron sets (Trigonal Planar Arrangement, AX_3 type)

Three electron sets around the central atom repel each other to lie at the corners of an equilateral triangle. This is the trigonal planar arrangement and the ideal bond angle is 120° : Example: BF_3 , NO_3^- , $HCHO$ (formaldehyde) etc. Another molecular shape is possible within this electron set arrangement, with two bonding and one lone pair (AX_2E type). The examples of this type include $PbCl_2$, $SnBr_2$, SO_2 , O_3 , etc. They have bent (V-shaped or angular) geometry. This is the first example of the effect of a lone-pair on adjacent bonding-pairs. Since a lone-pair is held on the central atom, it exerts a stronger repulsion than bonding pair.

Thus, in AX_2E type species a lone-pair repels bonding-pairs more strongly than bonding-pairs repel each other. The repulsion increases the angle between lone-pair and bonding-pair, which decreases the angle between bonding-pairs. Note the large decrease from the ideal 120° angle in AX_2E type molecules. (Example $SnCl_2 = 95^\circ$).

Boron trifluoride contains three covalent bonds, or bonding pairs. In the most stable arrangement, the three BF bonds point to the corners of an equilateral triangle with B in the center of the triangle:

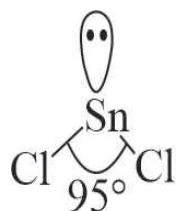


Triangular Planar

Molecular Shapes with Four Electron Sets (Tetrahedral Arrangement, AX_4 Type)

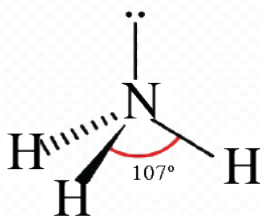
Note that AX_3E and AX_2E_2 types are with four electron sets like AX_4 type but differ in shape and bond angle because of the presence of lone pair(s) on the central atom.

In three dimensions, the four electron sets can move farther apart and point to the corners of a tetrahedron, giving a bond angle of 109.5° . Therefore, all molecules or ions with four electron sets around the central atom adopt the tetrahedral arrangement. Some of the examples of this type are CH_4 , NH_4^+ , SO_4^{2-} . When one of the four electron sets in the tetrahedral arrangement is a lone pair, the molecular shape is that of a trigonal pyramid (AX_3E type). The measured bond angle is slightly less than the ideal 109.5° . In ammonia (NH_3), for example, the lone pair forces the N–H bonding pairs together, and the H–N–H bond angle is 107.3° . When the four sets of electrons around the central atom include two bonding and two non-bonding sets, the molecular shape is bent, V-shaped or angular (AX_2E_2 type). Recall that one of the shapes in the trigonal planar arrangement - that with two

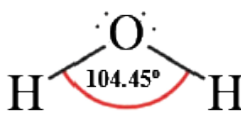


Trigonal Pyramidal

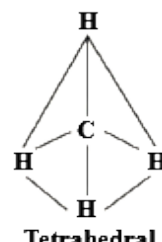
bonding sets and one lone pair-is also called angular or bent or V-shaped (AX_2E type), but its ideal bond angle is 120° , not 109.5° . Water is the most important V-shaped molecule in the tetrahedral arrangement of electrons. We would expect the repulsions between its two lone-pairs have a greater effect on bond angle than the repulsions from the single lone pair in NH_3 , and observation confirms this: two lone-pairs on the central O atom compress the H–O–H bond angle to 104.5° .



NH_3 - Pyramidal



H_2O - V-shaped



CH_4 - Tetrahedral

Exercise

For each of the following molecular formula (AX_mE_n), A is the central atom, X is the terminal atom, m is the number of terminal atoms bonded to the central atom; E is the lone-pair of electrons and n is the number of lone-pairs possessed by the central atom. The first item has been completed as an example. Fill the rest of the table after drawing in your note book.

Table 5

No of electron pairs	Molecular formula	Molecular geometry	Electron set geometry	Sketch	Examples
2	AX_2	Linear	Linear	$X - A - X$	$Cl - Be - Cl$
3	AX_2E AX_3				
4	AX_4 AX_3E AX_2E_2				

5.5 MOLECULAR GEOMETRY

Molecular geometry is the three-dimensional arrangement of atoms in a molecule. A molecule's geometry affects its physical and chemical properties, such as melting point, boiling point, density, and the types of reactions it undergoes. In general, bond lengths and bond angles must be determined by experiment. However, there is a simple procedure that enables us to predict with considerable success the overall geometry of a molecule or ion if we know the number of electrons surrounding a central atom in its Lewis structure. The basis of this approach is the assumption that electron pairs in the valence shell of an atom repel one another.

The valence shell is the outermost electron occupied shell of an atom; it holds the electrons that are usually involved in bonding. In a covalent bond, a pair of electrons, often called the bonding pair, is responsible for holding two atoms together. However, in a polyatomic molecule, where there are two or more bonds between the central atom and the surrounding atoms, the repulsion between electrons in different bonding pairs causes them to remain as far apart as possible.

The geometry that the molecule ultimately assumes (as defined by the positions of all the atoms) minimizes the repulsion. This approach to the study of molecular geometry is called the valence-shell electron-pair repulsion (VSEPR) model, because it accounts for the geometric arrangements of electron pairs around a central atom in terms of the electrostatic repulsion between electron pairs.

The basic principle of VSEPR theory is that the pair of valence shell electrons around the central atom stays as far apart from each other as possible to minimize repulsion among them. For simplicity, a set of electrons is defined as any number of electron pairs that occupies a localized region around a central atom. This set

may consist of a single bond (–), a double bond (=), a triple bond (\equiv), lone-pair(s) or in some cases even a lone (single) electron. Each of these is a separate set of electrons that repels the other groups and occupies as much space as possible around the central atom - one set of electrons repels the other sets to maximize the angles between them. The three-dimensional arrangement of these sets gives rise to the shape of the molecule.

Electron Pair Arrangement and Molecular geometry

The electron pair arrangement is defined by the sets of electrons, both bonding and non-bonding (lone-pair), around the central atom. On the other hand the molecular shape is defined by the relative positions of the atomic nuclei. Molecular shapes that occur when all the surrounding electron sets are bonding sets differ from molecular shapes when some of the electron sets are non-bonding sets. Thus, the same electron set arrangement can give rise to different molecular shapes. To classify molecular shapes, the AX_mE_n designation is assigned, where A is the central atom, X is the terminal atom, E is lone-pair (nonbonding) electron sets, m and n are positive integers.

The arrangements that best minimize repulsions naturally depend on the number of electron sets. Repulsive forces among valence pairs diminish in the following order:

Lone pair - lone pair > lone pair - bond pair > bond pair - bond pair.

In this section, you will see electron sets ranging from two to six around the central atom. Two electron sets locate themselves on opposite sides of the atom in a linear arrangement, three sets form a trigonal planar structure, four sets arrange themselves at the corners of a tetrahedron, five sets define a trigonal bipyramid, and six sets form an octahedron.

Guidelines for Applying VSEPR Model

The VSEPR model is an approach of using the number of electrons surrounding a central atom to study the molecular structures, based on the theory that the structure around a given atom is determined principally by minimizing electron-pair repulsion. Here are some guidelines for applying the VSEPR model:

1. First, write the Lewis structure of the molecule. In this model, consider only the electron sets (pairs) of the central atom.
2. Count the total number of electron sets around the central atom, including both the bonding pairs and lone pairs. Usually a lone pair, a single unpaired electron and any bond: single, double or triple, each

- count as one area of electron density. For species with more than one central atom, treat each central atom separately.
3. Use the VSEPR geometry to predict the shape of the molecule.
 4. In predicting bond angles, bear in mind that the repulsion of the lone pairs is stronger than between those of the bonding pairs. VSEPR theory is best used as a tool to explain why a given structure is distorted rather than as a predictive tool.

In explaining why the distortion occurs you need to balance three competing influences:

- Lone pairs take up more space than bonding pairs. Repulsions between electrons in a lone-pair and the other electrons (in either bonds or other lone pairs) may cause distortions in the structure.
- Triple bonds are fatter than double bonds, which are fatter than single bonds. The amount of electron-electron repulsion experienced between a bond and lone pairs or other bonding pairs decreases in the order:
$$\text{triple} > \text{double} > \text{single}$$
- Therefore, we expect that a triple bond may cause more distortions in the structure than either a double or single bond and that a double bond will cause more distortion when compared to a single bond.
- Bonds which involve a significant difference in electronegativity between the atoms in the bond will have the electrons in the bond distorted toward the more electronegative atom. This will decrease electron density near the central atom and lessen the repulsion between this bonding-pair and other electron pairs in the molecule.

5.6 METALLIC BONDING

ACTIVITY 7

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?

In simple terms, metallic bonding is referred to as bonding in metal atoms. It is also defined as interaction between metal nuclei and the delocalized electrons.

Delocalized electrons are also called as conduction electrons. Metal nuclei are the positive ions and so metallic bonding can be imagined as sea of electrons in which positive metal ions are embedded. Positive metal ions are called Kernels. Thus metallic bonding can be summarized as: The force of attraction which binds together the positive metal ions or Kernels with the electrons within its sphere of influence.

Is a metal made up of atoms or ions?

The strength of the metallic bond depends on the:

1. Number of electrons in the delocalised 'sea' of electrons. (More delocalised electrons results in a stronger bond and a higher melting point.)
2. Packing arrangement of the metal atoms. (The more closely packed the atoms are the stronger the bond is and the higher the melting point.)

For example, as illustrated in Figure 13, a sodium metal crystal is a lattice-like array of Na^+ ions surrounded by a sea of mobile bonding valence electrons.

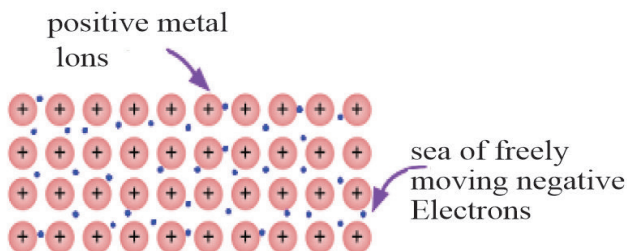


Figure 13. Metallic bonding in sodium metal

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

Properties of metals related to the concept Metallic Bonding

ACTIVITY 8

Form a group and discuss the following questions:

1. Use the Periodic Table to arrange the following in order of increasing melting point K, Br_2 , Mg and O_2 . Explain the reason for the order you choose.
2. List four physical characteristics of a solid metal .
3. Briefly account for the following relative values:
 - (a) Lithium boils at 1317°C and melts at 179°C . The boiling point is about 1138°C higher than its melting point.

- (b) The melting point of Li and Be are 180°C and 1287°C , respectively.
- (c) The melting point of sodium is 89°C , where as that of potassium is 63°C .
4. For each of the following pairs, which element will have the greater metallic character:
- Li or Be
 - Li or Na
 - Sn or Pb and B or Al?

Explain briefly. Share your ideas with the rest of the class

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.

Exercise

- Describe how a metallic bond is different from those of an ionic bond and a covalent bond.
- Explain thermal and electrical conductivity in metals.
- Is metallic bonding responsible to form compounds?
- Give some properties of metals in relation with metallic bonding

5.7 INTERMOLECULAR FORCES

ACTIVITY 9

Form a group and discuss the following questions:

- Why does a polar liquid generally have a higher normal boiling point than a non-polar liquid of the same molecular mass?
- State the principal reasons why CH_4 is a gas at room temperature whereas H_2O is a liquid.

3. Only one of these substances is a gas at STP: NI_3 , BF_3 , PCl_3 , and CH_3COOH . Which do you think it is and why? Share your ideas 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:

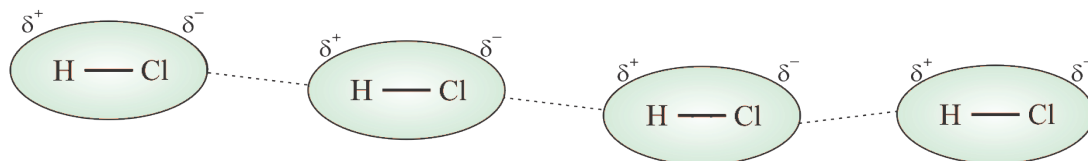


Figure 14. Dipole-dipole interactions in HCl molecules.

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.

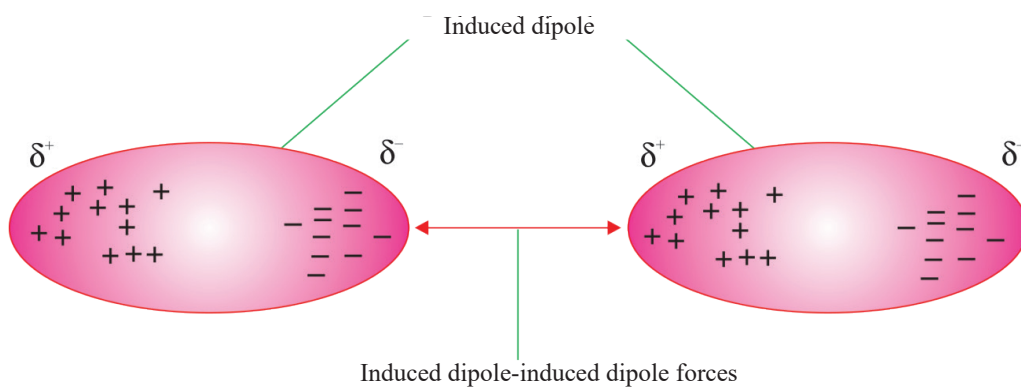


Figure 15. Induced dipole-induced dipole forces between non-polar molecules.

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, H₂, and oxygen, O₂), and (chlorine, Cl₂, and bromine, Br₂).

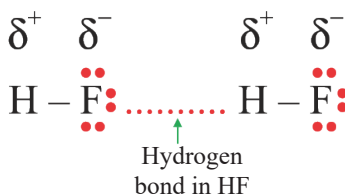
C. Hydrogen bonding

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?

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 (H₂O) and ammonia (NH₃).

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.



5.8 COORDINATE COVALENT BOND

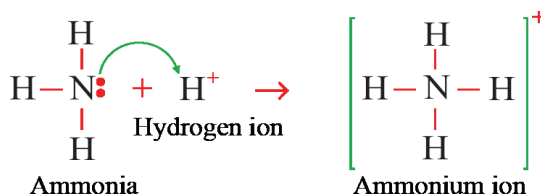
A covalent bond in which one atom donates both electrons of the bond is called a coordinate covalent bond. It is also called a dative bond. Such a bond is hypothetically represented as:



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.

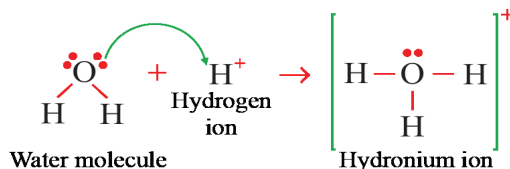
For an atom to act as a donor, it must contain lone pair of electrons in its valence shell and the acceptor atom must have at least one vacant orbital.

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 electronic arrangements, oxygen donates a pair of electrons to the carbon atom. In the process a coordinate covalent bond is formed between the two atoms.



KEY TERMS

- 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
- Ionization energy
- Ionic bonding
- Inter-molecular forces
- London forces
- Lone-pair electrons
- Malleability
- Metallic bond
- Mobile electrons
- Noble gases
- Polar bond
- Polar covalent bond
- Polarity

- Sea of electrons
- Single-bond
- Triple-bond
- Valence electrons

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 $N \equiv N$.
- 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.
- In a covalent bond between atoms of different electronegativity, electrons are displaced toward the atom with the higher electronegativity. In terms of electronegativity differences chemical bonds vary over the range: non-polar to polar covalent to ionic.
- 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 non-polar 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)
- Molecular geometry is determined by the spatial orientations of the atomic orbitals involved in bonding.
- The VB method often requires that bonding atomic orbitals be hybridized in order to rationalize known structures of molecules. A hybridized orbital is some combination of s , p and d orbitals, such as sp , sp^2 , sp^3 .

- The geometric distribution of hybridized orbitals in the valence bond method is the same as the electron set geometry proposed by VSEPR theory.
- Hybrid orbitals overlap in the usual way (end-to-end) and form σ bonds. Unhybridized p orbitals overlap in a side-by-side manner and give rise to π bonds. A double bond consists of one σ bond and one π bond; a triple bond, one σ bond and two π bonds.

Exercises

Part I. Multiple Choice Questions

1. In the formation of ionic bonding, valence electrons are:
 - (a) shared
 - (b) delocalized
 - (c) transferred
 - (d) not affected
2. “Two atoms each provide two electrons that are shared by the two atoms” This is a description of
 - (a) triple covalent bond
 - (b) double covalent bond
 - (c) coordinate covalent bond
 - (d) single covalent bond
3. A non-metal usually forms two covalent bonds in its compounds. How many electrons will it have in its valence shell?
 - (a) 2
 - (b) 4
 - (c) 6
 - (d) 8
4. The noble gases do not usually form chemical compounds because:
 - (a) they have very stable nuclei.
 - (b) the bonds between their atoms are very strong
 - (c) they already have paired valence shell electrons.
 - (d) they are not polar.
5. Which of the following compounds contain both ionic and covalent bonds?
 - (a) CO_2
 - (b) Cl_2O
 - (c) Na_2CO_3
 - (d) BaCl_2

6. Which of the following molecules would you expect to be non-planar?
- (a) CH_4
 - (b) XeF_4
 - (c) BCl_3
 - (d) HCHO
7. If a molecule has a trigonal pyramidal shape, how many non-bonding pairs of electrons are present in the valence shell of the central atom?
- (a) 1
 - (b) 2
 - (c) 3
 - (d) 4
8. Of the following, the most polar bond is:
- (a) P – Cl
 - (b) Si – F
 - (c) S – O
 - (d) C – N
9. The carbon atoms in ethane (C_2H_6), ethene (C_2H_4) and ethyne (C_2H_2) provide, respectively, examples of the three common types of hybridization corresponding to:
- (a) sp, sp^2, sp^3
 - (b) sp^3, sp^2, sp
 - (c) sp, sp^3, sp^2
 - (d) sp^3, sp, sp^2
10. Which of the following correctly describes a π bond?
- (a) It is formed by the end-to-end interaction of p -orbitals and has a low electron density on the internuclear axis.
 - (b) It is formed by the parallel interaction of p -orbitals and has a high electron density just above and below the internuclear axis.
 - (c) It is formed by the interaction of s -orbitals and has a low electron density on the internuclear axis.
 - (d) It is formed by the interaction of s -orbitals and has a high electron density on the internuclear axis.
11. In which of the following compounds would inter-molecular hydrogen-bonding occur?
- (a) HCHO
 - (b) CH_3OH

- (c) PH_3
(d) COCl_2
12. Which one of the following usually produces the weakest interaction between particles of similar molar mass?
(a) Van der Waals forces
(b) dipole-dipole forces
(c) covalent bonds
(d) hydrogen-bonding
13. If an element in group IIA of the periodic table formed a compound with an element in group VIIA of the periodic table, the compound formed is likely to:
(a) be a crystalline solid
(b) dissolve in non-polar solvents
(c) have a low boiling point
(d) conduct electricity in the solid state
14. Which of the following substances would you expect to have the lowest boiling point?
(a) CsCl
(b) SrSO_4
(c) Sc_2O_3
(d) AsCl_3
15. Which of the following is a compound?
(a) Cl_2
(b) HCl
(c) Na
(d) liquid oxygen

Part II. Answer the following questions:

16. Give a simple explanation of the following, using an example:
(a) covalent bond
(b) hydrogen bonding
(c) an ionic bond and metallic bond
(d) a dative bond
17. Predict the shape for the following molecules:
(a) SiF_4
(b) PCl_3
(c) H_2S

- (d) NF_3
(e) CCl_4
(f) PCl_5
18. Classify the given molecules as polar or non-polar:
- (a) SiF_4
(b) NF_3
(c) PCl_3
(d) H_2S
(e) CCl_4
(f) CO_2
19. Carbon and oxygen can bond either by a single bond (as in $\text{CH}_3\text{-OH}$), a double bond (as in $\text{O}=\text{C}=\text{O}$), or a triple bond (as in $\text{HC}\equiv\text{CH}$).
- (a) Describe these three types of bonds in terms of σ -bonds and π -bonds.
(b) How would you expect the length of the carbon-oxygen bond to vary in the three given examples?
20. Explain the following in terms of the intermolecular forces that exist in them:
- (a) Water is a liquid at room temperature and atmospheric pressure, but hydrogen sulphide is a gas.
(b) At room temperature and atmospheric pressure chlorine is a gas, bromine is a liquid and iodine is a solid.
(c) Pentan-1-ol boils at 137°C , where as pentan-3-ol boils at 116°C .
(d) The boiling point of sulfur dioxide is 24°C higher than that of chlorine.



C10CH06

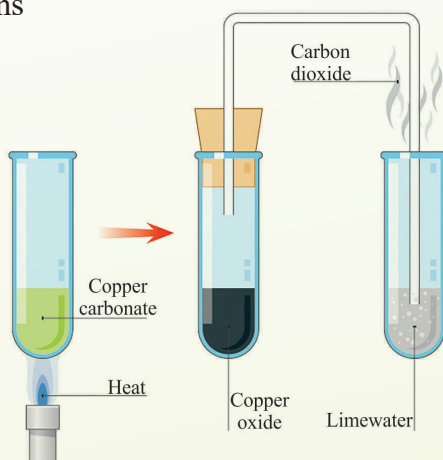
CHAPTER

6

ATOMS, MOLECULES, IONS AND STOICHIOMETRY

Chapter Contents

- 6.1 Fundamental Laws of Chemical Reactions
- 6.2 Molecules and Ions
- 6.3 The Mole Concepts/Stoichiometry
- 6.4 Percentage Composition of Compounds
- 6.5 Determination of Formulas
- 6.6 Chemical Reactions
 - Key Terms
 - Summary
 - Exercises



Chapter Outcomes

Upon completion of this chapter, students will be able to:

- distinguish the fundamental chemical law;
- distinguish atoms, molecules and ions;
- discuss the mole concept;
- determine the percent of elements in compounds;
- determine the formula of a compound;
- discuss the kinds and types of chemical reactions;
- analyze the techniques in balancing chemical reactions and
- determine the limiting reagent/reactant of a chemical reaction.

ACTIVITY 1

Form a group and perform the following task.

1. List some chemical processes that occur in your daily life. Identify the reactants and products in each of these chemical processes.
2. 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 findings to the class.

6.1 FUNDAMENTAL LAWS OF CHEMICAL REACTIONS

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

- The law of conservation of mass
- The law of definite proportions
- The law of multiple proportions

The Law of conservation of mass

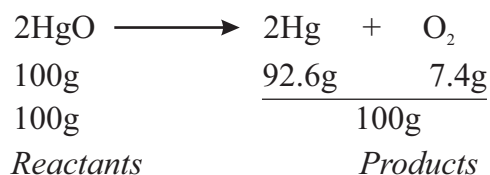
In 1774, the French chemist Antoine Lavoisier performed an experiment in which he heated a sealed glass container that held a sample of tin in air. He measured the mass of the substances before and after heating and found them to be the same. This and other similar experimental observations became the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

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 and shown as:

$$\text{Mass of reactants} = \text{Mass of products.}$$

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:



Experiment 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 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 color 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.

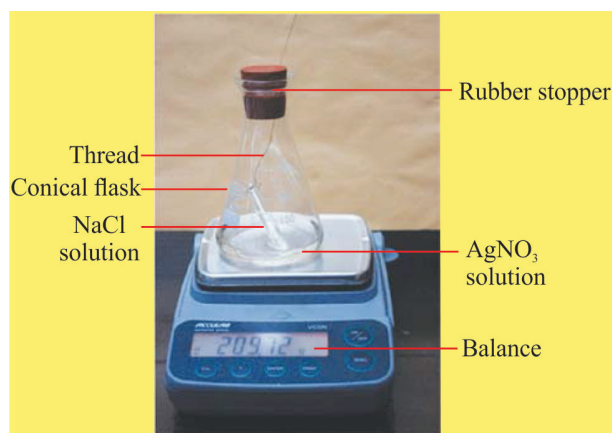


Figure 1. Conservation of mass.

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 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 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.

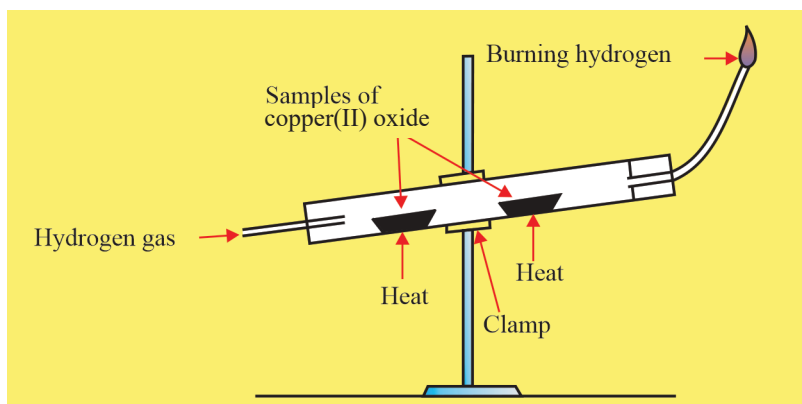


Figure 2. Reduction of copper(II) oxide by hydrogen

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 (CO₂). In CO₂, 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.

ACTIVITY 2

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 oxide
Dinitrogen monoxide	N ₂ O		
Nitrogen monoxide	NO		
Dinitrogen trioxide	NO ₃		
Nitrogen dioxide	NO ₂		
Dinitrogen pentoxide	N ₂ O ₅		

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.

$$\frac{2.67\text{g of oxygen in CO}_2}{1.33\text{g of oxygen in CO}} = \frac{2}{1} = 2:1$$

Exercise

1. Iron and chlorine form two compounds, A and B. Compound A contains 1.27g 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.
2. 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.

6.2 MOLECULES AND IONS

Molecules

A molecule is an aggregate of at least two atoms in a definite arrangement held together by chemical forces (also called chemical bonds). A molecule may contain atoms of the same element or atoms of two or more elements joined in a fixed ratio, in accordance with the law of definite proportions. Thus, a molecule is not necessarily a compound, which, by definition, is made up of two or more elements. Hydrogen gas, for example, is a pure element, but it consists of molecules made up of two H atoms each. Water, on the other hand, is a molecular compound that contains hydrogen and oxygen in a ratio of two H atoms and one O atom. Like atoms, molecules are electrically neutral.

The hydrogen molecule, symbolized as H₂, is called a diatomic molecule because it contains only two atoms. Other elements that normally exist as diatomic molecules are nitrogen (N₂) and oxygen (O₂), as well as the Group 7A elements—fluorine (F₂), chlorine (Cl₂), bromine (Br₂), and iodine (I₂). Of course, a diatomic molecule can contain atoms of different elements. Examples are hydrogen chloride (HCl) and carbon monoxide (CO).

The vast majority of molecules contain more than two atoms. They can be atoms of the same element, as in ozone (O₃), which is made up of three atoms of oxygen, or they can be combinations of two or more different elements. Molecules containing

more than two atoms are called polyatomic molecules. Like ozone (O_3), water (H_2O) and ammonia (NH_3) are polyatomic molecules.

Ions

An ion is an atom or a group of atoms that has a net positive or negative charge. The number of positively charged protons in the nucleus of an atom remains the same during ordinary chemical changes (called chemical reactions), but negatively charged electrons may be lost or gained. The loss of one or more electrons from a neutral atom results in a cation, an ion with a net positive charge.

ACTIVITY 3

In units you have learnt about the formation of ionic bond.

In your group discuss on the following questions.

1. What is the difference between atoms and ions?
2. What are cations and anions.
3. Explain how sodium chloride is formed from Na and Cl.

6.3 THE MOLE CONCEPTS/STOICHIOMETRY

ACTIVITY 4

Form a group and discuss the following

- (a) A bicycle mechanic has 10 frames (body parts) and 16 wheels in the shop. How many complete bicycle can he assemble using these parts? Which parts of the bicycle are left over.
- (b) Based on your conclusion in (a), do you think that the masses of reactants are always completely converted to product 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.

- The composition of any substance in the chemical equation should be expressed by a definite formula.
- The law of conservation of mass must be obeyed (the mass of reactants equals the mass of products).

The Mole Concept

ACTIVITY 5

Form a group and discuss the following:

1. Suppose you have one million eggs. Would it be simpler for you to sell them by counting from one to million or by grouping them in boxes containing one hundred eggs each? Why?
2. Tablets are also sold in containers or boxes containing specific numbers. Why is that so?

We may group different items in dozens or grosses. Analogous to the grouping of items, particles like atoms and molecules are grouped in moles in chemistry. Since atoms and molecules are very small, scaling factor called mole is used to convert their small masses into masses in grams. The mole provides the basis for relating masses in grams to number of atoms, molecules or formula units.

The mole is the amount of a substance that contains the same number of particles as the number of atoms in exactly 12 g of carbon-12. 12 g of carbon-12 contains an Avogadro's number of carbon-12 atoms.

$$\text{Avogadro's number} = 6.022 \times 10^{23}$$

A mole of any substance is a group containing 6.022×10^{23} particles of that substance whether the particles are atoms, molecules or ions.

Thus, one mole of atoms contains 6.022×10^{23} atoms. One mole of molecules contains 6.022×10^{23} molecules. One mole of an ionic compound contains 6.022×10^{23} formula units. The mass of one mole of atoms, one mole of molecules and one mole of an ionic compound is equal to the atomic mass, molecular mass and formula mass expressed in grams, respectively.

For example:

- One mole of carbon atoms has a mass of 12 g.
- One mole of water molecules has a mass of 18 g.
- One mole of sodium chloride has a mass of 58.5 g

The mass of one mole of atoms, molecules or a compound is called molar mass. It is denoted by 'M' and expressed by the unit gram per mole (g/mol). It is possible to calculate the mass of moles of a substance in grams. This can be done by multiplying the number of moles by the molar mass.

Mass of the substance = Given number of moles \times Molar mass (g/mol).

Example

Calculate the mass of 2.5 mol of sodium atom (Atomic mass: Na = 23)

Solution

You can also proceed to calculate the mass as follows.

$$1 \text{ mol of sodium} = 23 \text{ g/mol}$$

$$2.5 \text{ mol of sodium} = x$$

Cross multiplying and solving for 'x' gives

$$x = 2.5 \text{ mol of sodium} \times 23 \text{ g/mol} = 57.5 \text{ g of sodium}$$

Example

What is the mass of 5 mol of sodium hydroxide, (NaOH)?
(Atomic mass: Na = 23, O = 16, H = 1)

Solution

First, calculate the molar mass of NaOH or formula mass of NaOH in grams.

$$\text{Molar mass of NaOH} = (1 \times 23) + (1 \times 16) + (1 \times 1) = 40 \text{ g/mol}$$

$$\text{Thus, mass in grams of 5 mol NaOH} = 5 \text{ mol} \times 40 \text{ g/mol} = 200 \text{ g}$$

It is possible to calculate the number of moles of a substance from a given mass or number of particles.

To calculate the number of moles of a substance from a given mass, we use the relation:

$$\text{Number of moles (n)} = \frac{\text{Given mass in grams}}{\text{Molar mass of the substance in grams per mole}} = \frac{m}{M}$$

In short, we can write this relation as:

$$n = m/M \quad \text{where } n = \text{number of moles}$$

$$m = \text{given mass}$$

$$M = \text{molar mass}$$

Example

How many moles are there in 81 g of aluminum, Al? Atomic mass: Al = 27

Solution

$$\text{Molar mass of Al (M)} = \text{Atomic mass of Al} = 27 \text{ g/mol.}$$

$$\text{Given mass of Al (m)} = 81 \text{ g}$$

$$\text{The number of mole of Al (n)} = \frac{\text{Given mass}}{\text{Molar mass}} = \frac{81\text{g}}{27\text{g.mol}} = 3 \text{ mol of Al}$$

Exercise

How many moles are contained in:

- (a) 20 g of CaCO_3 (Atomic mass: Ca = 40, C = 12, O = 16)
 (b) 147 g of H_2SO_4 ? (Atomic mass: H = 1, S = 32, O = 16)

The number of moles of a substance from a given number of particles is calculated using the relation:

$$\text{Number of moles (n)} = \frac{\text{Given number of particles}}{6.022 \times 10^{23} \text{ particles/mol}}$$

Example

How many moles are 9.033×10^{23} molecules of ammonia, NH_3 ?

Solution

Given number of molecules = 9.033×10^{23} NH_3 molecules

$$\begin{aligned} \text{Number of moles NH}_3 &= \frac{\text{Given number of particles}}{6.022 \times 10^{23} \text{ particles / mol}} \\ &= \frac{9.033 \times 10^{23} \text{ NH}_3 \text{ molecules}}{6.022 \times 10^{23} \text{ molecules / mol}} \\ &= 1.5 \text{ mol of NH}_3 \end{aligned}$$

You can also proceed as follows:

$$1 \text{ mol} = 6.022 \times 10^{23} \text{ molecules}$$

$$x = 9.033 \times 10^{23} \text{ molecules}$$

Cross multiplying and solving for 'x' gives you 1.5 mol again.

Generally, when the number of particles (atoms, molecules or formula units) is given, simply divide the given number of particles by Avogadro's number to calculate the number of moles.

To calculate the number of particles in a given mass (amount) of a substance, first calculate the number of moles of the substance and then multiply by Avogadro's number.

$$\text{Number of particles} = \frac{m}{M} \times 6.022 \times 10^{23} \text{ particles / mol}$$

$$\text{Number of particles} = \frac{\text{Given mass}}{\text{Molar mass}} \times 6.022 \times 10^{23} \text{ particles / mol}$$

Example

What is the number of molecules of water in 27 g of water? (Atomic mass: H = 1, O = 16)

Solution

Molar mass of $\text{H}_2\text{O} = (2 \times 1) + (1 \times 16) = 18 \text{ g/mol}$.

Given mass of $\text{H}_2\text{O} = 27 \text{ g}$

To calculate the number of molecules, we proceed as follows:

$$\begin{aligned} \text{Number of molecules of water} &= \frac{\text{Given mass}}{\text{Molar mass}} \times 6.022 \times 10^{23} \text{ particles/mol} \\ &= \frac{27\text{g}}{18\text{g/mol}} \times 6.022 \times 10^{23} \text{ molecules of water /mol} \\ &= 9.0333 \times 10^{23} \text{ molecules of water} \end{aligned}$$

Exercise

- Calculate the mass of:
 - 0.25 mol of $\text{Al}_2(\text{SO}_4)_3$
 - 0.8 mol of KOH
 - 3 mol of MgCl_2
 - 0.4 mol NaCl
 - 1.2 mol NaOH
 - 2.5 mol $\text{C}_6\text{H}_{12}\text{O}_6$
- How many moles are there in:
 - 18.5 g $\text{Ca}(\text{OH})_2$
 - 1.5055×10^{23} molecules of H_2
 - 250 g CaCO_3
 - 7.5275×10^{23} atoms of Nitrogen
 - 9.033×10^{24} molecules of CO_2
 - 157.5 g HNO_3

- (g) 1.8066×10^{24} molecules of CH_4
(h) 2.4088×10^{23} atoms of sodium
3. Calculate the number of atoms in:
- (a) 14 g of Li
(b) 1.5 mol of carbon
(c) 20 g of calcium
4. Calculate the number of molecules in:
- (a) 2 mol of CO_2
(b) 51 g of NH_3
(c) 33 g of CO_2
(d) 5 mol of H_2
(e) 2 mol of SO_2
(f) 24.5 g of H_2SO_4
5. Calculate the number of formula units present in:
- (a) 60 g NaOH
(b) 2 mol of KCl
6. How many moles are equivalent to:
- (a) 1.8066×10^{25} molecules of CO_2
(b) 1.5055×10^{24} atoms of sodium
(c) 9.033×10^{22} formula units of NaCl
(d) 2.4088×10^{24} atoms of chlorine
(e) 5.4198×10^{24} molecules of H_2
(f) 3.6132×10^{25} atoms of carbon.

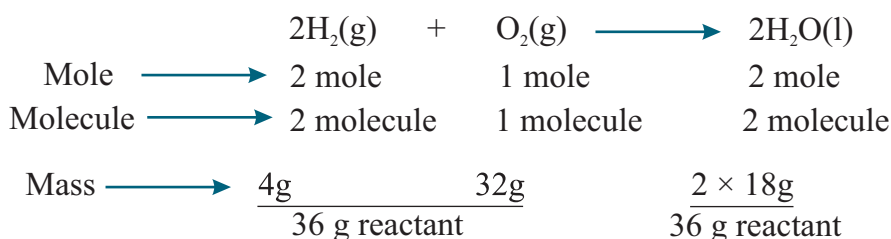
Calculations based on formulas

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:



Calculations based on chemical equations (stoichiometric problems) are classified into mole–mole problem, mass–mass problems and mass–mole, mole atom, particle, molecule problems.

Mass–Mass problems

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:

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).

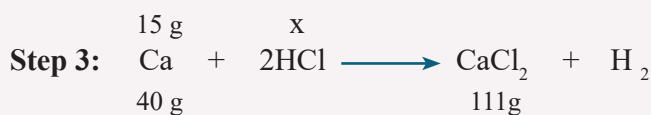
Step 4: Set up the proportion.

Step 5: Solve for the unknown mass, x.

Example

How many grams of calcium chloride are formed when 15 g of calcium metal reacts with hydrochloric acid?

Solution



$$\text{Step 4: } \frac{15\text{g}}{40\text{g}} = \frac{x}{111\text{g}}$$

$$\text{Step 5: } x = 41.63 \text{ g}$$

Therefore 41.63 g of CaCl_2 is produced

Example

How many grams of oxygen are produced by the decomposition of 145 grams of potassium chlorate?

Solution:



$$\text{Step 4: } \frac{145\text{g}}{245\text{g}} = \frac{x}{96\text{g}}$$

$$\text{Step 5: } x = 56.8\text{g}$$

Mole–Mole problems

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.

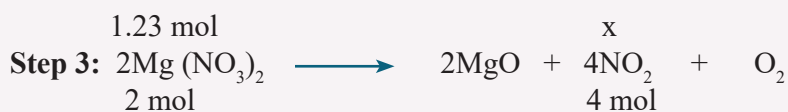
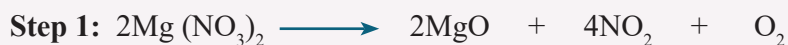
Step 3: Place the coefficients as the number of moles under the formula of each substance involved.

Step 4: Set up the proportion.

Step 5: Solve for the unknown value, x ; and convert the moles obtained into mass.

Example

What mol of nitrogen dioxide is produced by the decomposition of 1.23 mol of magnesium nitrate?

Solution

$$\text{Step 4: } \frac{1.23 \text{ mol}}{2 \text{ mol}} = \frac{x}{4 \text{ mol}}$$

$$\text{Step 5: } x = 2.46 \text{ moles of NO}_2$$

Mole–mass problems

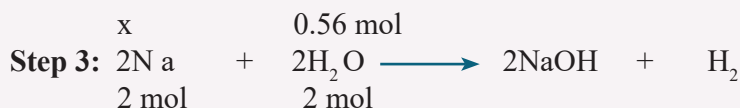
In such kind of problems either grams of one substance is given and moles of the other substance is asked or moles of one substance is given and mass of another substance is going to be calculated, or both can be given in mass and asked to convert them into moles.

Example

How many grams of sodium metal are needed to react with 10.0 g of water?

Solution

$$\text{Step 2: } \text{moles of H}_2\text{O} = \frac{\text{given mass}}{\text{Molar mass}} = \frac{10.0\text{g}}{18\text{g}} = 0.56 \text{ mol}$$



$$\text{Step 4: } \begin{array}{ccc} x & = & 0.56\text{mol} \\ 2 \text{ mol} & & 2 \text{ mol} \end{array}$$

$$\text{Step 5: } x = 0.56 \text{ mol of Na}$$

Since you are asked grams of sodium metal produced, convert 0.56g of Na to grams

$$\text{Mass of Na} = \text{mole} \times \text{molar mass} = 0.56 \text{ mol} \times 23\text{g/mol} = 12.88\text{g}$$

Exercise

- How many grams of CaCO_3 are needed to react with 15.2 g of HCl in according to the following equation?



- How many grams of NaOH are needed to neutralize 50 grams of H_2SO_4 ?
- Calculate the mass of CaCl_2 formed when 5 moles of chlorine reacts with calcium metal.
- How many moles of H_2O are required to produce 4.5 moles of HNO_3 according to the following reaction:



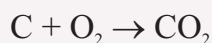
- In the decomposition of KClO_3 , how many moles of KCl are formed in the reaction that produces 0.05 moles of O_2 ?
- How many moles of CaO are needed to react with excess water to produce 370g of calcium hydroxide?

Mole-atom, particle, molecule problems

In this type of problem mole of substance is given and atoms of the other substance is going to be calculated. In this case we apply the concept of Avogadro's number.

Example

How many atoms of Carbon are reacted with one mole of oxygen molecule to form carbon dioxide.

Solution

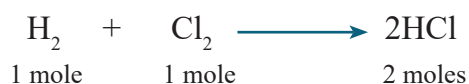
$$\frac{\begin{array}{c} x \\ \text{C} \\ \hline 1\text{mol} \\ (6.02 \times 10^{23} \\ \text{atom}) \end{array}}{6.02 \times 10^{23} \text{ atom}} + \frac{\begin{array}{c} 1\text{mol} \\ \text{O}_2 \\ \hline 1\text{mol} \end{array}}{1\text{mol}} \rightarrow \text{CO}_2$$

$$x = 6.02 \times 10^{23} \text{ atoms of carbon}$$

Limiting and excess reagents

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:

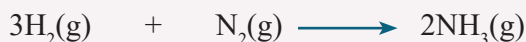


According to the equation, 1 mole of H_2 reacts with 1 mole of chlorine to produce 2 mole of HCl . Thus, all the reactants are completely consumed and only products appear. However if 1 mole of H_2 reacts with 1.5 mole of Cl_2 , there is insufficient H_2 to react with all of the Cl_2 . Therefore, Cl_2 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

How much ammonia is produced if 10 g of hydrogen reacts with 18 g of nitrogen?

Solution



First determine the number of moles;

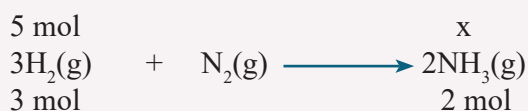
$$\text{Moles of H}_2 = \frac{10\text{g}}{2\text{g/mol}} = 5 \text{ mol}$$

$$\text{Moles of N}_2 = \frac{18\text{g}}{28\text{g/mol}} = 0.64 \text{ mol}$$

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 quantity of H_2 in calculation

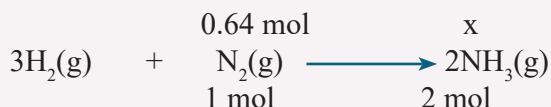


$$\frac{5 \text{ mol}}{3 \text{ mol}} = \frac{x}{2 \text{ mol}}$$

$$x = 3.33 \text{ mol NH}_3$$

$$\text{Mass of NH}_3 = 3.33 \text{ mol} \times 17 \text{ g/mol} = 56.6 \text{ g}$$

(ii) Using the quantity of N₂ in calculation

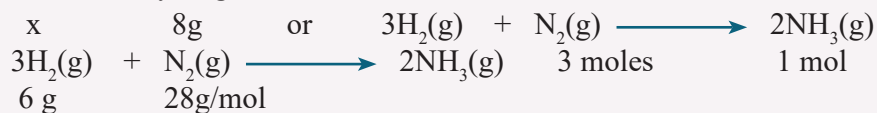


$$0.64 \text{ mol} = \frac{x}{2 \text{ mol}}$$

$$x = 1.28 \text{ mol NH}_3$$

$$\text{Mass of NH}_3 = 1.28 \text{ mol} \times 17 \text{ g/mol} = 21.8 \text{ g}$$

Therefore, the limiting reactant is nitrogen, because it gives less amount of NH₃, i.e., 21.8 g NH₃. In the reaction, 0.64 mole (18 g) of N₂ is consumed. Hydrogen is in excess. The amount of hydrogen consumed will be x 0.64 moles



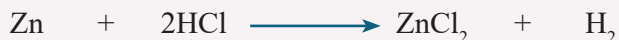
$$\frac{x}{6} = \frac{18}{28} = 3.86 \text{ g of H}_2$$

$$\frac{x}{3 \text{ mol}} = \frac{0.64 \text{ moles}}{1 \text{ mol}} = 1.92 \text{ mol of H}_2$$

Therefore, 3.86 g or 1.92 moles of H₂ is used in the reaction, and 6.14 g or 3.08 moles of H₂ is left unreacted.

Example

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 ZnCl₂ will be produced?

Solution

$$\text{Moles of Zn} = \frac{9.5 \text{ g}}{65 \text{ g/mol}} = 0.15 \text{ mol}$$

$$\text{Moles of HCl} = \frac{9.5 \text{ g}}{36.5 \text{ g/mol}} = 0.26 \text{ mol}$$

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 in calculation



$$\frac{0.15 \text{ mol}}{1 \text{ mol}} = \frac{x}{1 \text{ mol}} = 0.15 \text{ mol ZnCl}_2$$

(ii) Using the quantity of HCl in calculation



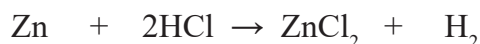
$$\frac{0.26 \text{ mol}}{2 \text{ mol}} = \frac{x}{1 \text{ mol}} = 0.13 \text{ mol ZnCl}_2$$

Hence, the limiting reactant is HCl

$$\text{Mass of ZnCl}_2 = 0.13 \text{ mol} \times 136 \text{ g/mol} = 17.68 \text{ g ZnCl}_2$$

Exercise

1. If 6.5 g of zinc reacts with 5.0 g of HCl, according to the following reaction.



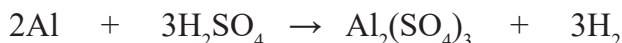
- Which substance is the limiting reactant?
 - How many grams of the reactant remains unreacted?
 - How many grams of hydrogen would be produced?
- What mass of Na_2SO_4 is produced if 49 g of H_2SO_4 reacts with 80 g of NaOH?
 - If 20 g of CaCO_3 and 25 g of HCl are mixed, what mass of CO_2 is produced?



4. If 3 moles of calcium reacts with 3 moles of oxygen, then:

- which substance is the limiting reactant?
- how many moles of calcium oxide are formed?

5. For the reaction:



How many grams of hydrogen are produced if 0.8 mole of aluminum reacts with 1.0 mole of sulfuric acid?

6.4 PERCENTAGE COMPOSITION OF COMPOUNDS

Percentage composition of a compound is the number of parts by mass of each element in one hundred parts by mass of the compound. The percent by mass of each element in a compound is calculated using the relation:

$$\% \text{ composition} = \frac{\text{Mass of element in the compound}}{\text{Formula (molar) mass of a compound}} \times 100$$

To determine the percentage composition of a compound, we use the following steps.

- Identify the elements and the number of atoms of each kind present in the compound.
- Calculate the total mass of each element by multiplying its number in the formula of the compound with the respective molar mass.
- Calculate the molar mass of the compound.
- Calculate the percentage composition of the compound by dividing the total mass of each element by the molar mass of the compound and then multiply by 100.

Example

Calculate the percentage composition of magnesium carbonate, MgCO_3 . (Atomic mass: $\text{Mg} = 24$, $\text{C} = 12$, $\text{O} = 16$)

Solution

Step 1: Elements: $\text{Mg} = 1$, $\text{C} = 1$, $\text{O} = 3$

Step 2: Mass: $\text{Mg} = 1 \times 24 = 24 \text{ g}$, $\text{C} = 1 \times 12 = 12 \text{ g}$, $\text{O} = 3 \times 16 = 48 \text{ g}$

Step 3: Molar mass of $\text{MgCO}_3 = 24 \text{ g} + 12 \text{ g} + 48 \text{ g} = 84 \text{ g}$

$$\text{Step 4: } \% \text{ of Mg} = \frac{24\text{g}}{84} \times 100 = 28.57\%$$

$$\% \text{ of C} = \frac{12\text{g}}{84} \times 100 = 14.29\%$$

$$\% \text{ of O} = \frac{48\text{g}}{84} \times 100 = 57.14\%$$

The sum of the percentages of elements in a compound is expected to be 100%. However, due to errors resulting from rounding off, the sum of the percentages of different elements in a compound may sometimes not add up to exactly 100%.

Exercise

- Calculate percentage composition by mass of the constituent elements for each of the following compounds:
 - $\text{Ca}(\text{NO}_3)_2$
 - Na_2CO_3
 - NaOH
 - N_2O_5
 - $\text{Fe}_2(\text{SO}_4)_3$
 - $\text{C}_9\text{H}_8\text{O}_4$
- Calculate the percentage composition by mass of nitrogen in each of the following fertilisers.
 - DAP, $(\text{NH}_4)_2\text{HPO}_4$
 - Urea, $(\text{NH}_2)_2\text{CO}$
 - Ammonium nitrate, NH_4NO_3
 - Ammonium sulphate, $(\text{NH}_4)_2\text{SO}_4$
- What is the percentage composition by mass of oxygen in each of the following?
 - PbO
 - Fe_2O_3
 - Al_2O_3
 - MnO_2
 - MgO
 - K_2O

6.5 DETERMINATION OF FORMULAS

ACTIVITY 6

Discuss the following in groups. When you complete, present your group's opinion to the rest of the class.

The formula of glucose is $C_6H_{12}O_6$. If we simplify the three subscripts by six, we get another formula; CH_2O .

- What do the two formulas indicate?
- What do we call such type of formulas?

Molecular formula is the formula that shows the actual number of atoms of each type of element present in a compound.

Empirical Formula is the formula that shows the number of atoms of each type of element present in the compound in the lowest possible ratio. It is the simplest formula of the compound. Table 1 shows molecular formulas and empirical formulas of some compounds.

Table 1 Names, molecular formulas and empirical formulas of some compounds.

Name of compound	Molecular formula	Empirical formula
Ethene	C_2H_4	CH_2
Butene	C_4H_8	CH_2
Hexene	C_6H_{12}	CH_2

The compounds, ethene, butene and hexene are different compounds. Their molecular formulas are different but have the same empirical formula. The empirical formula CH_2 , shows that each of these compounds consists of carbon atoms and hydrogen atoms in the ratio 1:2.

Sometimes a compound has the same empirical formula and molecular formula. For example, the empirical formula and molecular formula of ethanol is C_2H_6O . The empirical formula and molecular formula of sulfuric acid is H_2SO_4 .

Determining an empirical formula of a compound

To determine the empirical formula of a compound from its composition expressed in percentage by mass of each element, we use the following steps.

- Write the symbols of the constituent elements.
- Write the percentage by mass of each element below its chemical symbol.
- Divide the percentage by mass of the element by the respective atomic mass to get the number of moles of each element.

4. Divide the numbers of moles of all element by the smallest number of mole obtained in step 3 to get their mole ratio.
5. If any number obtained in step 4 is not a whole number multiply all numbers by the appropriate integer to make that number a whole number.
6. Write the symbols of the constituent elements side by side and the numbers obtained as subscripts to get its empirical formula.

Example

What is the empirical formula of the compound composed of 92.3% carbon and 7.7% hydrogen? (Atomic mass: C = 12, H = 1).

Solution

The percentage by mass of an element can be considered as its mass in hundred grams of the compound.

Step 1: C	H
Step 2: 92.3 g	7.7 g
Step 3: $\frac{92.3\text{g}}{12}$	$\frac{7.7\text{g}}{1}$
= 7.7 mol	= 7.7mol
Step 4: $\frac{7.7}{7.7}$	$\frac{7.7}{7.7}$
= 1	= 1

Step 5: Is not necessary because the numbers are whole numbers.

Step 6: Empirical formula of the compound is CH. It is not important to write the subscript 1.

The empirical formula, CH, tell only the simplest ratio of carbon to hydrogen is 1:1 and it tells nothing about the molecular formula of the compound.

Example

What is the empirical formula of the compound composed of 52.94% aluminum and 47.06% oxygen? (Atomic mass: Al = 27, O = 16)

Solution:

Step 1: Al	O
Step 2: 52.94 g	47.06 g

$$\text{Step 3: } \frac{52.93\text{g}}{27\text{g/mol}} \quad \frac{47.06\text{g}}{16\text{g/mol}}$$

$$= 1.96\text{ mol} \quad = 2.94\text{ mol}$$

$$\text{Step 4: } \frac{1.96}{1.96} \quad \frac{2.94}{1.96}$$

$$= 1 \quad = 1.5 = 3/2$$

Step 5: Since one of the numbers is not a whole number, multiply both numbers by 2

$$1 \times 2 = 2$$

$$3/2 \times 2 = 3$$

Step 6: The empirical formula of the compound is Al_2O_3 .

Al_2O_3 is the empirical formula and molecular formula of aluminum oxide.

Determining molecular formula of a compound

As explained above, some compounds have the same empirical formula and molecular formula. However, there are compounds having different empirical formula and molecular formula. In case of these compounds, the molecular formula is a whole number multiple of the empirical formula. In addition to that, molecular mass of the compound is also a multiple of the empirical formula mass. That is:

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

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

Where 'n' is a number greater or equal to 1.

It is possible to determine the molecular formula of a compound from its percentage composition and the molecular mass of the compound using the following steps.

1. Determine the empirical formula of the compound.
2. Calculate the empirical formula mass.
3. Divide the molecular mass by the empirical formula mass to get the number that should multiply the empirical formula.
4. Multiply the empirical formula by the number obtained when the molecular mass is divided by the empirical formula mass.

Example

What is the molecular formula of a compound having percentage composition of 85.71% carbon and 14.29% hydrogen whose molecular mass is 140? (Atomic mass: C = 12, H = 1) First, determine the empirical formula from the given percentage composition using the steps applied in examples given to determine empirical formula.

Solution

Step 1: C	H
Step 2: 85.71 g	4.29 g
Step 3: $\frac{85.71\text{g}}{12\text{g/mol}}$	$\frac{4.29\text{g}}{1\text{g/mol}}$
= 7.145 mol	= 4.29 mol

Step 4: The smaller number is 4.29. So, divide by 4.29

$\frac{7.145}{4.29}$	$\frac{4.29}{4.29}$
= 1	= 1

Step 5: Not important in this case.

Step 6: The empirical formula of the compound is CH_2 .

After determining the empirical formula, use the steps followed to get the molecular formula from empirical formula and molecular mass.

Step 1: Empirical formula is CH_2 .

Step 2: Empirical formula mass = $(1 \times 12) + (2 \times 1) = 14$

Step 3: $n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{140}{14} = 10$

Step 4: Molecular formula = Empirical formula $\times n$

$$= (\text{CH}_2) \times 10$$

$$= \text{C}_{10}\text{H}_{20}$$

So, the molecular formula of the compound is $\text{C}_{10}\text{H}_{20}$

Example

If the empirical formula of a compound is CH_2 and its molecular mass is 70, what would be the molecular formula of the compound?

Solution

Step 1: Empirical formula is CH_2 .

Step 2: Empirical formula mass is 14.

Step 3: $n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{70}{14} = 5$

$$\begin{aligned}
 \text{Step 4: } \text{Molecular formula} &= \text{Empirical formula} \times n \\
 &= (\text{CH}_2) \times 5 \\
 &= \text{C}_5\text{H}_{10}
 \end{aligned}$$

Exercise

- Determine the molecular formula of a compound having a simplest formula of CH and formula mass of 78.
- A compound with a molecular mass of 42 is found to contain 85.64% carbon and 14.36% hydrogen by mass. Find its molecular formula.
- Analysis of a 10.15 g sample of a compound known to contain only phosphorus and oxygen is found to contain 5.717 g of oxygen. What is the simplest formula of this compound?
- What is the empirical formula of a compound found to contain 32.38% sodium, 22.65% sulfur and 44.49% oxygen?

6.6 CHEMICAL REACTIONS

Chemical reactions are the processes by which substances interact to form new substances with different compositions. Chemical reactions are constantly occurring in the world around us; everything from the rusting of an iron fence to the metabolic pathways of a human cell are all examples of chemical reactions. Chemistry is an attempt to classify and better understand these reactions.

A chemical reaction is typically represented by a chemical equation, which represents the change from reactants to products. The left hand side of the equation represents the reactants, while the right hand side represents the products. A typical chemical reaction is written with stoichiometric coefficients, which show the relative amounts of products and reactants involved in the reaction. Each compound is followed by a parenthetical note of the compound's state: (*l*) for liquid, (*s*) for solid, (*g*) for gas. The symbol (*aq*) is also commonly used in order to represent an aqueous solution, in which compounds are dissolved in water.

A reaction might take the following form:



In the above example, A and B, known as the reactants, reacted to form C and D, the products.

Kinds of chemical reactions

Irreversible reactions

Some chemical reactions can occur in only one direction. These reactions are called irreversible reactions. The reactants can change to the products, but the products cannot change back to the reactants. These reactions are like making a cake. The ingredients of a cake—such as eggs and flour—are the reactants. They are mixed together and baked to form the cake, which is the product. The cake can't be “unbaked” and “unmixed” to change it back to the raw eggs, flour, and other ingredients. So making a cake is irreversible.

Combustion reactions are generally irreversible. Combustion occurs whenever a fuel burns. In this type of reaction, the fuel may combine with oxygen (in the air) and produces carbon dioxide and water vapor. The chemical equation for a combustion reaction is:

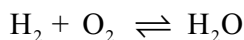


In a complete combustion reaction, fuel and oxygen are the reactants and the products are carbon dioxide and water. These two products cannot react to reform the fuel and oxygen, so the reaction is irreversible. The one-way arrow in the equation shows that the reaction can go in only one direction.

Reversible reactions

Many chemical reactions can occur in both directions. These reactions are called reversible reactions. Not only can the reactants change to the products, but the products can change back to the reactants, at least under certain conditions.

In a chemical equation, a reversible reaction is represented with two arrows, one pointing in each direction. This shows that the reaction can go both ways. How would you represent the reversible reaction in which hydrogen and oxygen gases combine to produce water?



In a reversible reaction, both forward and reverse directions of the reaction generally occur at the same time. While reactants are reacting to produce products, products are reacting to produce reactants. Often, a point is reached at which forward and reverse directions of the reaction occur at the same rate. When this happens, there is no overall change in the amount of reactants and products, even though the reactions keep occurring in both directions. This point is called *equilibrium*. The

term *equilibrium* means “state of balance,” and it is used to refer to a state of balance between any opposing changes.

Types of Chemical Reactions

ACTIVITY 7

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.

Combination Reactions

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.



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.

Experiment 3

Investigation of Combination Reaction.

Objective: To investigate the reaction between sulfur and iron.

Apparatus: Test tube, stand, Bunsen burner, watch glass.

Chemicals: Sulfur powder, iron filings.

Procedure:

1. Mix about 3 g of iron filings and 2 g of powdered sulfur 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 3.
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 colors of iron filings and sulfur before the reaction?
2. What was the color of the resulting compound after the reaction?

3. Write a balanced chemical equation for the reaction.
4. Identify the type of reaction.

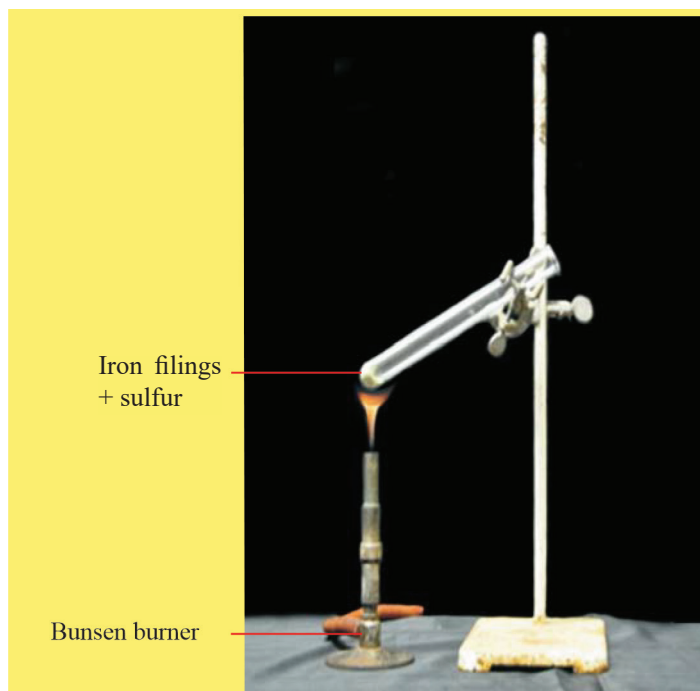
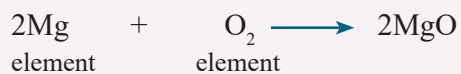


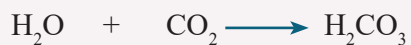
Figure 3. The reaction between iron and sulfur

Example

- Magnesium burns in oxygen to form magnesium oxide.



- Water and carbon dioxide combine to form carbonic acid.



Exercise

Complete and balance the following combination reactions.

1. $\text{CO} + \text{O}_2 \rightarrow$
2. $\text{H}_2\text{O} + \text{SO}_2 \rightarrow$
3. $\text{Na}_2\text{O} + \text{H}_2\text{O} \rightarrow$
4. $\text{CaO} + \text{CO}_2 \rightarrow$

Decomposition reactions

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:



where the reactant AB must be a compound and the products A and B could be elements or compounds.

Experiment 4

Investigation of Decomposition Reaction

Objective: To investigate the decomposition of copper(II) carbonate.

Apparatus: Test tube, stand, Bunsen burner, cork, delivery tube.

Chemicals: Copper(II) carbonate and lime water.

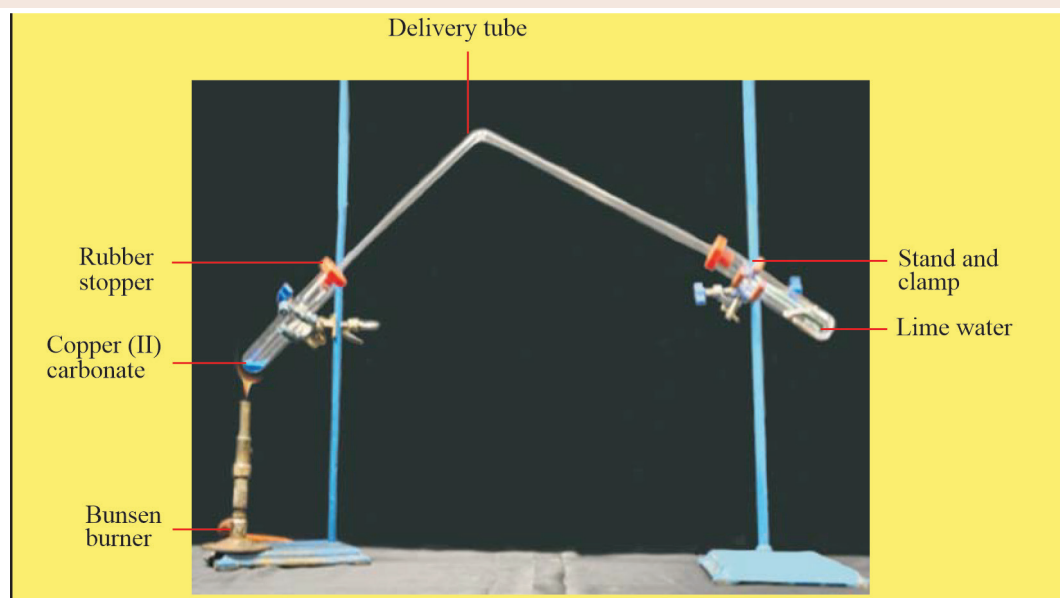


Figure 4. Decomposition of copper(II) carbonate.

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. Fit a cork and a delivery tube to the test tube.

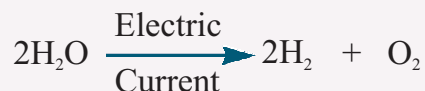
Put another test tube containing lime water at the end of the delivery tube. Heat the copper(II) carbonate with a burner.

Observations and analysis:

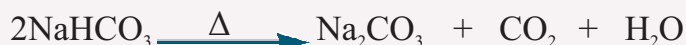
1. What was the color of copper(II) carbonate before heating?
2. What was the color during heating and after cooling?
3. What change did you observe in the lime water?
4. Write a balanced chemical equation for the reaction.

Example

- Water is decomposed to hydrogen and oxygen gases when electricity is passed through it.



- When sodium bicarbonate is heated, it decomposes to give sodium carbonate, carbon dioxide, and water.



Let us consider the decompositions of nitrates and carbonates:

1. Decomposition of metallic nitrates

(a) Decomposition of group IA nitrates produces nitrites and oxygen



(b) Decomposition of all metal nitrates, except group IA metals, gives nitrogen dioxide, metal oxide and oxygen gas.

Example



2. Decomposition of metallic carbonates

All metal carbonates, except sodium and potassium, decompose when heated to form the metal oxide and carbon dioxide.



Single displacement reactions

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.



If A is a metal, it will displace B to form AC, provided A is a more active metal than B.



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.

Experiment 5

Investigation of Single Displacement Reaction

Objective: To investigate the displacement reaction between iron and copper(II) sulfate.

Apparatus: Iron rod and beaker

Chemicals: Copper(II) sulfate.

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) sulfate solution as shown in Figure 5 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

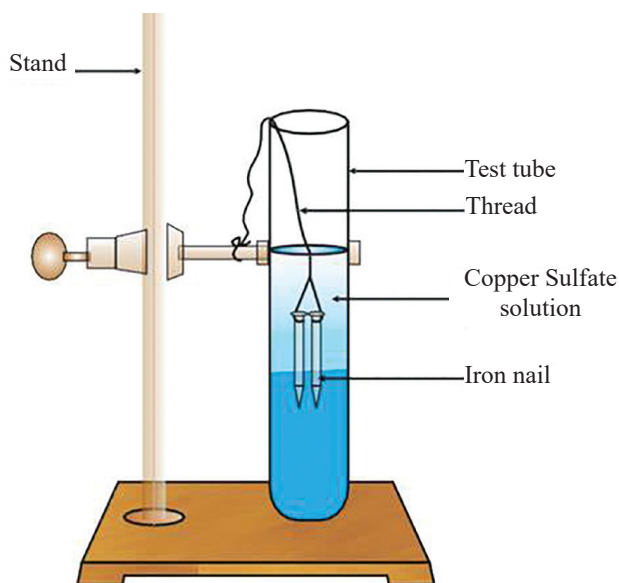


Figure 5. Reaction between iron and copper(II) sulfate

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.



- (b) Reactive metals, such as potassium, calcium, and sodium react vigorously with water to displace hydrogen:



Exercise

Complete and balance the following single displacement reactions:

- $\text{Zn} + \text{CuSO}_4 \rightarrow$
- $\text{Cu} + \text{Zn}(\text{NO}_3)_2 \rightarrow$
- $\text{F}_2 + \text{CaCl}_2 \rightarrow$
- $\text{Br}_2 + \text{NaCl} \rightarrow$

Double Displacement Reactions

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.

**Experiment 6**

Investigation of Double Displacement Reaction

Objective: To observe the displacement reaction between Na_2SO_4 and $\text{Ba}(\text{NO}_3)_2$.

Apparatus: Beaker, stirrer, filter paper, filter funnel.

Chemicals: Na_2SO_4 and $\text{Ba}(\text{NO}_3)_2$.

Procedure:

- Take solution of $\text{Ba}(\text{NO}_3)_2$ into a beaker and add dropwise Na_2SO_4 solution. Then stir it continuously.
- Filter the precipitate using a filter paper and funnel. Collect the filtrate or the solution in a clean beaker.

Observations and analysis:

- Write the names of the compounds that are formed as a precipitate and as solution at the end of the reaction.
- What was the color of the precipitate.
- Write the balanced chemical equation for the reaction.

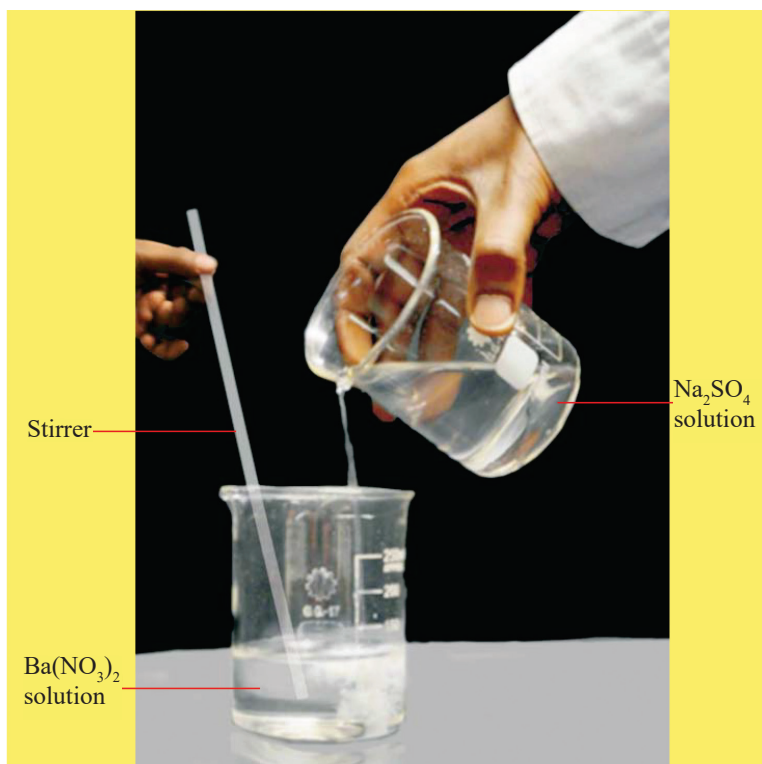


Figure 6. The double displacement reaction between Na_2SO_4 and $\text{Ba}(\text{NO}_3)_2$.

Example

- The two soluble compounds AgNO_3 and NaCl react to produce an insoluble precipitate of AgCl and a soluble NaNO_3 solution.



- When aqueous solutions of BaCl_2 and Na_2SO_4 react, a precipitate of BaSO_4 is formed.



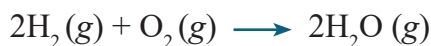
Exercise

- What type of reaction does usually take place in each of the following reactions?
 - a metal reacting with water.
 - 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) $\text{FeO} + \text{C} \rightarrow \text{Fe} + \text{CO}$
 - (b) $2\text{NH}_3 + \text{H}_2\text{SO}_4 \rightarrow (\text{NH}_4)_2\text{SO}_4$
 - (c) $\text{CaCO}_3 + 2\text{HCl} \rightarrow \text{CaCl}_2 + \text{CO}_2 + \text{H}_2\text{O}$
 - (d) $2\text{Cu}(\text{NO}_3)_2 \rightarrow 2\text{CuO} + 4\text{NO}_2 + \text{O}_2$
 - (e) $2\text{Na}_3\text{PO}_4 + 3\text{Ca}(\text{OH})_2 \rightarrow \text{Ca}_3(\text{PO}_4)_2 + 6\text{NaOH}$
 - (f) $\text{CuSO}_4 \cdot 5\text{H}_2\text{O} \rightarrow \text{CuSO}_4 + 5\text{H}_2\text{O}$

Combustion reactions

A combustion reaction is a reaction in which a substance reacts with oxygen gas, releasing energy in the form of light and heat. Combustion reactions must involve O_2 as one reactant. The combustion of hydrogen gas produces water vapor:



Notice that this reaction also qualifies as a combination reaction.

Many combustion reactions occur with a hydrocarbon, a compound made up solely of carbon and hydrogen. The products of the combustion of hydrocarbons are carbon dioxide and water. Many hydrocarbons are used as fuel because their combustion releases very large amounts of heat energy. Propane (C_3H_8) is a gaseous hydrocarbon that is commonly used as the fuel source in gas grills.



Oxidation-reduction reactions

An oxidation–reduction or redox reaction is a reaction that involves the transfer of electrons between chemical species (the atoms, ions, or molecules involved in the reaction).

Redox reactions are all around us: the burning of fuels, the corrosion of metals, and even the processes of photosynthesis and cellular respiration involve oxidation and reduction. Detail explanation on oxidation reduction reactions will be given in the next unit.

Writing and balancing chemical equations

Even though chemical compounds are broken up and new compounds are formed during a chemical reaction, atoms in the reactants do not disappear, nor do new atoms appear to form the products. In chemical reactions, atoms are never created or destroyed. The same atoms that were present in the reactants are present in the products—they are merely reorganized into different arrangements. In a complete chemical equation, the two sides of the equation must be present on the reactant and the product sides of the 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:



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.



3. Balance the equation.



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.

Balancing chemical equation

ACTIVITY 8

Form a group and discuss each of the following

- What is the difference between chemical reaction and chemical equation?
- 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

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

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

Solution

Step 1: Magnesium + Oxygen \longrightarrow Magnesium oxide

Step 2: $\text{Mg} + \text{O}_2 \longrightarrow \text{MgO}$ (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.



- 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.



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

**Exercise**

Balance the following chemical equation, using the inspection method:

- $\text{Na} + \text{H}_2\text{O} \rightarrow \text{NaOH} + \text{H}_2$
- $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
- $\text{H}_2\text{O}_2 \rightarrow \text{H}_2\text{O} + \text{O}_2$
- $\text{Al} + \text{H}_3\text{PO}_4 \rightarrow \text{AlPO}_4 + \text{H}_2$
- $\text{HNO}_3 + \text{H}_2\text{S} \rightarrow \text{NO} + \text{S} + \text{H}_2$

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

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

Solution

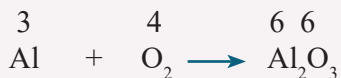
Step 1: Represent the reaction by a word equation.



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



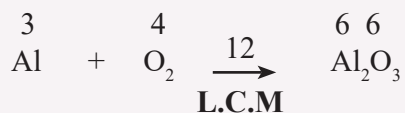
Step 3: Place the total valency of each atom above it.



Now the equation shows

- The valency of aluminum as 3.
- The total valency of oxygen is $2 \times 2 = 4$.
- The total valency of aluminum 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.



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.



Checking: There are 4 aluminum and 6 oxygen atoms on both sides of the equation. Hence, the

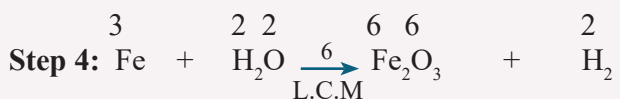
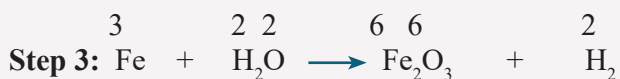
Chemical equation is correctly balanced.

Example

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

Solution

Step 1: Iron + water \longrightarrow Iron(III) oxide + hydrogen



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

Exercise

- Write the balanced chemical equation to represent the following reactions.
 - Sulfur dioxide reacts with oxygen to produce sulfur trioxide.
 - Potassium chlorate when heated produces potassium chloride and oxygen.
 - Sodium carbonate reacts with hydrochloric acid to form water, carbon dioxide and sodium chloride.
 - Silver oxide decomposes to silver and oxygen gas.
- Balance the following equations by the LCM method.
 - $\text{PCl}_5 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{PO}_4 + \text{HCl}$
 - $\text{Mg} + \text{H}_2\text{O} \rightarrow \text{Mg}(\text{OH})_2 + \text{H}_2$
 - $\text{Zn}(\text{NO}_3)_2 \rightarrow \text{ZnO} + \text{NO}_2 + \text{O}_2$
 - $\text{H}_2\text{SO}_4 + \text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2\text{O}$
 - $\text{NH}_3 + \text{O}_2 \rightarrow \text{NO} + \text{H}_2\text{O}$

KEY TERMS

- Chemical reaction
- Combination reaction
- Decomposition reaction
- Double displacement reaction
- Empirical formula
- Irreversible reaction
- Law of conservation of mass
- Law of definite proportion
- Law of multiple proportion
- Mole
- Molecular formula
- Oxidation reduction reaction
- Percentage composition
- Reversible reaction

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.
- 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.
- Stoichiometry is the quantitative relationship between reactants and products.
- Mass-mass problems, mass-mole and mole – mole problems are the main types of stoichiometric calculations.
- The mole is the amount of a substance that contains the same number of particles as the number of atoms in exactly 12 g of carbon -12.
- A mole is Avogadro's number (6.022×10^{23}) of particles (atoms, molecules or ions).
- Molar mass is the mass of one mole of atoms, molecules or ionic compounds and expressed by the unit gram per mole (g/mol).
- Percentage composition of a compound is the number of parts by mass of each element in one hundred parts by mass of the compound.
- Molecular formula is the formula that shows the actual number of atoms of each element present in a compound.
- Empirical formula is the formula that shows the number of atoms of each element present in the compound in the lowest possible ratio.
- Some compounds have the same empirical and molecular formula.
- In case of some compounds, the molecular formula is a whole number multiple of the empirical formula.

Exercises**Part I. True-False Type Questions**

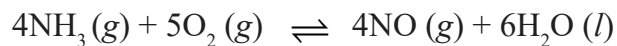
1. In any chemical reaction, each type of atoms is conserved.
2. In a balanced chemical equation, both sides of the equation have the same number of moles.
3. One mole of any substance contains 6.022×10^{23} particles of that substance.
4. Empirical formula shows the actual number of atoms of each kind of element present in the compound.

5. Molecular formula is a short hand representation of molecule of an element or a compound.
6. One mole of different substances have the same mass.
7. The molecular formula and empirical formula of any compound are identical.
8. The atom that serves as a standard in determining relative atomic mass is carbon-12 atom.
9. Atomic mass unit is used as a unit to describe the molar mass of a substance.
10. The sum of the percentages of all elements in a compound is nearly 100%.

Part II. Give short answers for each of the following questions.

11. What is the mass, in grams, of each of the following ?
 - (a) 4.25 mol Na
 - (b) 6.5 mol Cu
 - (c) 3.6 mol CO₂
 - (d) 7.2 mol Au
12. How many moles of atoms and number of atoms are there in:
 - (a) 150 g S
 - (b) 100 g Ca
 - (c) 48 g O
 - (d) 140 g Fe
13. How many moles of atoms are equivalent to:
 - (a) 6.022×10^{23} atoms Ne
 - (b) 3.011×10^{23} atoms Mg
 - (c) 2.25×10^{25} atoms Zn
14. What is the mass, in grams, of each of the following?
 - (a) 3.011×10^{24} atoms F
 - (b) 8.42×10^{23} atoms Br
 - (c) 1.505×10^{24} atoms Mg
 - (d) 6.022×10^{25} atoms Cl
15. How many grams and moles contain:
 - (a) 3.011×10^{24} atoms CO₂
 - (b) 9.033×10^{22} molecules N₂
16. Find the simplest formula of a compound that contains 52.2% carbon, 13.0% hydrogen and 34.8% oxygen.
17. A compound is composed of 53.33% carbon, 11.11% hydrogen and 35.56% oxygen. If the molecular mass of the compound is 90, what is the molecular formula of this compound?

18. How many grams of oxygen can be prepared by the decomposition of 12 grams of mercury(II) oxide?
19. 25 g of NH_3 is mixed with 4 moles of O_2 in the given reaction:



- (a) Which is the limiting reactant?
- (b) What mass of NO is formed?
- (c) What mass of H_2O is formed?



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CHAPTER

7

OXIDATION-REDUCTION REACTIONS

Chapter Contents

- 7.1 Definition of Oxidation and Reduction
- 7.2 Oxidation Number or Oxidation State
- 7.3 Oxidizing and Reducing Agents
- 7.4 Balancing Simple Oxidation-Reduction (Redox) Equations
 - Summary
 - Exercises

Chapter Outcomes

Upon completion of this topic, learners will:

- discuss the terms oxidation and reduction;
- discuss the difference between oxidizing and reducing agents;
- apply the concept of calculating oxidation numbers and;
- apply the rules for balancing redox reactions.

ACTIVITY 1

Form a group and discuss the following phenomenon:

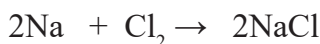
When you dry your meal dishes with a towel, the towel can be termed as a drying agent and the dish as wetting agent. What happens to the towel and the dish, in terms of getting wet and dry, after cleaning? Relate this idea with oxidizing agent and reducing agent, oxidation and reduction of substances. Present your conclusion to the class.

7.1 DEFINITION OF OXIDATION AND REDUCTION

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

Oxidation

The term oxidation was originally used to describe reactions in which an element combines with oxygen. For example, the reaction between magnesium and oxygen involves the oxidation of magnesium: $2\text{Mg(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{MgO(s)}$. However, nowadays it has a broader meaning that includes reactions not involving only oxygen. Now, oxidation is the loss of electrons and is identified by an increase in oxidation number. For example, in the reaction



Each sodium atom has lost one electron and has turned to a sodium ion. Hence, sodium is oxidized.

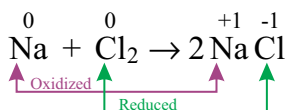
**Reduction**

The term reduction was originally used to describe the production of a metal from an ore. The term has been modified through time to include other reactions. Now, reduction is the gain of electrons and is identified by a decrease in oxidation number. For example, in the reaction between sodium and chlorine each chlorine atom has gained an electron and has changed to chloride ion. Thus, chlorine is reduced;



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.

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 Cl₂ = 0, O in O₃ = 0, S in S₈ = 0.

Rule 2: The oxidation number of a monatomic ion is equal to the charge on the ion. Example: Na⁺ = +1, Mg²⁺ = +2, S²⁻ = -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:

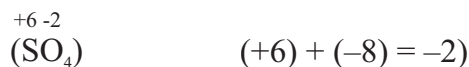
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|--------------------------------|--|
| (i) peroxides is -1. | Example: Na ₂ O ₂ |
| (ii) superoxides is -1/2. | Example: KO ₂ |
| (iii) oxygen difluoride is +2. | Example: OF ₂ |

Rule 4: The oxidation number of hydrogen in its entire compounds is +1 except in metal hydrides, (like NaH, CaH₂ and AlH₃), where its oxidation number is -1.

Rule 5: The sum of the oxidation number of all the atoms in a neutral compound is zero.



Rule 6: In a polyatomic ion, the sum of the oxidation numbers of the constituent atoms equals the charge on the ion, for example:



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

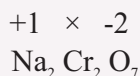
What is the oxidation number of chromium in $\text{Na}_2\text{Cr}_2\text{O}_7$?

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 .



Since the sum of the oxidation numbers of Na, Cr, and O in $\text{Na}_2\text{Cr}_2\text{O}_7$ is 0 (Rule 5)

$$\begin{aligned} \overset{+1}{\text{Na}}_2 \overset{x}{\text{Cr}}_2 \overset{-2}{\text{O}}_7 \\ (1 \times 2) + (x \times 2) + (-2 \times 7) = 0 \\ 2 + 2x - 14 = 0 \\ x = +6 \end{aligned}$$

Therefore, the oxidation number of Cr in $\text{Na}_2\text{Cr}_2\text{O}_7$ is $+6$.

Example

What is the oxidation number of manganese in MnO_4^- ?

Solution

Let the oxidation number of Mn be x .

$$\begin{array}{c} x - 2 \\ (\text{MnO}_4) \end{array}$$

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 $+7$.

Exercise

- Determine the oxidation number of the specified element in each of the following:
 - C in $\text{H}_2\text{C}_2\text{O}_4$
 - N in NH_4F
 - S in $\text{Na}_2\text{S}_4\text{O}_6$
 - P in $\text{Ca}_3(\text{PO}_4)_2$
 - H in AlH_3
 - N in NH_4HCO_3
 - Fe in $\text{K}_4[\text{Fe}(\text{CN})_6]$
- Determine the oxidation number of the specified element in each of the following.
 - S in S^{-2}
 - Cl in ClO_3^-
 - N in NH_4^+
 - P in PO_4^{3-}
 - Cr in Cr_2O_7^-
 - S in $\text{S}_2\text{O}_8^{2-}$
- Determine whether the following processes are oxidation or reduction reactions:
 - $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$
 - $\text{K} \rightarrow \text{K}^+ + e^-$
 - $\text{O} + 2e^- \rightarrow \text{O}^{2-}$

- (d) $\text{S}^{2-} \rightarrow \text{S} + 2e^{-}$
 (e) $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + e^{-}$
 (f) $\text{N} + 3e^{-} \rightarrow \text{N}^{3-}$

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

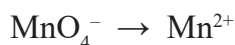
Reducing agents are substances that:

- are oxidized (lose electrons)
- contain elements whose oxidation number increases

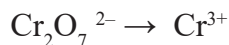
Tests for an oxidizing agent are accomplished by mixing it with a substance which is easily oxidized to give a visible color change when the reaction takes place.

For example,

- (i) Permanganate ion (MnO_4^{-}) in acidic solution changes color from purple to colorless.



- (ii) Dichromate in acidic solution changes color from orange to green.



Other common oxidizing agents are chlorine, potassium chromate, sodium chlorate and manganese(IV) oxide.

Similarly, certain reducing agents undergo a visible color 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, $2\text{I}^{-} \rightarrow \text{I}_2$. 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.

7.4 BALANCING SIMPLE OXIDATION-REDUCTION (REDOX) EQUATIONS

Because atoms are neither created nor destroyed in an ordinary chemical reaction, chemical equation must have an equal number of atoms of each element on the reactant and product sides. In addition, the net electrical charges in the reactant side must be equal to the net electrical charges in the product side.

Balancing redox reactions using change in oxidation number method

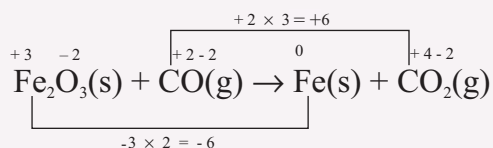
In a redox reaction, the total number of electrons gained by the oxidizing agent is equal to the total number of electrons lost by the reducing agent. The change in oxidation number method for balancing redox reaction is based on this fact. This method is used to balance redox reactions that do not occur in aqueous solutions, and also reactions that do not involve ions.

Use the following steps to balance chemical equations, using the change in oxidation state method

Step 1: Write an unbalanced chemical equation.

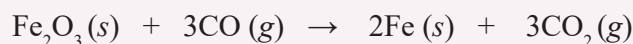
Step 2: Assign oxidation numbers to each atom, and determine which atoms are changing their oxidation numbers.

Step 3: Determine the number of electrons lost by each atom per molecule undergoing oxidation and the number of electrons gained by each atom per molecule being reduced.



Step 5: Balance the remaining substances by counting atoms.

Check the balancing for both atoms and charge. Occasionally, a coefficient may need to be placed in front of a molecular formula that was not involved in the redox process. In the current example, the equation is now balanced.



Step 6: Check the final equation to be sure that each atom as well as the net charges on either side is balanced.

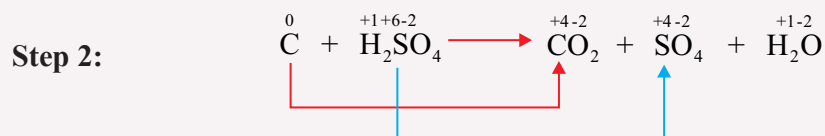
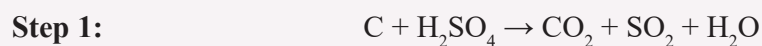
Reactants	Products
Fe (2)	Fe (2)
C (3)	C (3)
O (6)	O (6)

Example

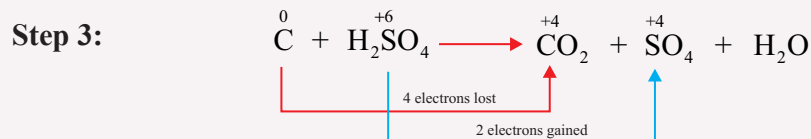
Balance the following chemical equation, using the change in oxidation-number method.

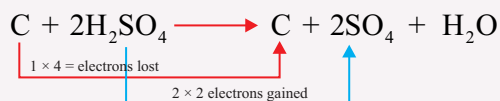


Solution



Changes in oxidation number are observed in carbon and sulfur.



Step 4:**Step 5:** Balance hydrogen atoms by making the coefficient of water 2.**Step 6:**

Reactants	Products
C (1)	C (1)
H (4)	H (4)
S (2)	S (2)
O (8)	O (8)

The chemical equation is balanced, because the number of atoms in the reactant side is equal to the number of atoms in the product side. It can also be confirmed that the net charge is zero on each side.

Exercise

Balance the following equations using change in Oxidation numbers method.

- $\text{Na}_2\text{Cr}_2\text{O}_7 + \text{FeCl}_3 + \text{HCl} \rightarrow \text{CrCl}_3 + \text{NaCl} + \text{FeCl}_2 + \text{H}_2\text{O}$
- $\text{Cu} + \text{HNO}_3 \rightarrow \text{Cu}(\text{NO}_3)_2 + \text{NO}_2 + \text{H}_2\text{O}$
- $\text{H}_2\text{S} + \text{HNO}_3 \rightarrow \text{S} + \text{NO} + \text{H}_2\text{O}$
- $\text{Cu} + \text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{SO}_2 + \text{H}_2\text{O}$
- $\text{P}_4 + \text{HNO}_3 \rightarrow \text{H}_3\text{PO}_4 + \text{NO}_2 + \text{H}_2\text{O}$

Ion-electron method for balancing redox reactions

In the ion-electron method (also called the half-reaction method), the redox equation is separated into two half-equations - one for oxidation and one for reduction. Each of these half-reactions is balanced separately and then combined to give the balanced redox equation.

Steps for balancing redox reaction using the ion-electron method:

Step 1: Identify the element undergoing oxidation and reduction; write separate equations for oxidation and reduction half-reactions.

Step 2: Balance the element oxidized or reduced.

Step 3: Determine the change in oxidation number of the element oxidized or reduced. Add the electrons gained or lost to the side on which the redox element has the higher oxidation number.

Step 4: Balance any elements other than hydrogen and oxygen that may be present.

Step 5: Balance charges by adding H^+ ions if the solution is acidic and OH^- ions if the solution is basic.

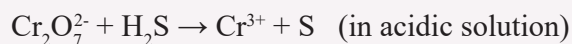
Step 6: Balance hydrogen by adding water molecules.

Step 7: Multiply each half-reaction by appropriate coefficients so that the number of electrons lost in oxidation equals the number of electrons gained in reduction.

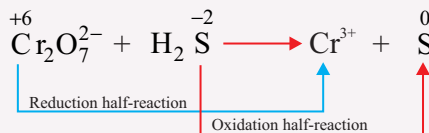
Step 8: Add the two half-reactions together and eliminate anything that appears in identical form on both sides of the equation.

Example

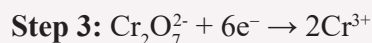
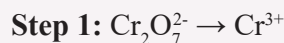
Balance the following chemical equation using the ion-electron method.



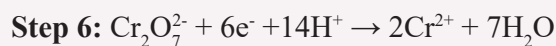
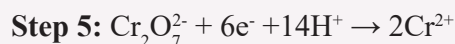
Solution



Reduction half-reaction



Step 4: There is no element other than oxygen.



Here, the reduction half-reaction is balanced.

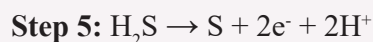
Oxidation half-reaction



Step 2: Sulfur is already balanced.

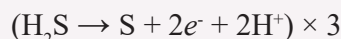


Step 4: There is no element other than hydrogen.

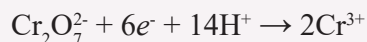


Step 6: Hydrogen is already balanced and the oxidation half-reaction is also balanced.

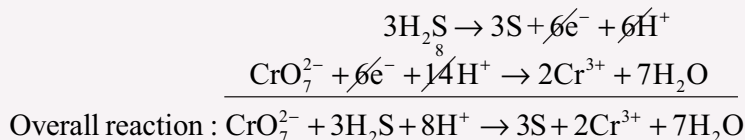
Step 7: Oxidation half-reaction



Reduction half-reaction



Step 8: Add the two half-reactions..



Exercise

Balance the following chemical equations, using the ion-electron method.

- $\text{MnO}_4^- + \text{SO}_3^{2-} \rightarrow \text{Mn}^{2+} + \text{SO}_4^{2-}$ (in acidic solution)
- $\text{Sn}^{2+} + \text{Bi}^{3+} \rightarrow [\text{Sn}(\text{OH})_6]^{2-} + \text{Bi}$ (in basic solution)

SUMMARY

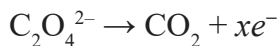
- Oxidation is loss of electrons, and reduction is gain of electrons.
- A reduced substance is an oxidizing agent and an oxidized substance is a reducing agent.
- Redox reactions are balanced, using the oxidation number change method or the ion- electron method.
- Oxidizing agents are the substances reduced, and reducing agents are the substances oxidized.

Exercises

Part I. Choose the best answer from the given alternatives.

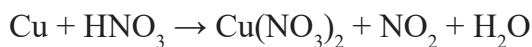
- What is the oxidation number of chromium in $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$?
 - +7
 - +3
 - +5
 - +6

2. In the following half-reaction,



The value of 'x' is:

- (a) 1
 - (b) 2
 - (c) 4
 - (d) 3
3. Balance the following equation, using the lowest possible whole number coefficients;



The sum of the coefficients is:

- (a) 8
 - (b) 12
 - (c) 10
 - (d) 7
- II. Give short answer to the following questions**

4. Define the following terms:

- (a) oxidation
- (b) reduction
- (c) oxidizing agent
- (d) reducing agent

5. In each of the following equations, identify the substance oxidized, the substance reduced, the oxidizing agent and reducing agent.

- (a) $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$
- (b) $\text{CuO} + \text{H}_2 \rightarrow \text{Cu} + \text{H}_2\text{O}$
- (c) $\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} + \text{NaNO}_3$
- (d) $3\text{Cu} + 8\text{HNO}_3 \rightarrow 3\text{Cu}(\text{NO}_3)_2 + 2\text{NO} + 4\text{H}_2\text{O}$
- (e) $\text{SO}_4^{2-} + \text{I}^- + \text{H}^+ \rightarrow \text{H}_2\text{S} + \text{I}_2 + \text{H}_2\text{O}$

Health Related Caution

What are the ways to avoid dengue and malaria fever?

- Time your outings.
- Reduce mosquito habitat.
- Sleep under mosquito-net.
- Put screens on windows and doors.
- Keep your house airy and well-lit.
- Do not let water stagnate anywhere.
- Wear long pants and long sleeves to cover your body.
- Apply mosquito repellent with DEET (diethyltoluamide) to exposed skin.
- Treat clothing, mosquito nets, tents, sleeping bags and other fabrics with an insect repellent called permethrin.



How can a person reduce the risk of getting HIV?

- Get tested for HIV.
- Do not inject drugs.
- Choose less risky sexual behaviors.
- Use condoms every time you have sex.
- Limit your number of sexual partners.
- Get tested and treated for STDs.
- Talk to your health care provider about pre-exposure prophylaxis (PrEP).

WHAT IS BULLYING?

Any unwanted written, verbal, graphic, or physical act by an individual or group toward another person(s) that causes harm or distress.

Types of Bullying

- Physical
- Verbal
- Social
- Emotional
- Cyber

Signs of Bullying

- Headaches
- Depression
- Loss of friends
- School absenteeism
- Academic problems



What You Can Do

PREVENT

- Be a role model for positive communication, healthy relationships, and self-care.
- Reinforce acts of kindness, respect, and inclusion.
- Set policies and rules about bullying.

RECOGNIZE

- Know the definition of bullying and its many forms.
- Talk with and actively listen to the youth who confide in you.
- Watch for warning signs of bullying.

INTERVENE

- If you witness bullying behavior
- Respond quickly and consistently to send the message that it is not acceptable.
- Separate the students involved.
- Meet any immediate medical or mental health needs.
- Stay calm and model respectful behavior.



Source: Teacher's Diary on *Cyber-Crime Awareness* by UNODC, Cybercrime and MoE, Republic of Liberia

WHAT IS CYBERCRIME?

Cybercrime is criminal activity that either targets or uses a computer, a computer network or a networked device. Most cybercrime is committed by cybercriminals or hackers who want to make money or take advantage of a person.



Types of Cybercrime

- Email and internet fraud.
- Identity fraud (where personal information is stolen and used).
- Theft of financial or card payment data.
- Theft and sale of corporate data.
- Cyber extortion (demanding money to prevent a threatened attack).
- Ransomware attacks (a type of cyberextortion).
- Cryptojacking (where hackers mine cryptocurrency using resources they do not own).
- Cyberespionage (where hackers access government or company data).
- Interfering with systems in a way that compromises a network.
- Infringing copyright.
- Illegal gambling.
- Selling illegal items online.
- Soliciting, producing, or possessing child pornography.

How to Prevent Cyber Crimes?

- Enforce concrete security and keep it up-to-date.
- Never give out personal information to a stranger.
- Check security settings to prevent cybercrime.
- Using an antivirus software helps to recognize any threat or malware before it infects the computer system.
- When visiting unauthorized websites, keep your information secure.
- Restriction on access to your most valuable data.
- Backup all data, system, and considerations.
- Don't use free USB sticks.



Source: Teacher's Diary on *Cyber-Crime Awareness* by UNODC, Cybercrime and MoE, Republic of Liberia

