

CHEMISTRY

Distance Module-I

GRADE 11

CHE.

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6.94 Li Lithium	4 9.01 Be Beryllium											5 10.81 B Boron	6 12.01 C Carbon	7 14.01 N Nitrogen	8 15.999 O Oxygen	9 18.998 F Fluorine	10 20.18 Ne Neon	
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Cs Cesium	56 137.33 Ba Barium	57 138.91 La Lanthanum	72 178.49 Hf Hafnium	73 180.95 Ta Tantalum	74 183.85 W Tungsten	75 186.21 Re Rhenium	76 190.20 Os Osmium	77 192.22 Ir Iridium	78 195.09 Pt Platinum	79 196.97 Au Gold	80 200.59 Hg Mercury	81 204.37 Tl Thallium	82 207.19 Pb Lead	83 208.98 Bi Bismuth	84 (209) Po Polonium	85 (210) At Astatine	86 (222) Rn Radon	
87 (223) Fr Francium	88 226,03 Ra Radium	89 227,03 Ac Actinium	104 (261) Rf Rutherfordium	105 (262) Db Dubnium	106 (266) Sg Seaborgium	107 (262) Bh Bohrium	108 (255) Hs Hassium	109 (266) Mt Meitnerium	110 (271)	111 (272)	112 (277)	(113)	(114)	(115)	116 (289)	(117)	(118) (293)	l t
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Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Cerium	Praseodymium	Neodymium	Promethium	Samarium	Europium	Gadolinium	Terbium	Dysprosium	Holmium	Erbium	Thulium	Ytterbium	Lutetium
90 232.04	91 231.04	92 238.03	93 237.05	94 (244)	95 (243)	96 (247)	97 (247)	98 (251)	99 (252)	100 (257)	101 (258)	102 (259)	103 (262)
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr
Thorium	Protactinium	Uranium	Neptunium	Plutonium	Americium	Curium	Berkelium	Californium	Einsteinium	Fermium	Mendelevium	Nobelium	Lawrencium



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CHEMISTRY

Distance Module-I

GRADE 11



FEDERAL DEMOCRATIC REPUBLIC OF ETHIOPIA | MINSTRY OF EDUCATION



GRADE 11

ATOMIC STRUCTURE, CHEMICAL BONDING, AND PHYSICAL STATES OF MATTER

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Introduction to the Module

Dear Learner! In your lower grades, you learned about the basic concept of chemistry. You have been introduced about matter, physical and chemical properties, chemical bonding, the periodic table of elements, organic and inorganic compounds, etc. In grade 11 chemistry you will learn in details the structure of atoms, chemical bonding, physical states of matter, chemical kinetics, chemical equilibrium, and oxygen-containing organic compounds.

As a student you are enrolled in an independent study, you will take a dual role as a student and a teacher. As a student, you are responsible for mastering the lessons and completing the learning activities and assignments. As a teacher, you are responsible for checking your work carefully, noting areas in which you need to improve and motivating yourself to succeed.

Grade 11 Chemistry is divided into two modules:

Module 1 deals with the Atomic Structure, Chemical Bonding, and Physical States of matter

Module 2 deals with Chemical kinetics, equilibrium, and some important oxygen-containing organic compounds

For each module, an assignment for submission is provided. After you go through the contents of each module try to answer the questions. Submit the assignments to your tutor. Ask your tutor for the feedback and comments so that you can improve your study. The assignments are part of the total evaluation of the course.



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Module objectives

At the end of this module, you will be able to:

- discuss the historical development of atomic structure
- explain electromagnetic radiation, atomic spectra and Bohr models of the atom
- describe the quantum mechanical model of the atom and the related postulates and principles
- explain periodic law and how electronic configurations of atoms are related to the orbital diagrams
- state the formation and general properties of substances containing ionic, covalent and metallic bond
- discuss the formation and nature of hydrogen bonds, dipole-dipole forces and London forces
- explain how the Valence Shell Electron Pair Repulsion (VSEPR) model can be used to predict molecular shape
- discuss the kinetic molecular theory and properties of the three physical states of matter
- explain the properties of gases based on the variables such as volume, temperature, pressure and number of moles
- develop skills of solving problems using the gas laws.

Module Assessment Methods

- 1. You should complete regularly self-assessments and your progress will be evaluated.
- 2. Contact your tutor regularly if possible face-to-face in order to have oral feedback on your study progress for each section of a module.
- 3. You should submit individual assignment on time and try to get feedback on your progress
- After completion of module 1, you will sit for written examination which will be composed of true/false items, multiple choice, fill in the black space, and short answer questions.

UNIT

ATOMIC STRUCTURE AND PERIODIC PROPERTIES OF THE ELEMENTS

Unit Introduction

Dear learner, this unit deals with the structure and periodic properties of the elements. An atom is pictured as a nucleus of protons and neutrons surrounded by electrons. The electrons surrounding the nucleus are arranged in orbitals on the basis of their energies. The electron configuration represents the electronic structure of atoms. Periodic patterns in the ionization energies, physical properties structures of the elements and chemical properties are explained in terms of periodicity in their electronic configuration. Therefore, understanding of atomic structure is required to explain why the properties of the elements are regularly repeated when they arranged in order of increasing atomic number. This is what you will study in this unit.

Unit Outcome

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

- discuss the historical development of atomic structure
- explain the experimental observations and inferences made by some famous scientists to characterize the atom
- list the subatomic particles
- identify atomic mass and isotope terms
- explain electromagnetic radiation, atomic spectra and Bohr models of the atom
- compute the calculations involving atomic structure

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- describe the quantum mechanical model of the atom and the related postulates and principles
- demonstrate periodic law and how electronic configurations of atoms are related to the orbital diagrams and can explain periodic trends.

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Section 1.6: The Quantum Mechanical Model of the Atom
Section 1.7: Electronic configurations and orbital diagrams
Section 1.8: Electronic Configurations and the Periodic Table of the Elements

^{*j*} The Required Study Time

5 weeks

If you set an average of one hour per day, you will be able to complete unit one by the end of the assigned week. Try to do all the learning activities. And compare your answers which are attached at the end of the Module.

If you do not get a particular exercise right in the first attempt, you should not get discouraged, but instead, go back and attempt it over and over again. If you still do not get it right after several attempts, then you would rather seek help from your friend or even your tutor.

Unit Learning Strategies

- You should monitor his/her performance through planning, monitoring, and self-regulation
- You should manage the learning environment and available resources
- You should try all self-test exercises and written assignments
- Surf on the internet for better understanding of contents in the module.
- Consult your tutor for further clarifications on some contents in the module.

Section 1.1: Historical Development of the Atomic Nature of Substances

Dear learner, in this section you will study the historical development of atomic theory of matter.

Dear learner, atomic theory originated as a philosophical concept in ancient Greece. In this section you study how the beliefs about the structure of matter evolved and developed to the modern concept of the atom.

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

• compare the views of different philosophers on the nature of matter.

Most ancient Greece philosophers, including Plato and Aristotle, believed that matter is continuous, that is, matter continuously divisible into smaller and smaller pieces. However, Democritus (460–370 BC) disagreed on the continuous nature of matter and suggested that if you divided matter into smaller and smaller pieces, you would eventually end up with tiny, indestructible particles called atomos, or "atoms", meaning "indivisible". However, the ideas of Democritus were not widely accepted until 1808, John Dalton, developed an atomic theory that had gained broad acceptance.

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Can we see atoms with our naked eyes?

The size of an atom is so small that it is not possible to see it with naked eyes.

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Activity 1.1

Dear learner, can you describe the early developments leading to the modern concept of the atom?

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

l can...

- discuss the historical development of atomic structure
- compare the views of different philosophers on the nature of matter

Section 1.2: Dalton's Atomic Theory and the Modern Atomic Theory

Dear learner, in this section you will study about the development of the modern atomic theory and the laws that led to it. Dalton's atomic theory was the first complete attempt to describe all matter in terms of atoms and their properties.

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

- state postulates of Dalton's atomic theory
- state postulates of modern atomic theory
- state the laws of conservation of mass, definite proportions, multiple proportions and the basis of each of these laws
- use postulates of Dalton's atomic theory to explain the laws of definite and multiple proportions
- evaluate postulates of Dalton's and the modern atomic theories.

Scientific ideas usually develop based on previous scientific findings. The laws that are the basis for Dalton's atomic theory are the law of conservation of mass and the law of definite proportions.



What is the law that Dalton formulated based on the law of conservation of mass and the law of definite proportions?

The law of conservation of mass states that in a chemical reaction mass is neither created nor destroyed. If one able to measure the oxygen used by the fire to burn the wood and collects the products of combustion, the mass before and after combustion is exactly the same.

Postulates of Dalton's Atomic Theory

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

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

Activity 1.2

Dear learner, try to answer the following questions.

- The mass of a piece of wood before and after it is burnt to ashes is not the same. Does this show that mass is created or destroyed? Explain it.
- 2. Which of Dalton's postulates about atoms are inconsistent with later observations?

Postulates of Modern Atomic Theory

The Modern Atomic Theory can be summarized as follows:

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

Self-test Exercise 1.1

- 1. List the postulates of Dalton that continue to have significance (are retained in the modern atomic theory).
- 2. Describe the limitations of Dalton's atomic theory.
- 3. List the three fundamental laws of chemistry.

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

l can...

- state the laws of conservation of mass, definite proportions, multiple proportions and the basis of each of these laws
- use postulates of Dalton's atomic theory to explain the laws of definite and multiple proportions
- evaluate postulates of Dalton's and the modern atomic theories.

Section 1.3: Early experiments to characterize the atom

Dear learner, in this section you will study the contribution of early experiments conducted to understand the structure of an atom. You will learn about the discovery of electron and the nucleus. Radioactivity and the types of radioactive emissions will also be discussed.

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

- discuss the discovery of the electron
- describe the properties of cathode rays
- define the terms: radioactivity, radioactive decay and radio-isotope
- describe the common types of radioactive emissions
- illustrate the alpha scattering experiment
- summarize and interpret the major contribution of experiments of Thomson, Millikan and Rutherford concerning atomic structure.

) How were electrons discovered?

Dear learner, recognize that one of the first experiments on subatomic particles was carried out by the English physicist J. J. Thomson in 1897. Figure 1.1 shows an experimental apparatus similar to the one used by Thomson. In this apparatus, two electrodes from a high-voltage source are sealed into a glass tube from which the air has been evacuated. The negative electrode is called the cathode; the positive one, the anode. When the high-voltage current is turned on, the glass tube emits a greenish light. Experiments showed that this greenish light is caused by the interaction of the glass with cathode rays, which are rays that originate from the cathode. This beam bends away from the negatively charged plate and toward the positively charged plate.



What did Thomson conclude from this observation?

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Figure 1.1: Formation of cathode rays

 \sim Cathode rays bend when a magnet is brought toward them.

Thomson Showed that the characteristics of cathode rays are independent of the material making up the cathode. He concluded that a cathode ray consists of a beam of negatively charged particles (or electrons) and that electrons are constituents of all matter.

Which Dalton's postulate is rejected due to the discovery of electrons?

Thomson was able to calculate the ratio of the mass of an electron, m_e , to its charge, e (- 5.686 x 10⁻¹² kg C⁻¹ (kilograms per coulomb)). In 1909, Robert A. Millikan, an American physicist, measured the charge of the electron by measuring the effect of an electrical field on the rate at which charged oil drops fell under the influence of gravity. Based on careful experiments, Millikan established the charge on an electron as $e = -1.602 \times 10^{-19}$ C. He used this value and Thomson's mass/charge ratio to calculate an electron's mass to be 9.109 x 10⁻³¹ kg.

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

= 9.109 × 10⁻³¹ kg

1.3.1 Radioactivity and the Discovery of the Nucleus

Dear learner, three types of rays are released by radioactive substances (Figure 1.2). Alpha (α) rays consist of positively charged particles, called α particles. They have a mass of about four times that of a hydrogen atom and a charge twice the magnitude of an electron ; they are identical to helium nuclei. Beta (β) rays, or β particles, are electrons coming from inside the nucleus and are deflected by the negatively charged plate. The third type of radioactive radiation consists of high–energy rays called gamma (γ) rays. They have no charge and are not affected by an external electric or magnetic field



Figure 1.2: Three types of ray emitted by radioactive elements

Radioactivity or radioactive decay is the spontaneous emission of particles and/or radiation from the unstable nuclei of certain atoms such as uranium, radium, etc.

Activity 1.3 Dear learner, attempt to answer the following questions.

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1. Which rays are used to see whether bones are broken or not?

2. Does radio activity support Dalton's idea of atoms?

Discovery of the Nucleus

In1911, Ernest Rutherford conducted an experiment by bombarding a thin sheet of gold with α -particles and then studied the trajectory of these particles after their interaction with the gold foil (Figure 1.3).



Figure 1.3: (a) α -particles aimed at a piece of gold foil. (b) Magnified view of α -particles passing through and being deflected by nuclei

Rutherford made the following conclusions from the results of the experiment.

- i. Most of the mass and all the positive charge of an atom are concentrated in a minute central core of an atom called nucleus.
- ii. Most of the volume of an atom is empty space.

Rutherford correctly concluded that an atom has a positively charged nucleus that is very small in comparison to the size of the atom.

1.3.2 Discovery of the Neutron

Dear learner, did you know that neutron was discovered in 1932 by the British physicist James Chadwick? The neutron is a nuclear particle having a mass almost identical to that of the proton but with no electric charge.

The mass of a neutron, $m_n = 1.67493 \times 10^{-27}$ kg, which is about 1840 times the mass of an electron.

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CHECKLIST-1.3

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

l can...

- discuss the discovery of the electron
- describe the properties of cathode rays
- define the terms: radioactivity, radioactive decay and radio-isotope
- describe the common types of radioactive emissions
- summarize and interpret the major contribution of experiments of Thomson and Rutherford concerning atomic structure.

Section 1.4: Make-up of the nucleus

Dear learner, in this section you will study about the nucleus of an atom, the three fundamental subatomic particle, atomic mass and isotopes of an element. In order to study the structure of an atom, you need to know its component particles. An atom is made up of entities called subatomic particles.

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

- describe the make-up of the nucleus
- define atomic mass
- define isotope, and
- calculate the relative atomic mass (atomic mass) of naturally occurring isotopic elements.

1.4.1 Subatomic Particles

Dear learner, you have to know that a proton is a nuclear particle having a positive charge equal in magnitude to that of the electron. A proton has a mass of $m_p = 1.67262 \times 10^{-27}$ kg, which is about 1840 times the mass of electrons. **Table 1.1** compares the relative masses and charges of the three subatomic particles (note that "amu" stands for "atomic mass unit", which is equal to 1/12 the mass of an atom of carbon-12).

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⁷ The three important fundamental subatomic particles are: electrons, protons, and neutrons. The protons in a nucleus give the nucleus its positive charge.

Table 1.1: Properties of subatomic particles

Particle	Actual mass(kg)	Relative mass(amu)	Actual charge(C)	Relative charge	
Proton (p)	1.672622 x 10 ⁻²⁷	1.007276	1.602 x 10 ⁻¹⁹	+1	
Neutron(n)	1.674927 x 10 ⁻²⁷	1.008665	0	0	
Electron(e⁻)	9.109383 x 10 ⁻³¹	5.485799 x 10 ⁻⁴	-1.602 x 0 ⁻¹⁹	-1	

The atomic number (Z) of an element equals the number of protons in the nucleus of each of its atoms. All atoms of a particular element have the same atomic number, and each element has a different atomic number from that of any other element.

⁷ The total number of protons and neutrons in the nucleus of an atom is its mass number (A).

The mass number and atomic number are often written with the symbol, $\frac{A}{Z}X$.

Activity 1.4

Dear learner, attempt to explain the similarities and differences between the pair of atomic notations given below.

- a. ${}^{12}_{6}C$ and ${}^{14}_{6}C$
- b. ${}^{16}_{8}$ O and ${}^{14}_{7}$ N

1.4.2 Atomic Mass and Isotopes

All atoms of an element are identical in atomic number but not in mass number.

Isotopes of an element are atoms that have the same atomic number but different mass numbers.

For example, the element carbon has three naturally occurring isotopes: ¹²C, ¹³C, and ¹⁴C. Most elements found in nature are mixtures of isotopes.

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Why do isotopes of the same elements have similar chemical properties?

The average mass for the atoms in an element is called the atomic mass of the element.

Atomic mass can be obtained as averages over the relative masses of the isotopes of each element, weighted by their observed fractional abundances. If an element consists of *n* isotopes, of relative masses $A_1, A_2...A_n$ and fractional abundances of $f_1, f_2... f_n$ then the average relative atomic mass (A) of the element is:

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 $A = A_1 f_1 + A_2 f_2 + \dots + A_n f_n.$

Example 1.1

There are two naturally occurring isotopes of silver. Isotope ¹⁰⁷Ag (106.90509 amu) accounts for 51.84% of the total abundance, and isotope ¹⁰⁹Ag (108.90476) accounts for the remaining 48.16%. Calculate the atomic mass of silver?

Solution:

Find the portion of the atomic mass from each isotope:

Portion of atomic mass from ¹⁰⁷Ag: = isotopic mass x fractional abundance = 106.90509 amu x 0.5184 = 55.42 amu

Portion of atomic mass from ¹⁰⁹Ag: = 108.90476 amu x 0.4816 = 52.45 amu

Find the atomic mass of silver:

Atomic mass of Ag = 55.42 amu + 52.45 amu = 107.87 amu

🛠) Self-test Exercise 1.2

- 1. How many protons and neutrons are in the nucleus of each of the following atoms?
 - a. $^{27}_{13}$ Al c. $^{64}_{30}$ Zn
 - b. ${}^{32}_{16}S$ d. ${}^{207}_{82}Pb$
- Element X is toxic to humans in high concentration but essential to life at low concentrations. It has four naturally occurring isotopes that contain 24 protons. Identify element X and give the atomic symbol for the isotopes with 26, 28, 29, and 30 neutrons.
- 3. Naturally occurring boron consists of two isotopes, ¹⁰B and ¹¹B, with the isotopic masses 10.013 amu and 11.009 amu, respectively. The observed atomic mass of boron is 10.811 amu. Calculate the abundance of each isotope.
- 4. The two naturally occurring isotopes of lithium, lithium-6 and lithium-7, have masses of 6.01512 amu and 7.01600 amu, respectively. Which of these two occurs in greater abundance?

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

l can...

- describe the make-up of the nucleus
- define atomic mass and isotope
- calculate the relative atomic mass (atomic mass) of naturally occurring isotopic elements.

Section 1.5: Electromagnetic Radiation and Atomic Spectra

Dear learner, in this section you will study the characteristics of electromagnetic radiation and its relationship with properties of matter.

At	the	end	of t	his	secti	on.	vou	will	be	able	to:
						··· /					

- characterize electromagnetic radiation (EMR) in terms of wavelength, frequency and speed
- calculate the wavelength and frequency of EMR
- explain about light which has both wave and particle nature
- describe emission spectra of atoms as consisting of a series of lines
- define a photon as a unit of light energy
- distinguish how the photon theory explains the photoelectric effect and the relationship between a photon absorbed and an electron released
- state Bohr's assumption of energy of the electron in a hydrogen atom
- calculate the radius of electron orbit, the electron velocity and the energy of an electron using Bohr's model
- identify the line spectrum of hydrogen demonstrates the quantized nature of the energy of its electron
- explain that atoms emit or absorb energy when they undergo transitions from one state to another
- compose the limitations of Bohr's theory.

1.5.1 Electromagnetic Radiation

Dear learner, under this subsection you will learn about the electromagnetic radiation, and its primary characteristics (wavelength, frequency and speed).

Activity 1.5

Dear learner, before reading the note given under this section attempt the following questions.

- 1. What is the importance of electromagnetic radiation (EMR) in chemistry?
- 2. What are the common features of different energy sources?

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How does energy travel in space?

Electromagnetic radiation (EMR) is the emission and transmission of energy in the form of electromagnetic waves.

James Clerk Maxwell (1873) proposed that light consists of electromagnetic waves. According to his theory, an electromagnetic wave has an electric field and magnetic field components. These two components vibrate in two mutually perpendicular planes (Figure 1.4).





What are the common features of different energy sources? Dear learner, electromagnetic waves have three primary characteristics: wavelength, frequency and speed. Wavelength (λ , Greek letter lambda), is the distance the wave travels during one cycle (Figure 1.5). It is expressed in meter (m) and for very short wavelengths, in nanometer (nm), picometer (pm), or angstrom (Å). Frequency (v, Greek letter nu) is the number of cycles the wave undergoes per second and expressed in units of l/second (1/s; also called hertz, Hz).



Figure 1.5: Frequency of waves

Wave has a speed which depends on the type of wave and the nature of the medium through which it travels(for example, air, water, or a vacuum).

How do you express the relationship between the speed of light and its frequency?

The speed (c) of a wave is the product of its wavelength and its frequency:

 $c = \lambda v \tag{1.1}$

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In a vacuum, electromagnetic waves travel at 3 \times 10⁸ m/s, which is called the speed of light.

What is electromagnetic spectrum? What is the importance of the different regions of EMR?

Dear learner, EMR comes in a broad range of frequencies called the electromagnetic spectrum (Figure 1.6). A rainbow is an example of a continuous spectrum. Different wavelengths in visible light have different colors, from red (λ = 750 nm) to violet (λ = 380 nm). Radiation provides an important means of energy transfer. For instance, the energy from the Sun reaches the Earth mainly in the form of visible and ultraviolet radiation. In microwave ovens, microwave radiation heats water in foods, causing the food to cook quickly.



Figure 1. 6: The electromagnetic spectrum

Example 1.3

Ethiopian National Radio, The Addis Ababa radio station broadcasts a frequency of 2400 kHz with its AM signal. What is the wavelength of the radio wave expressed in meters?

Solution:

We obtain the wavelength of the radio wave by rearranging Equation1.1. So,

$$\lambda = \frac{c}{v} = \frac{3.00 \times 10^8 \frac{\text{m}}{\text{s}}}{2.4 \times 10^6 \text{ s}} = 125 \text{ m}$$

$\mathbf{\xi}$)Self-test Exercise 1.3

- 1. The most intense radiation emitted by the Earth has a wavelength of about 10.0 µm. What is the frequency of this radiation in hertz?
- 2. The Addis Ababa Fana FM radio station, broadcasts electromagnetic radiation at a frequency of 98.1 MHz. What is the wavelength of the radio waves, expressed in meters?

1.5.2 The Quantum Theory and Photon

Dear learner, in this subsection you will learn about the Max Planck theory: a new view of matter and energy and the photoelectric effect described by Albert Einstein.

Why are waves treated as particles and particles as the waves too?

A German physicist, Max Planck (1900), came to a new view of matter and energy. He made a revolutionary proposal; energy is discontinuous like matter.

According to Planck, atoms and molecules could emit or absorb energy only in discrete quantities, like small packages or bundles. Each of these small "packets" of energy is called a *quantum*. The energy of a quantum is proportional to the frequency of the radiation. The energy E of a single quantum is given by:

$$E = hv$$

(1.2)

Where h is called Planck's constant and v is the frequency of radiation. The value of Planck's constant is 6.63×10^{-34} J. s.

) What is the relationship between energy of a quantum and wavelength of light?

From the relation, $v = \frac{hc}{\lambda}$, Equation 1.2 can also be expressed as: $E = \frac{hc}{\lambda}$ (1.3)

Dear learner, according to quantum theory, energy is always emitted or absorbed in integral multiples of hv; for example, hv, 2hv, 3hv, etc. A system can transfer energy only in whole quanta. Thus, energy seems to have particulate properties.

Quantum is a small "packet" of energy

Example 1.4

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The blue color in fireworks is often achieved by heating copper (I) chloride (CuCl) to about 1200 °C. Then the compound emits blue light having a wavelength of 600 nm. What is the increment of energy (the quantum) emitted at 600 nm by CuCl?

Solution:

The quantum of energy can be calculated from the Equation 1.2:

E = hv

The frequency (v) for this case can be calculated as follows:

$$v = \frac{C}{\lambda}$$

= $\frac{3.0 \times 10^{9} \text{ m s}^{-1}}{6.0 \times 10^{-7} \text{ m}}$
= $0.5 \times 10^{15} \text{ s}^{-1}$
So, $E = hv$
= $6.63 \times 10^{-34} \text{ J s} \times 0.50 \times 10^{15} \text{ s}^{-1}$
= $3.315 \times 10^{-19} \text{ J} = 3.32 \times 10^{-19} \text{ J}$

The Photoelectric Effect

?) What can happen if the surface of a metal is exposed to light?

Dear learner, Albert Einstein (1905) used quantum theory to explain the photoelectric effect. The photoelectric effect is a phenomenon in which electrons are ejected from the surface of certain metals exposed to light of at least a certain minimum frequency called the threshold frequency, n_o.

These observations can be explained by assuming that EMR is quantized (consists of photons), and the threshold frequency represents the minimum energy required to remove the electron from the metal's surface.

Photons are particles of light or energy packets. The minimum energy required to remove an electron is:

$$E_{o} = hv_{o} \tag{1.4}$$

Where E_{o} is the minimum energy of the photon, and v_{o} , the threshold frequency.

Dear learner, a photon with energy less than Eo ($v < v_o$) cannot remove an electron, or a light with a frequency less than the v_o produces no electrons. On the contrary, if a light has $> v_o$, the energy in excess of that required to remove the electron is given to the electron as kinetic energy (KE):

$$KE_{e} = \frac{1}{2} mv^{2} = hv - hv_{o}$$
 (1.5)

Where KE_e is the kinetic energy of an electron, **m** is the mass of an electron, **v** is the velocity of the electron, **hv**, is the energy of an incident photon, and hv_o is the energy required to remove an electron from the metal's surface.

What is the relation between the intensity of light and the number of photons?

The intensity of light is a measure of the number of photons present in a given part of the beam: a greater intensity means that more photons are available to release electrons (as long as v > vo for the radiation).

What does the dual nature of light mean?

In his theory of relativity in 1905, Einstein derived the famous Equation:

$$E = mc^2 \tag{1.6}$$

Rearranging this equation, we have,

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$$m = \frac{E}{C^2}$$
(1.7)

Where E is energy, m is mass, and c is speed of light.

The main significance of Equation 1.7 is that energy has mass. Using this equation, you can calculate the mass associated with a given quantity of energy or the apparent mass of a photon.

For electromagnetic radiation of wavelength, λ , the energy of each photon is given by the expression:

$$E_{\text{photon}} = \frac{hc}{\lambda} \tag{1.8}$$

Then the apparent mass of a photon of light with wavelength is given by:

$$m = \frac{E}{C^2} = \frac{hC}{\lambda C^2} = \frac{h}{\lambda C}$$
(1.9)

You can summarize the important conclusions from the work of Planck and Einstein as follows:

Energy is quantized. It can occur only in discrete units called photon or quanta.

Dear learner, EMR, which was previously thought to exhibit only wave properties, seems to show certain characteristics of particulate matter as well. This phenomenon is sometimes referred to as the dual nature of light and is illustrated **Figure 1.7**.



Figure 1.7: The dual nature of light

Example 1.5

Compare the wavelength for an electron (mass = 9.11×10^{31} kg) traveling at a speed of 1.00×10^7 m/s with that for a ball (mass = 0.10 kg) traveling at 35 m/s.

Solution:

We use the equation $m = \frac{h}{\lambda C}$ Where $h = 6.63 \times 10^{-34} \text{ J s} = 6.63 \times 10^{-34} \text{ kg m}^2 \text{ s}$ Since 1 J = 1 kg m²/s² For the electron,

$$\lambda = \frac{6.63 \times 10^{-34} \text{ kg m}^2 \text{ s}}{9.11 \times 10^{-31} \text{ m} \times 1.00 \times 10^7 \text{ m/s}} = 7.28 \times 10^{-11} \text{ m} = 728 \text{ nm}$$

For the ball

$$\lambda = \frac{6.63 \times 10^{-34} \text{ kg m}^2 \text{ s}}{0.10 \text{ m} \times 35 \text{ m/s}} = 1.89 \times 10^{-34} \text{ m} = 189 \times 10^{-25} \text{ nm}$$

The maximum kinetic energy of the photoelectrons emitted from a give metal is 1.5×10^{-20} J when a light that has a 750 nm wavelength shines on the surface. Determine the threshold frequency, v_o , for this metal. Calculate the corresponding wavelength, λ_o .

Solution:

a) Determination of the threshold frequency, $v_{_{\rm o}}$

Solve for v from c = v × λ

Thus,
$$v = \frac{C}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s}}{4.5 \times 10^{-7} \text{ m}}$$

 $= 0.4 \times 10^{15} \text{ s}^{-1} = 4.0 \times 10^{14} \text{ s}^{-1}$

Similarly, rearrange Equation 1.5 and solve v_{o}

$$\begin{aligned} v_{o} &= \frac{hv - KE}{h} \\ &= \frac{\left(6.63 \times 10^{-34} \text{ J s} \times 4.0 \times 10^{14} \text{s}^{-1}\right) - 1.5 \times 10^{-20} \text{ J}}{6.63 \times 10^{-34} \text{ J s}} \\ &= \frac{26.52 \times 10^{-20} \text{ J} - 1.5 \times 10^{-20} \text{ J}}{6.63 \times 10^{-34} \text{ J s}} = \frac{24.02 \times 10^{-20} \text{ J}}{6.63 \times 10^{-34} \text{ J s}} = 3.77 \times 10^{14} \text{ s}^{-1} \\ \text{Therefore, a frequency of } 3.77 \times 10^{14} \text{ Hz is the minimum (threshold)} \\ \text{required to cause the photoelectric effect for this metal.} \end{aligned}$$

b) Calculate the corresponding wavelength $\lambda_{_{\rm o}}$

$$\lambda_{\circ} = \frac{C}{v_{\circ}} = \frac{3.0 \times 8 \text{ m/s}}{3.77 \times 10^{14} \text{s}^{-1}} = 0.796 \times 10^{-6} \text{m} = 796 \text{ nm}$$

🛠) Self-test Exercise 1.4

- 1. The following are representative wavelengths in the infrared, ultraviolet, and X-ray regions of the electromagnetic spectrum, respectively: 1.0×10^{-6} m, 1.0×10^{-8} m, and 1.0×10^{-10} m.
 - a. What is the energy of a photon of each radiation?
 - b. Which has the greatest amount of energy per photon?
 - c. Which has the least?
- A clean metal surface is irradiated with light of three different wavelengths l₁, l₂, and l₃. The kinetic energies of the ejected electrons are as follows: l₁: 7.2 x 10⁻²⁰ J; l₂: approximately zero; l₃: 5.8 x 10⁻¹⁹ J. Which light has the shortest wavelength and which has the longest wavelength? Determine the threshold frequency, v_a, for this metal.

1.5.3 Atomic Spectra

Dear learner, in this subsection you will learn how atom spectra are formed. You will also learn how atoms form discrete (line) spectrum. You will learn that spectral lines are characteristic of the atom (element) producing them and used for their identification.

) What happens when a narrow beam of light passes through a prism?

Dear learner, different kinds of spectrum are observed when an electric discharge, or spark, passes through a gas such as hydrogen. The electric discharge is an electric current that excites, or energizes, the atoms of the gas. More specifically, the electric current transfers energy to the electrons in the atoms raising them to excited states. The atoms then emit the absorbed energy in the form of light as the electrons return to lower energy states.

When a narrow beam of this light is passed through a prism, you do not see a continuous spectrum, or rainbow, as sunlight does. Rather, only you observe a few colors, displayed as a series of individual lines. This series of lines is called the element's atomic spectrum or emission spectrum. The wavelengths of these spectral lines are characteristic of the element producing them, and used for their identification. For example, the emission (line) spectrum of hydrogen atom is show in **Figure 1.8**.

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Figure 1.8: The hydrogen line spectrum, containing only a few discrete wavelengths

Dear learner, changes in energy between discrete energy levels in hydrogen will produce only certain wavelengths of the emitted light, as shown in **Figure 1.9.** For example, a given change in energy from a high to a lower level would give a wavelength of light that can be expressed using Planck's Equation:

$$\Delta E = hv = hc/\lambda$$







Atomic or line spectra are produced from the emission of photons of electromagnetic radiation (light).

When compounds of the alkali metals: lithium, sodium, and potassium are burned they give different colored flames, as shown in **Figure 1.9**.

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Section 1.5: Electromagnetic Radiation and Atomic Spectra



Atomic or line spectra are produced from the emission of photons of electromagnetic radiation (light).

Dear learner, different elements have different flame colors because their electrons have different allowed energy levels.

Note that every element has a unique emission spectrum which serves as an 'atomic finger print' and can be used to identify the element.

1.5.4 The Bohr Model of the Hydrogen Atom

Dear learner, in this subsection you will learn the Bohr model of hydrogen atom. You will learn how the absorption and emission of light by hydrogen atoms are related to the change in energy of the electrons within the atoms. You will also study the assumption that Bohr proposed based on Planck's and Einstein's ideas about quantized energy and position of electron in hydrogen atom.



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Explain why an electron does not enter the nucleus even though they are oppositely charged?

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A Danish physicist, Niels Bohr (1913), explained why the orbiting electron does not radiate energy when it moves around the nucleus. He introduced the fundamental idea that the absorption and emission of light by hydrogen atoms to energy changes of the electrons within the atoms. The fact that only certain frequencies are absorbed or emitted by an atom tells you that only certain energy changes are possible. Thus, energy changes in an atom are quantized.

Bohr used Planck's and Einstein's ideas about quantized energy and proposed the following assumptions:

- 1. The electron in a hydrogen atom travels around the nucleus in a circular orbit.
- 2. The energy of the electron in an atom is proportional to its distance from the nucleus. The further an electron is from the nucleus; the more energy it has.
- 3. Only limited number of orbits with certain energies are allowed. This means, the orbits are quantized.
- 4. The only orbits that are allowed are those for which the angular momentum of the electron is an integral multiple of $h/2\pi$.
- 5. As long as an electron stays in a given orbital, it either gains or loses energy. That means, the atom does not change its energy while the electron moves within an orbit.
- 6. The electron moves to a higher energy orbit only by absorbing energy and emitting light when it falls to a lower energy orbit. The energy (photon) of the light absorbed or emitted is exactly equal to the difference between the energies of the two orbits.

Furthermore, Bohr showed that the radii, **r**, of the permitted orbits or energy levels for a hydrogen atom are related to Planck's constant, **h**, the electron's charge, **e**, and its mass, **m**.



What is the total energy of an electron in Bohr atomic model?

Dear learner, consider hydrogen atom with an electron with constant speed, v, circulating the nucleus in an orbit of radius, r. The total energy of the electron is the sum of the kinetic energy and the potential energy.

$$E = \frac{1}{2}mv^2 + \frac{e^2}{r}$$
(1.10)

For an electron to exist in a stable orbit of a constant radius, the centripetal force (attracting the electron to the nucleus), e^2/r^2 , and the centrifugal force (pulling away the electron from the nucleus), mv^2/r must be equal.

$$\frac{e^2}{r^2} = \frac{mv^2}{r}$$
(1.11)



What is quantum number in Bohr model?

Bohr then introduced an additional requirement that the angular momentum, mvr, of the electron can take only certain permitted values that is an integral multiple of $h/2\pi$. This requirement is called a quantum condition:

$$mvr = \frac{nh}{2\pi}$$
(1.12)

Solving equation (1.12) for v gives

$$v = \frac{nh}{2\pi mr}$$
(1.13)

Substituting for v in equation 1.11 and solving for r gives

$$r = \frac{n^2 h^2}{4\pi m e^2} = n^2 \left(\frac{h^2}{4\pi m e^2}\right)$$
(1.14)

Here, n is positive integer (n = 1, 2, 3...) and is called quantum number. It is known that h, π , m, and e are constants, thus Equation 1.14 ca be simplified to

$$r = n^2 \alpha_{\rho} \tag{1.15}$$

According to Equation 1.15 the only permitted values of the radii of the electron path in the hydrogen atom are those proportional to the square of a whole number, **n**. The numerical values of a_0 is 0.53 Å (angstroms). Thus,

$$r = (0.53 \text{ Å})n^2$$

For instance, for n = 1, r = 0.53 Å; which is the first Bohr radius. The larger the values of n, the further the electron from the nucleus. Figure 1. 11 shows Bohr's electron orbits of hydrogen atom



Figure 1.11: Energy levels of a hydrogen atom

Bohr showed that the energies that an electron in hydrogen atom can occupy are given by:

$$E_n = -\frac{R_H}{n^2}$$
(1.16)

Where $R_{\rm H}$ is the Rydberg constant for the hydrogen atom and has the value 2.18 x 10⁻¹⁸ J. Thus, Equation 1.16 can be written as:

$$E_n = -\frac{2.18 \times 10^{-18} J}{n^2}$$

Dear learner! The negative sign in **Equation 1.16** is an arbitrary convention, signifying that the energy of the electron in the atom is lower than the energy of a free electron, which is an electron that is infinitely far from the nucleus.

As the electron gets closer to the nucleus (as **n** decreases), E_n becomes larger in absolute value, but also more negative. The most negative value, then, is reached when n = 1, which corresponds to the most stable energy state. We call this the ground state or ground level, which refers to the lowest energy

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state of an atom. The stability of the hydrogen electron diminishes for n = 2, 3.... Each of these levels is called an excited state, or excited level, which is higher in energy than the ground state.

Ground state/level: refers to the lowest energy state of an atom.

Excited state/level: refers highest energy than the ground state.

A hydrogen electron for which n > 1 is said to be in an excited state.



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Using Bohr's model of an atom, how can the energy (E_n) and energy level (n) are related?

Using Bohr's Equation 1.16, relating the energy (E_n) and energy level (n) for an electron it is possible to calculate the energy of a single electron in a ground state or excited state, or the energy change when an electron moves between two energy levels.

Example 1.6

 Consider the n = 5 state of hydrogen atom. Using the Bohr model, calculate the radius of the electron orbit, the velocity and the energy of the electron.

Solution:

To determine the radius, Equation 1.15 is used:

$$r = (0.53 \text{ Å}) \times 5^2 = 15.24 \text{ Å} = 1.524 \text{ nm}$$

Velocity of the electron is determined using Equation 1.13

$$v = \frac{nh}{2\pi mr}$$

$$= \frac{5(6.63 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1})}{2 \times 3.14 \times 9.11 \times 10^{-31} \text{ kg} \times 1.524 \times 10^{-9} \text{ m}}$$

$$= 3.8 \times 10^5 \text{m/s}$$
Equation 1.15 is used to solve E₅
E₅ = $-\frac{2.18 \times 10^{-18} \text{ J}}{5^2} = 8.72 \times 10^{-20} \text{ J}$
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Calculate the energies of the hydrogen electron in n = 1, n = 2, and the n = 3Solution: Using Equation 1.16 n = 1, $E_1 = -\frac{2.18 \times 10^{-18} \text{ J}}{1^2} = -2.18 \times 10^{-18} \text{ J}$ For n = 2, $E_2 = -\frac{2.18 \times 10^{-18} \text{ J}}{2^2} = -5.54 \times 10^{-19} \text{ J}$ n = 3, $E_3 = -\frac{2.18 \times 10^{-18} \text{ J}}{2^2} = -2.42 \times 10^{-19} \text{ J}$

Dear learner, Bohr's model also quantitatively explained the line spectra of hydrogen atom. He proposed that the absorptions and emissions in line spectra correspond to the transfer of the electron from one orbit to another.

Energy must be absorbed for the electron to move from one orbit to another one having a bigger radius. Whereas, energy is emitted when an electron moves from the higher orbital energy level, n_i , to the lower orbital, $n_{r'}$ energy level. Thus, the change in energy, ΔE , is the difference in the energy between the final and the initial state electron:

$$\Delta E = E_f - E_i \tag{1.17}$$

This equation is similar to:

$$\Delta E = -R_{H} \left(\frac{1}{n_{f}^{2}} - \frac{1}{n_{i}^{2}} \right)$$

$$= -2.18 \times 10^{-18} J \left(\frac{1}{n_{f}^{2}} - \frac{1}{n_{i}^{2}} \right)$$
(1.18)

Where \mathbf{n}_i and \mathbf{n}_i represent quantum numbers for initial and final states respectively. But, $\Delta E = hv$, thus we have

$$\Delta E = h\nu = -2.18 \times 10^{-18} J\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$
(1.19)

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Note that when $n_f > n_i$, ΔE is positive, indicating that the system has absorbed energy. However, when $n_i > n_f$, ΔE is negative and this corresponds to emission of energy.

Example 1.7

 Calculate the energy emitted when an electron moves from the n = 3 to the n = 2 energy level. Determine the wavelength of the emitted energy.
 Solution: Equation 2.19 is used, to determine the emitted energy

$$\Delta E = -2.18 \times 10^{-18} \text{ J} \left(\frac{1}{2^2} - \frac{1}{3^2}\right) = 3.03 \times 10^{-19} \text{ J}$$

To obtain λ , Equation 1.19

$$\Delta E = hv = \frac{hc}{\lambda}$$
Thus, $\lambda = \frac{hc}{\Delta E} = \frac{6.63 \times 10^{-34} \text{ Js} \times 3.0 \times 10^8 \frac{\text{m}}{\text{s}}}{3.03 \times 10^{-19} \text{ J}}$

$$= 6.56 \times 10^{-7} m = 656 \text{ nm}$$

Self-test Exercise 1.5

- 1. What is the wavelength of a photon (in nanometers) emitted during a transition from the n = 5 state to the n = 2 state in the hydrogen atom?
- 2. Calculate the frequency of the green line arising from the electron moving from n = 4 to n = 20 in the visible spectrum of the hydrogen atom using Bohr's theory.



When we study a large number of hydrogen atoms, we observe all possible transitions and hence the corresponding spectral lines. For instance, **Figure 1.12** illustrates line spectra of hydrogen atom, when its electron moves from n = 4 to n = 1, n = 3 to n = 1 and n = 2 to n = 1.

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Each horizontal line represents an allowed energy level for the electron in a hydrogen atom. The energy levels are labeled with their principal quantum numbers.

The emission spectrum of hydrogen includes a wide range of wavelengths from the infrared to the ultraviolet. Table 1.2 shows the series of transitions in the hydrogen spectrum; they are named after their discoverers. The Balmer series was particularly easy to study because some of its are lines fall in the visible range.

Series	n_f	<i>n</i> _i	Spectrum Region
Lyman	1	2, 3, 4,	Ultraviolet
Balmer	2	3, 4, 5,	Visible and ultraviolet
Paschen	3	4, 5, 6,	Infrared
Brackett	4	5, 6, 7,	Infrared

Tallala	1 0 .	TI				las salura as a sa		and a shirt was
IDDIE	1 2 .	INP	Various	series in	atomic	nvaroaen	emission	SDECTIUM
IGDIC	· • 🗠 •		V GHOOS	201102111	aronnic	nyarogon	01111351011	speciforn

⁷ For a larger orbit radius (i.e., a higher atomic energy level), the electron drops further, the energy is greater (higher v, shorter λ) of the emitted photon.

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Self-test Exercise 1.6

- 1. Calculate the energies of the states of the hydrogen atom with n = 2and n = 3, and calculate the wavelength of the photon emitted by the atom when an electron makes a transition between these two states.
- 2. What is the wavelength of a photon emitted during a transition

from the n_i = 10 state to the n_f = 2 state in the hydrogen atom?

1.5.5 Limitations of the Bohr Model

Dear learner, the Bohr Model was an important step in the development of atomic theory. He introduced the idea of quantized energy states for an electron in a hydrogen atom.



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Can you state the drawbacks of Bohr's atomic model?

The model was adequate for explaining atoms and ions containing only one electron such as H, He⁺ and Li²⁺.

However, the model has the following drawbacks:

- It doesn't explain the atomic spectra of more complicated atoms and ions, even that of helium, the next simplest element.
- It doesn't explain about further splitting of spectral lines in the hydrogen spectra on application of a magnetic field.
- It considers electrons to have both known radius and orbit, which is impossible according to Heisenberg's uncertainty principle.

1.5.6 The wave-particle duality of matter and energy

Dear learner, in this subsection you will study about Louis de Broglie principle which explain the dual character of energy and matter, i.e., matter and energy show both wave and particulate properties. This concept is important to understand how any particle of mass, *m*, whether a planet, ball, or electron-moving at speed has a certain wavelength; a wave like properties.

Bohr's model assumed that an atom has only certain allowable energy levels in order to explain the observed line spectrum. However, his assumption had no basis in physical theory.

A French physicist, Louis de Broglie (1924), by combining Einstein's equation for the quantity of energy equivalent to a given amount of mass ($E = mc^2$) with Planck's equation for the energy of a photon ($E = hv = hc/\lambda$), he derived an equation for the wavelength of any particle of mass, m, whether a planet, ball, or electron-moving at speed, v.

$$\lambda = \frac{h}{mv} \tag{1.20}$$

According to this **Equation 1.20**, matter and energy show both wave and particulate properties. The distinction between a particle and a wave is only meaningful in the macroscopic world, it disappears at the atomic level.

This dual character of matter and energy is known as the waveparticle duality.

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Example 1.8

1. Calculate the de Broglie wavelength of an electron that has a velocity of 1.00×10^6 m/s. (electron mass, $m_e = 9.11 \times 10^{-31}$ kg; $h = 6.63 \times 10^{-34}$ J. s)

Solution:

$$\lambda = \frac{h}{m_2 v} = \frac{6.63 \times 10^{-34} \text{ kg} \frac{\text{m}^2}{\text{s}}}{1.00 \times 10^6 \frac{\text{m}}{\text{s}} \times 9.11 \times 10^{-31} \text{ kg}}$$

 $= 0.73 \times 10^{-9} m = 0.728 nm$

2. What is the speed of an electron that has a de Broglie wavelength of 1.00 nm?

Solution:

$$v = \frac{h}{m_2 \lambda} = \frac{6.63 \times 10^{-34} \text{ kg} \frac{\text{m}^2}{\text{s}}}{1.00 \times 10^{-9} \text{ m} \times 9.11 \times 10^{-31} \text{ kg}}$$

$$= 0.728 \times 10^6 \frac{\text{m}}{\text{s}} = 7.28 \times 10^5 \frac{\text{m}}{\text{s}}$$

Self-test Exercise 1.7

- 1. What is the characteristic wavelength of an electron (in nm) that has a velocity of 5.97 ×10⁶ m/s (m_e = 9.11 x 10⁻³¹ kg)?
- 2. Calculate the wavelength (in nanometers) of an H atom (mass = 1.67 x 10^{-27} kg) moving at 7.00 x 10^2 cm/s.
- 3. Calculate the wavelength (in nm) of a photon emitted when a hydrogen atom undergoes a transition from n = 5 to n = 2.

\checkmark CHECKLIST - 1.5

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

l can ..

- characterize EMR in terms of wavelength, frequency and speed
- calculate the wavelength and frequency of EMR
- explain the wave and particle nature of light
- describe emission spectra of atoms as consisting of a series of lineS
- define a photon as a unit of light energy
- distinguish how the photon theory explains the photoelectric effect and the relationship between a photon absorbed and an electron released
- state Bohr's assumption of energy of the electron in an hydrogen atom
- calculate the radius of electron orbit, the electron velocity and the energy of an electron using Bohr's model
- identify the line spectrum of hydrogen demonstrates the quantized nature of the energy of its electron
- explain that atoms emit or absorb energy when they undergo transitions from one state to another
- compose the limitations of Bohr's theory.

Section 1.6: The Quantum Mechanical Model of the Atom

Dear learner, in this section you will be introduced about quantum mechanics, which you will learn in detail when join higher education. Now, you will learn how the quantum mechanical model helps you to understand the distribution of electrons around the nucleus of an atom.

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

- state Heinsberg's uncertainty principle
- describe the significance of electron probability distribution.
- explain the quantum numbers n, l, m_l , m_s
- write all possible sets of quantum numbers of electrons in an atom.
- describe the shapes of orbitals designated by s, p and d.

) If electrons possess particle nature, it should be possible to locate electrons. How can an electron be located?

Dear learner, the Bohr model of an electron orbiting around the nucleus, looks like a planet around the Sun which does not explain many properties of atoms. The planetary view of one charged particle orbiting another particle of opposite charge does not match some of the best-known laws of classical physics.

Scientists have developed Quantum mechanics, which presents a different view of how electrons are arranged about the nucleus in the atom. This view depends on two central concepts: the wave behavior of matter and the uncertainty principle.

⁷ The combination of these two ideas lead to a mathematical description of electronic structure.

1.6.1 The Heisenberg's Principle

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Dear learner, the discovery of the wave properties of matter raised a new and very interesting question. If a subatomic particle can exhibit the properties of a wave, it is possible to say precisely where that particle is located. However, a wave, as a whole, extends in space, so its location is not defined precisely. A German physicist, Werner Heisenberg (1827), proposed the Heisenberg uncertainty principle, stating that it is not possible to know with great certainty both an electron's position and its momentum p (where p = mv) at the same time.

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Heisenberg was able to show that if Δp is the uncertainty in the momentum and Δx is the uncertainty of the position of the electron, then:

$$\Delta \times \Delta p \geq \frac{h}{4\pi}$$
 or $\Delta \times \Delta mv \geq \frac{h}{4\pi}$

This equation shows that if we measure the momentum of a particle precisely (i.e., if Δp is very small), then the position will be correspondingly less precise (i.e, Δx will become larger).

In simpler terms, if we know precisely where a particle is, we cannot know precisely where it has come from or where it is going.

If we know precisely how a particle is moving, we don't know exactly where it is.

The Heisenberg uncertainty principle has no practical consequence in everyday objects in the macroscopic world. However, it has very significant consequences when applied to atomic and subatomic particles.

1.6.2 Quantum Numbers

Dear learner! Under this topic you will learn the quantum mechanical description of the atom based on the assumption that there are waves associated with both matter and radiations. The mathematical description given to wave provides information about the energy of electrons and their position.

Erwin Schrödinger (1927), suggested that an electron or any other particle exhibiting wavelike properties can be described by a mathematical equation called a wave function (devoted Greek latter psi, ψ). This equation is very complex and it has series of solutions that describe the allowed energy states of electrons.

⁷ The square of a wave function, ψ^2 , gives the probability of finding an electron in a certain region of a space.

The results of a quantum mechanics indicate that the electron may be visualized as being in a rapid motion within a given region of space around the nucleus. Though we cannot determine the precise position of an electron, the probability of the electron being at a definite location can be calculated.



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The region in which an electron is most likely found is called an orbital.

The electron may occupy anywhere within an orbital at any instant in time. So, an electron can be considered as a particle that can rapidly move from place to place, behaving like an "electron cloud" whose density varies within the orbital.

Dear learner! According to quantum mechanics model, each electron in an atom is descried by four quantum numbers.

Three of these quantum numbers specify the wave function that gives the probability of finding an electron at various places in space. The fourth one is used to describe the spin (magnetic property) of the electrons that occupy the orbitals.

The four quantum numbers are:

1. The principal quantum number (n): describes the main energy level, or shell, an electron occupies. It may be any positive integer, n = 1, 2, 3, 4, etc.

Principal quantum number describes the size and energy of the shell in which the orbital resides. It is also similar to the energy levels in Bohr's model.

 The angular momentum quantum number (?): designate the shape of the atomic orbitals. A shell may contain different sublevels or subshells, each with a characteristic shape. It takes values from 0 to n - 1. It is also called azimuthal or the subsidiary quantum number.

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Orbitals of the same n, but different ℓ are said to belong to different subshells. A given energy level, n, has n^2 total number of orbitals.

For example, in n = 4, ℓ has values of 0, 1, 2, and 3. Each value of ℓ corresponds to an orbital label and an orbital shape. The following letters are usually used to denote the values of ℓ :

Value of <i>t</i>	Corresponding subshell
0	S
1	p
2	d
3	f

Example 1.9

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What do "4" and "d" represent in 4d?
```

Solution:

The number 4 gives the principal quantum number (n = 4) and the letter d gives the type of orbital (l = 2).

Activity 1.7

What are the values of *n* and *l* for the following subshells?

 a. 1s
 b. 3p
 c. 4s
 d. 4f
 e. 3d

 Write the subshell notations that correspond to

 a. n = 3 and *l* = 0
 c. n = 7 and *l* = 0
 b. n = 3 and *l* = 1
 d. n = 3 and *l* = 2

3. The magnetic quantum number (m_t) is related to the orientation of the orbital in space relative to the other orbitals in the atom. It is also called the orbital-orientation quantum number.

It has integral values between \cdot^{t} and \cdot^{t} , including 0. The number of possible m_{t} values or orbitals for a given ℓ value is $\star^{t} \ell + 1$.

4. The electron spin quantum number (m_s) refers to the spin of an electron and the orientation of the magnetic field produced by this spin. For every set of n, ℓ , and m_ℓ values, m_s can take the value $+\frac{1}{2}$ or $-\frac{1}{2}$.

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Each atomic orbital cannot accommodate more than two electrons, one with $m_s = +\frac{1}{2}$ and another with $m_s = -\frac{1}{2}$.

Example 1.10

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1. What are the allowed values of the quantum numbers through n = 2?

Solution: The allowed quantum number values in n = 2 are;

n	ł	m,	m _s	Electron capacity of subshell 4 ? + 2	Electron capacity of shell 2n ²
1	0 (1s)	0	+1/2, -1/2	2	2
	0 (2s)	0	+1/2, -1/2	2	
2	1 (2p)	-1,0,1	$\pm \frac{1}{2}$ for each value of <i>m</i> ,	6	8

What values of the angular momentum quantum number (ℓ) and magnetic quantum number (m_{ℓ}) are allowed for a principal quantum number, n = 3? How many orbitals are allowed for n = 3?

Solution:

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 $\ell = 2, m_{\ell} = -2, -1, 0, +1, +2$

Number of orbitals in n = 3 is $n^2 = 32 = 9$ orbitals

3s: 1 orbital

3p: 3 orbitals

3d: 5 orbitals

Total = 9 orbitals

Self-test Exercise 1.8

- 1. What are the allowed values of the quantum numbers through n = 4?
- 2. Write an orbital designation corresponding to the quantum numbers

$n = 4, \ell = 2, m_{\ell} = 0$	c. n = 5, <i>l</i> = 1, m _l = -1	
$n = 3, \ell = 1, m_s = 1$	d. n = 4, <i>l</i> = 3, m, = -3	

- 3. What values l and ml are allowed for a principal quantum number (n) of 3? How many orbitals are allowed for n = 3?
- 4. Can an orbital have the quantum number n = 2, l = 2, ml = 2
- 5. For an orbital with n = 3 and $m\ell = 1$, what is (are) the possible value(s) of ℓ ?
- 6. Which of the following orbitals do not exist: 1p, 2s, 2d, 3p, 3d, and 3f?

1.6.3 Shapes of atomic orbitals

What do the shapes of the orbitals?

Dear learner! Under this topic you will learn about the three-dimensional orientation of atomic orbitals. Shapes of atomic orbitals are usually represented by drawing boundary surface diagrams that encloses the highest probability (about 90%) of the total electron density in an orbital.

Figure 1.13 shows that the region of the greatest probability of finding an s electron is in a spherically symmetrical space whose origin is the atomic nucleus.

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The 1s orbital is the smallest, the 2s orbital is larger than 1s, and so on. Regardlessoftheirprincipal quantum numbers, all sorbitals are spherical.



Figure 1.13: The three s-orbitals

The p orbitals of any principal quantum number are arranged along three mutually perpendicular axes, x, y, and z, so that the region of the highest electron density are in dumb bell-shaped boundary surface (Figure 1.14).



Figure 1.14: The three p-orbitals

What is the difference between a $2p_x$ and a $2p_y$ orbital?

Two lobes of each p orbital lie along a line with the nucleus at their center. For instance, the three 2p orbitals are classified as $2p_x$, $2p_y$ and $2p_z$.

Dear learner, higher principal energy levels (n > 3) have five d orbitals in addition to, one s orbital, three 3p orbitals. The special orientations of d orbitals are much more complex in shape than p orbitals. **Figure 1.15** shows the different distribution of the five d atomic orbitals.

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Figure 1.15: The five d orbitals.

The principal energy level whose values are $n \ge 4$ have additional fourth sublevels. For instance, in n = 4, there are a 4s subshell with one orbital, a 4p subshell with three orbitals, a 4d subshell with five orbitals, and a 4f subshell with seven orbitals. The f orbitals have more complex shapes than the d orbitals.

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

 \square

I can...

- state Heinsberg's uncertainty principle
- describe the significance of electron probability distribution
- explain the quantum numbers n, l, m_μ, m_s
- write all possible sets of quantum numbers of electrons in an atom.
- describe the shapes of orbitals designated by s, p and d.

Section 1.7: Electronic configurations and orbital diagrams

Dear learner, in this section you will learn basic principles governing the distribution of electrons among atomic orbitals. These principles will help you to easily distribute/configure electron in atomic orbitals. You will also learn how orbitals are represented using diagrams.s

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

- explain the Aufbau Principle
- explain pauli's Exclusion Principle
- explain Hund's Rule
- write ground state electron configurations of multi-electron atoms.



?) What does the Aufbau'S Principle state about electrons?

The electron configuration for any atom is governed by the following three principles.

 Aufbau (Building up) Principle: this is a scheme used to reproduce the electron configurations of the ground state of atoms by successively filling with electrons in a specific order (the building up order). According to Aufbau principle, electrons occupy the lowest-energy orbital available before entering the higher energy orbital.

Thus, the ground state electron configurations of atoms are obtained by filling the subshell in the successive order: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, and so on, as indicated in **Figure 1.16.**





- Hund's Rule: states that equal energy orbitals (degenerate orbitals) are each occupied by a single electron before the second electrons of opposite spin enters the orbital. In other words, each of the three 2p orbitals (2p_x, 2p_y and 2p_z) will hold a single electron before any of them receives a second electron.
- 3. Pauli's Exclusion Principle: No two electrons can have the same four quantum numbers. That means, they must differ in at least in one of the four quantum numbers.

1.7.1 Ground State Electronic Configuration of the Elements

What is electron configuration?

The electronic configuration of an atom describes the distribution of the electrons among atomic orbitals in the atom.

Dear learner, you can use the following two general methods to denote electron configurations. The first method involves the use of subshell (sublevel) notation that uses numbers to designate the principal energy levels and the letters s, p, d and f *notations* to identify the sublevels.

A superscript number following a letter indicates the number of electrons in the designated subshell (e). The designation can be written as nl^e.

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For instance, the ground state electron configuration of H (Z =1) is $1s^1$ and Li (Z = 3) is $1s^22s^1$

Do you think that the notation nle includes the electron spin quantum number?

The second mechanism is to present electron configurations through an orbital diagram.

In orbital diagram representations each orbital is represented by a rectangular box, circle, or just a line, with an arrow indicating the electron's presence and its direction of spin.

Traditionally, \uparrow represents, $m_s = +\frac{1}{2}$ whereas \downarrow designates $m_s = -\frac{1}{2}$.

The orbital diagrams for the first ten elements are:



Activity 1.8

Dear learner, attempt to answer the following question.

- 1. What does each box in an orbital diagram represent?
- 2. What are the similarities and differences between a 1s and a 2s orbital?
- 3. What is the difference between a $2p_x$ and a $2p_y$ orbital?
- 4. What is the meaning of the symbol 4d⁶?

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The three orbitals in the 2p sublevel are degenerate (have equal energy) and have the same *n* and *l* values), thus, the fifth electron of boron can go into any one of the 2p orbitals.

A p sublevel has l = 1, the m_l (orientation) values can be -1, 0, or +1. For convenience, the boxes are labeled from left to right, -1, 0, +1, and assume it enters the $m_l = -1$ orbital: n = 2, l = 1, $m_l = -1$, and $m_s = +\frac{1}{2}$.

The three 2p orbitals have equal energy, but they can be along the x, y, or z axes.

Self-test Exercise 1.9

1. Write the expected electron configuration for these atoms:

a. Na (Z = 11)	d. V (Z = 23)
b. Al (Z = 13)	e. Mn (Z =25)
c. P (Z =15)	f. Fe (Z = 26)

- 2. Indicate the total number of:
 - a. p electrons in N (Z = 7) c. 3d electrons in S (Z = 16)

b. s electrons in Si (Z = 14)

3. Indicate which of the following sets of quantum numbers in an atom are unacceptable and explain why:

a. (1, 0, ½, ½)	c. (3, 2, 1, 1)	e. (4, 3, -2, +½)
b. (3, 0, 0, +½)	d. (2, 2, 1, +½)	



For Li, the first 2 electrons occupy the 1s orbital, and the third electron must occupy the first orbital with n = 2, the 2s orbital. Thus, the electron configuration for Li is $1s^22s^1$.



How do you write the condensed electron configuration of an atom?

Dear learner, to avoid writing the inner-level electrons, this configuration is often abbreviated as [He] $2s^1$, where [He] represents the electron configuration of He, $1s^2$. Similarly, Mg (Z =12), has the configuration $1s^22s^22p^63s^2$, or [Ne] $3s^2$. Then the next six elements, aluminum to argon, have configurations obtained by filling the 3p orbitals one electron at a time.

The ground state electron configurations of some atoms are presented in Table 1.3

Atom	Atomic number (Z)	Electron configuration
Н	1	1s ¹
Не	2	1s ²
Be	4	1s ² 2s ² or [He]2s ²
0	8	1s ² 2s ² 2p ⁴ or [He]2s ² 2p ⁴
Ne	10	1s ² 2s ² 2p ⁶ or [He]2s ² 2p ⁶
CI	17	1s ² 2s ² 2p ⁶ 3s ² 3p ⁵ or [Ne] 3s ² 3p ⁵
Ar	18	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ or [Ne] 3s ² 3p ⁶
Са	20	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ² or [Ar] 4s ²
Sc	21	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ² 3d ¹ or [Ar] 4s ² 3d ¹
Cr	24	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ¹ 3d ⁵ or [Ar] 4s ¹ 3d ⁵

Table 1.3: Ground state electron configurations of some atoms



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What is valance electrons?

In the electron configuration of an atom, the electrons in the outermost principal quantum level of an atom are called valence electrons.

For example, the valence electrons of the Be atom, are the 2s electrons. For the Ne atom, the valence electrons are the electrons in the 2s and 2p electrons.

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Valence electrons are the most important electrons to chemists since they are involving in bonding, as you will learn in the next chapter.

The inner electrons are known as core electrons.



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Why do some elements have exceptional configuration?

Dear learner! Atoms of some elements, example, chromium and copper have exceptional electron configurations. The expected configurations, according to the Aufbau's principle, are not observed through the emission spectra and magnetic properties of these elements.

	Expected	Observed
Cr (Z = 24)	[Ar] 4s ² 3d ⁴	[Ar] 4s¹3d⁵
C9 (Z = 29)	[Ar] 4s ² 3d ⁹	[Ar] 4s ¹ 3d ¹⁰

The reasons for these exceptions to the Aufbau principle are not completely understood, but it seems that the half-filled 3d subshell of chromium (3d⁵) and the fully filled 3d subshell of copper (3d¹⁰) gives a special stability to the electron configurations.

Note that a half-filled 4s subshell and a half-filled 3d subshell gives a lower energy state for a Cr atom than having a filled 4s subshell.

Section 1.7: Electronic configurations and orbital diagrams

×	Activity	1.9		
Dear learner, evaluate yourself by completing the following table and write				
the first-row transition metals in your note book:				
	Atomic	Ground state	e electron	Condensed electron
Elemeni	number	configur	ation	configuration
Sc	21			
Ti	22			
V	23			
Cr	24			
Mn	25			
Fe	26			
Со	27			
Ni	28			
Cu	29			
Zn	30			

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

I can...

explain the Aufbau Principle
explain Pauli's Exclusion Principle
explain Hund's Rule
write ground state electron configurations of multi-electron atoms

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Section 1.8: Electronic Configurations and the Periodic Table of the Elements

Dear learner, in this section you will learn how elements are arranged in the periodic table. You will also find out how an understanding of atomic structure can help us to explain and correlate chemical and physical properties of the elements. You will study periodicity of physical properties; atomic and ionic size, ionization energy, and electron affinity, electronegativity. You will learn how this property vary across a period or down a group.

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

- understand how the electron configurations of elements show which group and period each element is in
- give a reasonable explanation for the shape of the periodic table
- classify elements as representative transition and inner-transition elements
- explain the general trends in atomic radius, ionization energy, electron affinity, electronegativity, and metallic character of elements within a period and group of the periodic table
- write the advantages of the periodic classification of elements.

1.8.1 The Modern Periodic Table

Dear learner, in your Grade 9 chemistry lessons, you learned about the development of Mendeleev and the modern periodic table.

What is the main difference between the modern and Mendeleev periodic tables?

Elements are arranged in the modern periodic table by using atomic number. When the elements rearranged according to increasing atomic number their physical and chemical properties repeat at regular intervals. A period is a horizontal row of the periodic table. All elements in a row have the same number of electron shells.

There are seven periods in the periodic table, with each one beginning at the far left. A new period begins when a new principal energy level begins filling with electrons.

A group is a vertical column of the periodic table, based on the organization of the outer shell electrons.

There are a total of 18 groups. Elements in a group form a family with some variations in physical properties. The chemical properties of an element depend on its electrons. Since elements in a group have the same number of electrons, their chemical properties are the same.

1.8.2 Classification of the Elements

Dear learner, in this subsection you will study basis of classifying elements and the importance of classifying elements in groups and periods. The periodic classification follows as a logical consequence of the electronic configuration of atoms. Elements are grouped according to their outer-shell electron configurations, which account for their similar chemical behavior. The period indicates the value of **n** for the outermost or valence shell.

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There are 18 groups and 7 periods in the modern periodic table.

What is the basis of classifying elements into representative, transition, and inner transition elements?

1.8.3 Periodic Properties

The electron configurations of the atoms display a periodic variation with increasing atomic number (nuclear charge). As a result, the elements show periodic variations of physical and chemical behavior.



How does the size of elements vary across a period?

Dear learner, now let us examine some periodic atomic properties like atomic radii, ionization energies, electron affinities, electronegativity, and metallic character.

Activity 1.10

Dear learner, can you explain the importance of classifying elements in groups and periods?

Atomic Size (Atomic Radii)



Why don't the sizes of atoms increase uniformly with increasing atomic number?

Dear learner, the atomic size can be estimated by measuring the distance between the nuclei of two adjacent atoms. This property is called the **atomic radius** (**Figure 1.17**).



Figure 1.17: Atomic radius for a diatomic molecule.

r = d/2

Atomic size varies within both a group and a period of the main-group elements. These variations in atomic size are the result of two opposing influences: changes in the principal quantum number (n) and changes in the effective nuclear charge (Z_{eff}). The net effect of these depends on the shielding of the increasing nuclear charge by the inner electrons.



 Z_{eff} is the nuclear charge an electron actually experiences.



How $Z_{_{eff}}$ and n affect atomic size down a group and across a period?

Dear learner, atomic radius generally increases in a group from top to bottom but it decreases in a period from left to right. These trends hold well for the main-group elements but not as for the transition elements (electrons enter an inner electron shell, not the valence shell).

Self-test Exercise 1.11

- What are the factors affecting the size of an atom of an element? Explain your answer using examples.
- 2. Using the periodic table, rank each set of main-group elements in order of decreasing atomic size:

a. Ca, Mg, Sr b. K, Ga, Ca c. Br, Rb, Kr d. Sr, Ca, Rb

 Why is the difference in atomic radius between the elements Z = 11 (Na; 186 pm) and Z = 12 (Mg; 160 pm) so large, whereas between Z = 24 (Cr; 125 pm) and Z = 25 (Mn; 124 pm) the difference is very small. (Note that "pm" is picometers.).

Ionization Energy (IE)

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The ionization energy (IE) is the energy (in kJ) required for the complete removal of 1 mol of electrons from 1 mol of gaseous atoms or ions.

Dear learner, multi-electron atoms can lose more than one electron, so the ionization energies required to remove each electron are numbered in sequence from the ground-state atom. For instance, boron atom has five electrons, two in an inner core ($1s^2$) and three valence electrons ($2s^22p^1$). The five ionization steps and their ionization energies, IE, through IE_s, are:

B (g)	\rightarrow B ⁺	(g) + e⁻	IΕ	= 801 kJ mol ⁻¹
B+ (g)	\rightarrow B ²⁺	(g) + e⁻	IE ₂	= 2427 kJ mol ⁻¹
B ²⁺ (g)	\rightarrow B ³⁺	(g) + e⁻	IE ₃	= 3660 kJ mol ⁻¹
B ³⁺ (g)	\rightarrow B ⁴⁺	(g) + e⁻	IE ₄	= 25,025 kJ mol ⁻¹
B ⁴⁺ (g)	\rightarrow B ⁵⁺	(g) + e⁻	IE	= 32,822 kJ mol ⁻¹

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The first ionization energy (IE_1) of an atom is the minimum energy needed to remove the highest-energy (the outermost) electron from the neutral atom in the gaseous state:

Atom (g) \rightarrow lon⁺(g) + e⁻ $\Delta E = IE_1 > 0$

The second ionization energy (IE_2) removes the second electron. Since the electron is being pulled away from a positively charged ion, IE_2 is always larger than IE_1 :

 $lon^+(g) \rightarrow lon^{2+}(g) + e^- \qquad \Delta E = lE_2 (always > lE_1)$

Atoms with a low IE_1 tend to form cations during reactions, whereas those with a high IE_1 (except the noble gases) often form anions.

Ionization energies display a periodic variation when plotted against atomic number, as **Figure 1.18** shows.

Within any period, ionization energies tend to increase with atomic number. Thus, the lowest values in a period are found for the Group IA elements (alkali metals) The largest ionization energies in any period occur for the noble-gas elements.



Figure 1.18: Ionization energy versus atomic number

Dear learner, small deviations from this general trend occur. A Group IIIA element (*ns*²*np*¹) has a smaller ionization energy than the preceding Group IIA element (*ns*²). Apparently, the np electron of the Group IIIA element is more easily removed than one of the ns electrons of the preceding Group IIA element. Also note that a Group VIA element (*ns*²*np*⁴) has a smaller ionization energy than the preceding Group VA element.



Explain why boron has a smaller first ionization energy than beryllium. The first ionization energy of nitrogen is higher than oxygen.

Self-test Exercise 1.12

1. Choose the element with the higher ionization energy from each pair.

a. As or Bi b. As or Br c. Al or In d. K or Ge

- The first and second ionization energies of potassium, K, are 419 kJ mol⁻¹ and 3052 kJ mol⁻¹ and those of calcium, Ca, are 590 kJ mol⁻¹ and 1145 kJ mol⁻¹, respectively, Compare their values and comment on the differences.
- 3. Based on their positions in the periodic table, predict which atom of the following pairs will have the larger first ionization energy:

a. Ga or Ge b. Br or Sb c. K or Cr d. Mg or Sr e. O or Ne

Electron Affinity (EA)

The electron affinity is the energy change for the process of adding an electron to a neutral atom in the gaseous state to form a negative ion:

Atom (g) + $e^- \rightarrow \text{Ion}^-(g) \Delta E = EA$

The first electron affinity (EA1) is usually negative, however, the second electron affinity (EA2) is always positive.

For example, the EA_1 and EA_2 of an oxygen atom can be written as:

 $O(g) + e^{-} \rightarrow O^{-}(g)$ $EA_{1} = -141 \text{ kJ mol}^{-1}$ $O^{-}(g) + 1e^{-} \rightarrow O^{2-}(g)$ $EA_{2} = +744 \text{ kJ mol}^{-1}$

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Figure 1.19 gives the electron affinities of the main-group elements. Negative values indicate that energy is released when the asnion forms. Positive values, which occur in group VIII A (18), indicate that energy is absorbed to form the anion; in fact, these anions are unstable and the values are estimated.

IA (1)	
Н -72.8	IIA (2)
Li	Be
-59.6	≤0
Na	Mg
-52.9	≤0
K	Ca
-48.4	-2.37
Rb	Sr
-46.9	-5.03
Cs	Ba
-45.5	-13.95

					VIIIA (18)
IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	He (0.0)
B	С	N	O	F	Ne
-26.7	-122	+7	-144	-328	(+29)
Al	Si	P	S	Cl	Ar
-42.5	-134	-72	-200	-349	(+35)
Ga	Ge	As	Se	Br	Kr (+39)
-28.9	-119	-78.2	-195	-325	
In	Sn	Sb	Te	I	Xe (+41)
-28.9	-107	-103	-190	-295	
Ti	Pb	Bi	Po	At	Rn (+41)
-19.3	-35.1	-91.3	-183	-270	

Figure 1.19: Electron affinities of the main-group elements (in kJ/mol)

Electron affinities, EA., also show a periodic variation, just as atomic radii and ionization energies do.

Dear learner, in broadly speaking, the general trend is toward more negative electron affinities from left to right in any period.



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What could be the reason for the IIA element and for the VA element tend to have smaller electron affinities than the preceding element?

Self-test Exercise 1.13

- Which group in the periodic table has elements with high (endothermic) EA₁ and very negative (exothermic) first electron affinities (EA₁)? Give the charge on the ions these atoms form.
- 2. Silicon has an electron affinity of -134 kJ/mol. The electron affinity of phosphorus is -72 kJ/mol. Give a reason for this difference.
- 3. Why are the electron affinities of the alkaline earth metals either negative or small positive values?

Electronegativity

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What is the electronegativity of an atom? How is it different from electron affinity?

Electronegativity is the extent of attraction by which the electrons of the bond pair are attracted by an atom linked by this bond.

Linus Pauling, an American scientist, in 1922 assigned arbitrarily an electronegativity value of 4.0 to fluorine, the element considered to have the greatest ability to attract electrons. This is known as the Pauling scale. Approximate values for the electronegativity of a few elements are given in Table 1.4.

Table 1. 4: (a) Electronegativity values (on Pauling scale) across the periods

Atom (Period II)	Li	Be	В	С	Ν	0	F
Electronegativity	1.0	1.5	2.0	2.5	3.0	3.5	4.0
Atom (Period III)	Na	Mg	Al	Si	Р	S	CI
Electronegativity	0.9	1.2	1.5	1.8	2.1	2.5	3.0

In a given family of the periodic table, what is the general relationship between electronegativity and size?

Table 1.4: (b) Electronegativity Values (on Pauling scale) down a group

Atom (Group I)	Electronegativity value	Atom (Group 17)	Electronegativity value
Li	1.0	F	4.0
Na	0.9	Cl	3.0
К	0.8	Br	2.8
Rb	0.8	l	2.5
Cs	0.7	At	2.2

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Electronegativity generally increases across a period from left to right and decreases down a group in the periodic table.

Metallic Character

- Can you predict the trend in metallic character as a function of position in the periodic table? Is it the same as the trend observed for atomic size and Ionization energy?
- Metallic properties arise from the elements, ability to lose electrons.

Dear learner, metallic character decreases as you move across a period in the periodic table from left to right. This occurs as atoms more readily accept electrons to fill a valence shell than lose them to remove the unfilled shell.

Metallic character increases as you move down an element group in the periodic table. This is because electrons become easier to lose as the atomic radius increases, where there is less attraction between the nucleus and the valence electrons because of the increased distance between them.

Self-test Exercise 1.14

- 1. In which region of the periodic table metals are located?
- In which region are the elements with general electronic configuration of ns²p⁵ located? Give the group number.
- 3. Write the group number and period of an element with atomic number 26.
- 4. Is an element with the electronic configuration 1s² 2s² 2p⁶ 3s¹ a metal or a non- metal? Will it form a cation or an anion readily? Explain your answer.

1.8.4 Advantages of Periodic Classification of the Elements

Dear learner, some of the advantages of periodic classification of elements are:

- The classification of elements is based on the atomic number, which is a fundamental property of an element.
- Isotopes are averaged in one place, as the classification is on the basis of atomic number(For example, chlorine (atomic number 17) is given the atomic mass 35.5.).
- It explains the periodicity of the properties of the elements and relates them to their electronic configurations.
- The position of the elements that were misfits on the basis of mass number (anomalous pairs like argon and potassium) could be justified on the basis of atomic number.
- The lanthanides and actinides are placed separately at the bottom of the periodic table.
- The table is a simple, systematic and an easy way of remembering the properties of various elements, as it is based on the electronic configuration.

CHECKLIST-1.8

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

l can...

- understand how the electron configurations of elements show which group and period each element is in
- give a reasonable explanation for the shape of the periodic table
- classify elements as representative transition and inner-transition elements
- explain the general trends in atomic radius, ionization energy, electron affinity, electronegativity, and metallic character of elements within a period and group of the periodic table
- write the advantages of the periodic classification of elements.

UNIT SUMMARY

- Modern chemistry began discoveries leading to the formulation of two basic laws of chemical combination: the law of conservation of mass and the law of constant composition (definite proportions).
- The first clues to the structure of atoms came through the discovery and characterization of cathode rays (electrons).
- Radioactive substances emit alpha particles, beta particles, and gamma rays.
- The nucleus is a tiny, dense, and positively charged central part of an atom
- Electromagnetic radiation is characterized by its wavelength (λ), frequency (v), and speed (c)
- Electromagnetic radiation can be viewed as a stream of "particles" called photons.
- The photoelectric effect is the emission of an electron from the surface of a metal, caused by electromagnetic radiation.
- Bohr's theory accounts for the observed atomic spectrum of hydrogen atom.
- The electron in a hydrogen atom can be viewed as a matter-wave enveloping the nucleus.
- A wave function is characterized by the value of four quantum numbers:
 n, l, m_l, and m_s.
- According to Heisenberg's Uncertainty Principle, it is impossible to determine both the momentum and position of an electron
- An orbital describes a region in an atom that has a high probability of containing an electron or a high electron change density.
- Orbitals with the same value of n are in the same principal energy level.
- Aufbau ("building up") Principle is a guide for predicting the order in which electrons fill subshells and shells in atoms.
- According to Pauli's Exclusion Principle no two electrons in the same atom may have identical sets of four quantum numbers

- Hund's Rule states that each orbital of a given subshell is occupied by a single electron before pairing begins.
- Electron configuration refers to the distribution of electrons among orbitals in an atom.
- Elements with similar valence-shell electron configurations fall in the same group of the periodic table.
- Atomic properties vary periodically, when atoms are arranged in terms of increasing atomic number.

Self-Assessment

Part I: Multiple Choice Type Questions

- 1. Which of the following scientists did not contribute to determining the structure of the atom?
 - a. Thomson c. Becquerel
 - b. Rutherford d. Dalton
- 2. What is the frequency in Hz of the gamma radiation from a radioactive cobalt-60 source if its wave length is 1.0×10^{-9} nm?
 - a. 3.3 × 10⁻²⁷ Hz
 b. 3.3 × 10⁻⁸ Hz
 c. 3.0 × 10¹⁸ Hz
 d. 3.0 × 10⁻²⁶ Hz
- 3. What is the energy of the electronic transition associated with the Sodium-D line having a wavelength of 589 nm?
 - a. 6.63 x 10⁻³⁴ J c. 3.38 x 10⁻¹⁹ J
 - b. 1.13 x 10⁻²⁷ J d. 5.82 x 10² J
- 4. Among the following, which color corresponds to light of the highest frequency?
 - a. Green c. Yellow
 - b. Red d. Blue
- 5. Which of the following orbital designations does not exist?

a.	1s	C.	Зр
b.	2d	d.	4f

Part II: Complete the following statements.

- 6. The atomic number gives the number of _____ in the nucleus.
- 7. In an atom, the number of electrons is equal to the number of _____.
- 8. When n = 2, the values of ℓ could be _____ and ____
- When *l* = 1, the values of m_l could be _____, ____ and _____ as well as the subshell has the letter label _____.
- When l = 1, the subshell is called a _____ subshell with _____ total orbitals.

Part III: Short answer questions and Problems

- 11. For the following atoms, determine the number of protons, neutrons, and electrons:
 - a. $\frac{114}{48}$ Cd b. $\frac{98}{43}$ Tc c. $\frac{199}{79}$ AU d. $\frac{222}{92}$ Rn
- 12. Calculate the wavelength of the light emitted when an electron falls from n = 3 to the n = 1 state in hydrogen atom.
- 13. Write the number and the letter for the orbital that corresponds to the following pairs of *n* and *l* quantum numbers:
 - a. n = 3, l = 1c. n = 3, l = 2b. n = 4, l = 0d. n = 5, l = 3
- 14. What type of orbital (i.e., 3s, 4p, ...) is designated by these quantum numbers?
 - a. n = 5, l = 1, $m_l = 0$ c. n = 2, l = 0, $m_l = 0$
 - b. $n = 4, l = 2, m_l = -2$ d. $n = 4, l = 3, m_l = -3$
- 15. Write the ground-state electron configurations for the following elements:
 - a. Br (Z = 35) c. W (Z = 74)
 - b. Mo (Z = 42)

Answer for the Activities of Unit One

ਤ) Activity 1.1

The word "atom" is derived from the Greek word "atomos" meaning indivisible. The Greek philosopher Democritus (460 – 370 B.C.) believed that all matter consists of very small, indivisible particles, which he named "atomos" (meaning uncuttable or indivisible). However, since Aristotle and other prominent thinkers of the time strongly opposed his idea, the concept of atom was overlooked for more than 2000 years.



- The law of conservation of mass can't be violated. If one able to measure the oxygen used by the fire to burn the wood and collects the products of combustion, the mass before and after combustion is exactly the same.
- Dalton's atomic theory was accepted by many scientists almost immediately. Most of it is still accepted today. Those which are inconsistent are these postulates:
 - i. All atoms of a given element are alike in mass and other properties.
 - ii. All matter is composed of extremely small, indivisible particles called atoms. However, scientists now know that atoms are not the smallest particles of matter. Atoms consist of several types of smaller particles, including protons, neutrons, and electrons.

Activity 1.3

 Bone x-ray uses a very small dose of ionizing radiation to produce pictures of any bone in the body. It is commonly used to diagnose fractured bones or joint dislocation. Bone x-rays are the fastest and easiest way for your doctor to view and assess bone fractures, injuries and joint abnormalities.
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2. The discovery of radioactivity contradicted Dalton's atomic theory because Dalton's theory proposed that atoms are indestructible. Radioactivity is the process through which the nuclei of heavy chemical elements disintegrate into the nuclei of larger elements with the release of energy.

Activity 1.4

- a. ${}^{12}_{\ 6}C$ and ${}^{14}_{\ 6}C$ have the same number of protons but different mass number and number of neutrons
- b. ${}^{14}_{6}C$ and ${}^{14}_{7}N$ have the same mass number but different atomic number because they are atoms of different elements.

Activity 1.5

- 1. EMR significant roles in chemistry. For instance, EMR emitted from atoms gives us clues about how electrons arrange in an atom. Several modern technologies are based on EMR, which are generated from different substances.
- 2. EMR can be obtained energy from different sources including the cosmos (e.g., the sun and stars); radioactive elements (ex. U, Ra, Th, ect); and manufactured devices (microwave oven, medical X-rays, ultrasound, etc). All radiant energy obtained from different sources exhibit the same type of wavelike behavior such as wavelength, frequency and amplitude as well as travel at the speed of light in a vacuum.

■ Activity 1.6

When compounds of different elements ignite, they release different wavelengths of light, which translate into different colors. That means each metal has a different pattern of spectral lines and different flame colors. Flame colors are produced from the movement of the electrons in the metal ions present in the compounds. Generally, different elements have different flame colors because their electrons have different allowed energy levels.

Section 1.8: Electronic Configurations and the Periodic Table of the Elements

Activity 1.7				
1. a. n = 1, <i>l</i> = 0	d. n = 3, l = 3	e. n = 4, {	= 4	
b. n = 3, ℓ = 1	c. n = 4, <i>l</i> = 0			
2. a. 3s	b. 3p	c. 7s	d. 3d	
E Activity 1.8				

- 1. Each box in an orbital diagram represents an orbital.
- 2. Similarities: 1s and 2s orbitals occupy 2 electrons each and are spherical in nature.

Difference: 1s and 2s orbitals belong to different energy levels. 2S orbital has a bigger size, but 1s is small.

- 2p_x and 2p_y are oriented differently in space (different magnetic quantum numbers). The 2px orbital lies on the x-axis, while the 2p_y orbital lies along the y-axis.
- 4. It means that the 4d subshell has 6 electrons. 4 represents the fourth energy shell and d is a subshell, and 6 electrons are present in the d orbitals of the subshell.



Element	Atomic	Ground state electron	Condensed electron
	number	configuration	configuration
Sc	21	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ² 3d ¹	[Ar] 4s ² 3d ¹
Ti	22	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ² 3d ²	[Ar] 4s ² 3d ²
V	23	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ² 3d ³	[Ar] 4s ² 3d ³
Cr	24	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ¹ 3d ⁵	[Ar] 4s ¹ 3d ⁵
Mn	25	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ² 3d ⁵	[Ar] 4s ² 3d ⁵
Fe	26	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ² 3d ⁶	[Ar] 4s ² 3d ⁶
Со	27	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ² 3d ⁷	[Ar] 4s ² 3d ⁷
Ni	28	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ² 3d ⁸	[Ar] 4s ² 3d ⁸
Сυ	29	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ¹ 3d ¹⁰	[Ar] 4s ¹ 310
Zn	30	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ² 3d ¹⁰	[Ar] 4s ² 3d ¹⁰

Activity 1.10

Classification helps us to predict the properties of the elements and their compounds based on their positions in the Periodic Table, and vice versa. It becomes easier to study, understand, compare and contrast the related properties among the elements and their compounds from different groups.

Answer key for self-test Exercises

Answer for self-Exercise 1.1

- i. Compounds are formed when atoms of different elements unite in fixed proportions.
 - ii. No atoms are created, destroyed, or broken apart in a chemical reaction.
- 2. The limitations of Dalton's Atomic Theory are:
 - The indivisibility of an atom was proved wrong: an atom can be further subdivided into protons, neutrons and electrons.
 - According to Dalton, the atoms of same element are similar in all respects. However, atoms of some elements vary in their masses and densities. These atoms of different masses are called isotopes.
 - Dalton also claimed that atoms of different elements are different in all respects. This has been proven wrong in certain cases: argon and calcium atoms each have the same atomic mass (40 amu).
 - According to Dalton, atoms of different elements combine in simple whole number ratios to form compounds. This is not observed in complex organic compounds like sugar (C₁₂H₂₂O₁₁).
- Modern chemistry is based on several fundamental laws, including the law of multiple proportions, law of definite proportions, and law of conservation of mass.

Section 1.8: Electronic Configurations and the Periodic Table of the Elements

Answer for self-Exercise 1.2				
 a. 13 protons and 14 neutrons b. 16 protons and 16 neutrons c. 30 protons and 34 neutrons d. 82 protons and 125 neutrons 				
 Element X is Cr; ⁵⁰₂₄Cr, ⁵²₂₄Cr, ⁵³₂₄Cr, and ⁵⁴₂₄Cr 19.9% boron-10 and 80.1% boron-11 Because of the weighted-average atomic mass of lithium, 6.941 amu, is much closer to 7.01600 amu than to 6.01512 amu, lithium-7 must be the more abundant isotope. 				
Self-test Exercise 1.3				
1. $3.0 \times 10^{16} \mathrm{s}^{-1} = 3 \times 10^{16} \mathrm{Hz}$ 2. $3.06 \mathrm{m}$				
Self-test Exercise 1.4				
1. a. i. Elnfrared = 19.89 x 10 ⁻²⁰ J \approx 2.0 x 10 ⁻¹⁹ J ii. Eultraviolet = 19.89 x 10 ⁻¹⁸ J \approx 2.0 x 10 ⁻¹⁷ J iii. Ex-ray = 19.89 x 10 ⁻¹⁶ J \approx 2.0 x 10 ⁻¹⁵ J b. x-ray has the greatest amount of energy per photon c. Infrared has the amount of energy per photon has 2. i. $\lambda_1 = 2.76 \times 10^{-6}$ m $\lambda_2 \approx 0$ $\lambda_3 = 3.43 \times 10^{-7}$ m ii. λ_2 has the shortest wavelength & λ_1 has longer wavelength iii. $v_{o1} = 1.1 \times 10^{14}$ Hz $v_{o2} = 0$ $v_{o1} = 8.7 \times 10^{14}$ Hz				
Self-test Exercise 1.5				
1. 4.3 x 10 ⁻¹⁵ m 2. 1.97 x 10 ¹⁵ Hz				
Self-test Exercise 1.6				
1. i. $E_{n=2} = -4.45 \times 10^{-19} \text{ J}$ ii. $E_{n=3} = -2.42 \times 10^{-19} \text{ J}$ iii) 6.6 x 10 ⁻⁸ m = 66 nm 2. $\lambda = 3.6 \text{ nm}$				

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8-	Self-test Exercise 1.7						
	1. 0.122 nm		2. 56.7 nm		3. 434 nm		
8-	ਡ Self	-test	Exerc	ise (1.8		
1.							
	n		ł		m,	m _s	
		0 (2s)	0		+1/2, -1/2	
	4	1(4p)		-1, 0, 1		$\pm \frac{1}{2}$ for each value of m,	
	4	2(4d)		-2, -1, 0, 1, 2		$\pm \frac{1}{2}$ for each value of m_{ℓ}	
		3(4f)		-2, -	1, 0, 1, 2	$\pm \frac{1}{2}$ for each value of m_{ℓ}	
2.	a.4d b.3p		С) 5p	d) 4f		
3.		n	ł		m,		
	i.	i. 3	0		0		
			1		-1, 0, 1		
			2		-2, -1, 0, 1	, 2	
	ii n	- 3 h	$a = 3^2 -$	9 ort	sitals		
II. $n = 3 \text{ mas } 3^2 = 9 \text{ orbitals}$							
4. 5	Possible			are	1 and 2		
J.	5. Fossible values of (are) and 2.						
0.	6. Ip, 2a and 4t orbitals do not exist.						
8 –	Self-test Exercise 1.9						

1. a. Na (Z =11): $1s^2 2s^2 2p^6 3s^1$ or [Ne]3s¹ b. Al (Z = 13): $1s^2 2s^2 2p^6 3s^2 3p^1$ or [Ne] $3s^2 3p^1$ c. P (Z = 15): $1s^2 2s^2 2p^6 3s^2 3p^3$ or [Ne]3s² 3p³ d. V (Z = 23): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$ or [Ar]4s² 3d¹ e. Mn (Z = 25): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$ or [Ar]4s² 3d⁵ f. Fe (Z =26): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$ or [Ar]4s² 3d⁶

Section 1.8: Electronic Configurations and the Periodic Table of the Elements

- 2. a. N (Z = 7): 1s²2s²2p³: 3 p electrons
 - b. Si (Z = 14): $1s^22s^22p^63s^23p^2$: 6 s electrons
 - c. $S(Z = 16): 1s^22s^22p^63s^23p^4: No d electron$
- 3. c. is unacceptable, because the value of m_s can't be 1.
 - d. is also unacceptable, because for n =2, ℓ value can't be 2

Answer for self-Exercise 1.10

- Here are three main factors that affect the size of the atoms: the nuclear charge of the atom, the shielding effect, and the number of energy levels that hold the electrons.
- a. Sr > Ca > Mg. These three elements are in Group IIA(2), and size decreases up the group.
 - b. K > Ca > Ga. These three elements are in Period 4, and size decreases across a period.
 - c. Rb > Br > Kr. Rb is largest because it has one more energy level and is farthest to the left. Kr is smaller than Br because Kr is farther to the right in Period 4.
 - d. Rb > Sr > Ca. Ca is smallest because it has one fewer energy level. Sr is smaller than Rb because it is farther to the right.
- 3. In a series of transition elements, electrons join the core electrons, not valence shell electrons. As a result, the effective nuclear charge for Cr and, Mn remains more or less constant instead of increasing.

Chemistry Grade 11 Distance Module -I							
Answer for self-Exercise 1.11							
1.	a. As b. Br c. Al d. G						
2.	 Potassium (Z=19) and electron configuration 1s²2s²2p⁶3s²3p⁶4s¹ removes its 4s electron with great ease, attaining the noble gas configuration (Ar), and it is extremely difficult to remove the second electron from a stable configuration. Calcium (Z=20) and electron configuration 1s² 2s² 2p⁶ 3s² 3p⁶ 4s² removes the two electrons to become like that of noble gases in its configuration. This is the reason why the energy difference between 1st and 2nd is small as compared to potassium. 						
3.	a.Ga <ge b.sb<br="" c.k<cr="" d.sr<mg="" e.o<n<="" th=""></ge>						
E	Answer for self-exercise 1.12						
 A high IE1 and a very negative EA1 suggest that the elements are halogens, in Group VIIA (17), which form 1- ions. This is due to the difference in electron configuration of Si and P. An electron added to P must go into a p orbital that already contains an electron. That is the first pairing of electrons begin in the 3p orbitals. The electron that is already there repels the added electron, and less energy is released during this electron addition, so the electron affinity is lower. The valence electron configuration of the alkaline earth metals is ns2, where n is the highest principal quantum number. Therefore, the extra electron must enter the np subshell, which is effectively shielded by the two ns electrons and the inner electrons. Consequently, alkaline earth metals have little tendency to pick up an extra electron 							
E —	Answer for self-exercise 1.13						
1. 2. 3. 4.	Left side of the periodic table. In group 17 (right side of the periodic table). Group 8, period 4. It is a metal and has high tendency to loss the 3s ¹ electron to achieve						

the electron configuration of the neon.



UNIT

CHEMICAL BONDING

Unit Introduction

Dear learner, this unit deals with the different types of chemical bonding, molecular geometry, and theories of chemical bonding. In Grade 9, you have learned about chemical bonding and its types such as ionic, covalent and metallic bonding and their characteristics. In this unit, you will learn some new concepts about chemical bonding, like intermolecular forces, molecular geometry, theories of chemical bonding and much more.

Unit Outecome

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

- explain that a chemical bond is an attractive force between particles
- demonstrate an understanding of the formation and general properties of substances containing ionic, covalent and metallic bond
- draw Lewis structure for simple ionic and covalent compounds
- identify the origin of polarity within molecules
- describe the formation and nature of hydrogen bonds, dipole-dipole forces and London forces
- explain the bonding models (Lewis model, valence bond model and molecular orbital model) and show the usefulness of the bonding theories in explaining and predicting molecular properties (bond angle, bond length, bond energy, etc.)
- explain how the properties of a substance (solid or liquid) depends on the nature of the particles present and the type of intermolecular forces
- discuss the importance of intermolecular forces in plant and animal life
- explain how the Valence Shell Electron Pair Repulsion (VSEPR) model can be used to predict molecular shape.

Unit Content

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The Required Study Time

5 weeks

If you set an average of one hour per day, you will be able to complete unit two by the end of the assigned week. Try to do all the learning activities. Additionally, compare your answers with the given answer at the end of the Module.

If you do not get a particular exercise right in the first attempt, you should not get discouraged, but instead, go back and attempt it over and over again. If you still do not get it right after several attempts, then you should rather seek help from your friend or even your tutor.

Unit Learning Strategies

- You should monitor your performance through planning, monitoring, and self-regulation
- You should manage the learning environment and available resources
- You should try all self-test exercises and written assignments
- You should surf on the internet for better understanding contents in the module.
- You should consult your tutor for further clarifications on some contents in the module.

Section 2.1: Introduction

Dear learner, you have studied about the electronic configuration of atoms of different elements and variation in the periodic properties of elements in unit one. Chemical bonds which join the atoms together will give various types of substances. The discussion will also highlight how these bonds are formed. The properties of substances depend on the nature of bonds present between their atoms. In this section you will learn that ionic compounds such as sodium chloride dissolve in water whereas covalent compounds like naphthalene do not. This is because the type of bonds is present between them are different. In addition to the difference in solubility, these two types of compounds differ in other properties.

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

- define chemical bonding
- explain why atoms form chemical bonds
- illustrate chemical bonding using the octet rule
- describe the types of chemical bonding and the mechanisms of the bonding process.

Why do atoms combine?

Atoms usually combine to reduce their potential energy.

A chemical bond is the attractive force which holds atoms, ions, and molecules together.

2.1.1 Octet Rule

All noble gases except helium (1s²) have ns²np⁶ electron configurations (where n indicates the highest occupied shell).

The noble gases are quite unreactive because they have very stable electron configurations.

Because all the noble gases (except helium) have outer shells with eight electrons, many atoms undergoing reactions also attain eight valence

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electrons (ns²np⁶). This rule has become known as the octet rule.

Octet rule: Atoms tend to gain or lose electrons until they have achieved an outer shell that contains an octet of electrons (eight electrons).

Do you know any compound whose central atom does not obey the octet rule?

2.1.2 Types of Chemical Bonding

Hello, in general, there is a gradual change from metallic to non-metallic character as you move from left to right across a period and from bottom to top within most groups in the periodic table. Accordingly, atoms can combine to form three types of bonds: metal with non-metal (ionic bond), non-metal with non-metal (covalent bond), and metal with metal (metallic bond).

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

l can...

- define chemical bonding
- explain why atoms form chemical bonds
- illustrate chemical bonding using the octet rule
- describe the types of chemical bonding and the mechanisms of the bonding process.



Activity 2.1

Dear learner, try to answer the following questions.

- 1. Why do different atoms form different types of bonding?
- 2. Why do metals tend to form cations and why do non-metals tend to form anions?

Section 2.2: Ionic Bond

Dear learner, in this section you will study the formation of ionic bond and the properties of ionic compounds.

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

- define ionic bonding
- use Lewis electron dot symbols for main group elements
- describe ionic bonding using Lewis electron dot symbols
- list the favor able conditions for the formation of ionic bonds
- explain the formation of ionic bonding
- give examples of ionic compounds
- define lattice energy
- calculate lattice energy of ionic crystals from given data using the Born-Haber cycle
- discuss the exceptions to octet rule
- describe the properties of ionic bonding.



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Why do elements form ions in certain chemical reactions?

An ionic bond is formed by the electrostatic attraction between positive and negative ions.

An ionic bond forms between two atoms when one or more electrons are transferred from the valence shell of one atom to the valence shell of the other.

Cations are positive ions that form when atoms lose electrons, and the atoms that gain electrons are called anions (negative ions).

) Activity 2.2

Dear learner, which of the following pairs of elements are likely to form an ionic compound?

a. sodium and chlorine b. nitrogen and fluorine c. lithium and oxygen

Hello, any given ion tends to attract as many neighboring ions of opposite charge as possible. When large numbers of ions gather together, they form **an ionic solid**. The solid normally has a regular, crystalline structure that allows for the maximum attraction of ions, given their particular sizes.

For example, the formation of NaCl from sodium and chlorine can be explained as:

Na \rightarrow Na⁺ + e⁻ [Ne]3s¹ [Ne] Cl + e⁻ \rightarrow Cl⁻ [Ne]3s²3p⁵ [Ne]3s²3p⁶ or [Ar] Na⁺ + Cl⁻ \rightarrow NaCl or Na⁺Cl⁻

The bond formed, as a result of the electrostatic attraction between the positive and negative ions is called the electrovalent bond or ionic bond.

Ionic compounds are usually formed when metal cations bond with non-metal anions.

Self-test Exercise 2.1

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- 1. Explain the formation of bonds in the following pairs of elements:
 - a. potassium and chlorine b. magnesium and oxygen
 - c. sodium and oxygen
- 2. Which of the following elements will form an ionic bond with chlorine, why?

 Identify the species found in the following ionic compounds: CaCl₂, MgO, and Al₂O₃.

2.2.1 Lewis Electron-Dot Symbols

Why sodium atom highly reactive, but sodium ion does not?

Dear learner, recognize that the American Chemist Gilbert N. Lewis (1875– 1946) developed a special set of symbols for his theory. A Lewis symbol consists of a chemical symbol to represent the nucleus and core (inner-shell) electrons of an atom, together with dots placed around the symbol to represent the valence (outer-shell) electrons. For example, the Lewis symbol for chlorine, which has the electron configuration, [Ne]3s²3p⁵, is





Self-test Exercise 2.2 1. Write Lewis structures for the following compounds: a. BaO b. potassium oxide c. aluminum oxide. 2. Use Lewis symbols to show electron transfer between the following atoms to form cations and anions: a. K and S b. Ca and O c. Al and N

2.2.2 Formation of Ionic Bonds

Dear learner, you may be surprised to learn that the electron-transfer process actually absorbs energy! So why does it occur? As you will see, the reason ionic substances exist at all is because of the enormous release of energy when the ions come together and form a solid.

Consider just the electron-transfer process for the formation of lithium fluoride, which involves two steps:

1. A gaseous Li atom loses an electron (IE_1) :

 $Li(g) \rightarrow Li(g) + e^{-}$ IE₁ = +520 kJ

2. A gaseous F atom gains an electron lost by Li atom (EA):

$$F(g) + e \rightarrow F(g)$$
 EA = -328 kJ

The two-step electron-transfer process by itself requires energy:

$$Li(g) + F(g) \rightarrow Li^{+}(g) + F^{-}(g)$$
 $IE_1 + EA = +192 \text{ kJ}$

Hello, note that the total energy needed for ion formation is even greater than this because metallic lithium and diatomic fluorine must first be converted to separate gaseous atoms, which also requires energy. Despite this, the standard heat of formation (ΔH_f°) of solid LiF is –617 kJ/mol; that is, 617 kJ is released when 1 mol of LiF(s) forms from its elements. That is, there must be some exothermic energy component large enough to overcome the endothermic steps.

These are:

• When 1 mol of $Li^+(g)$ and 1 mol of $F^-(g)$ form 1 mol of LiF(g):

 $Li^+(g) + F^-(g) \rightarrow LiF(g)$ $\Delta H^\circ = -755 \text{ kJ}$

• When the gaseous ions coalesce into a crystalline solid. That occurs because each ion attracts others of opposite charge:

$$Li^+(g) + F^-(g) \rightarrow LiF(s)$$
 $\Delta H^\circ = -1050 \text{ kJ}$

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The negative of this value, 1050 kJ, is the lattice energy of LiF.

The lattice energy (U) is the enthalpy change that occurs when 1 mol of ionic solid separates into gaseous ions.

The lattice energy indicates the strength of ionic interactions, which influence melting point, hardness, solubility, and other properties.

Lattice Energies from the Born-Haber Cycle

Dear learner, the lattice energy of an ionic solid is usually found indirectly from experiment by using the Born-Haber cycle.

Hess's law states that an overall reaction's enthalpy change is the sum of the enthalpy changes for the individual reactions that make it up:

$$\Delta H_{\text{total}} = \Delta H_1 + \Delta H_2 + \Delta H_3 + \dots$$

Consider the Born-Haber cycle for the formation of NaCl. We think that solid sodium chloride being formed from the elements by two different routes, as shown in **Figure 2.1.** In one route, NaCl(s) is formed directly from Na(s) and $\frac{1}{2}Cl_{2}(g)$; $\Delta_{r}H^{\circ} - 411$ kJ mol⁻¹.





(B

The second route consists of the following five steps, along with the enthalpy change for each.

Step1: Metallic sodium is vaporized to a gas of sodium atom:

Na(s) \rightarrow Na(g) ΔH°_{step1} = + 108 kJ mol⁻¹

Step 2: Chlorine molecules are dissociated to atoms:

$$1/_2\text{Cl}_2(g) \rightarrow \text{Cl}(g) \Delta H^\circ_{\text{step2}} = 1/_2 \text{ bond energy of } \text{Cl}_2$$

$$= \frac{1}{2} (240 \text{ KJ}) = + 120 \text{ kJ mol}^{-1}$$

Step 3: Sodium atoms are ionized to Na⁺ ions:

$$Na(g) \rightarrow Na^{+}(g) + e^{-} \Delta H^{\circ}_{step3} = IE_1 = +496 \text{ kJ mol}^{-1}$$

Step 4: Formation of chloride ion:

$$CI(g) + e \rightarrow CI^{-}(g)$$
 $\Delta H^{\circ}_{sten4} = EA = -349 \text{ kJ mol}^{-1}$

Step 5: Formation of NaCl(s) from ions. The ions Na⁺ and Cl⁻ combine to give solid sodium whose enthalpy changes (the lattice energy) is unknown: Na⁺(chloride g) + Cl⁻(g) \rightarrow NaCl(s) $\Delta H^{\circ}_{step5} = U$ (lattice energy) =?

We know the enthalpy formation (Δ_{f} H°) of NaCl (Direct route) and equals -411 kJ mol⁻¹.

Therefore, we can calculate the lattice energy using Hess's law: Solving for U_{NaCl} gives:

 $U_{NaCl} = \Delta_{f}H^{o} - (\Delta H^{o}_{step1} + \Delta H^{o}_{step2} \Delta H^{o}_{step3} + \Delta H^{o}_{step4}) = -411 \text{ kJ mol}^{-1} - [108 \text{ kJ} \text{ mol}^{-1} + 120 \text{ kJ mol}^{-1} + 496 \text{ kJ mol}^{-1} + (-349 \text{ kJ mol}^{-1})]$

–786 kJ mol⁻¹

(F

Ionic solids exist only because the lattice energy exceeds the energy required for the electron transfer.

Self-test Exercise 2.3				
1. Draw a Lewis electron-dot symbol for each atom:				
a. Rb b. As c. I				
2. Give the group number and general electron configuration of an element with each electron-dot symbol				
a. •X b. •X•				
 Use condensed electron configurations and Lewis electron dot symbols to depict the monatomic ions formed from each of the following atoms, and predict the formula of the compound the ions produce. 				
a. Ba and Cl b. Sr and O c. Al and F d. Rb and O				
4. Identify the main group to which X belongs to in each ionic compound formula:				
a. X ₂ O ₃ b. XCO ₃ c. Na ₂ X				
5. For each pair, choose the compound with the lower lattice energy, and explain your choice:				
a. CaS or BaS b. NaF or MgO c. LiCl or CsCl d. BaS or CsC				

Factors Affecting the Formation of Ionic Bonds



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Why is an ionic compound unlikely to consist of ions of noble gases?

Dear learner, you have to know that the formation of ionic bonding is influenced by various factors. Some of the major factors are presented below.

Ionization energy (IE): The lesser the ionization energy, the greater is the ease of the formation of a cation.

Low ionization energy of metallic elements such as alkali and alkaline earth metals favors the formation of an ionic bond.

Electron affinity (EA): A higher electron affinity favors the formation of an anion.

⁷ Elements having higher electron affinity such as halogens favor the formation of an ionic bond.

Thus, low ionization energy of a metal atom and high electron affinity of a non-metal atom facilitate the formation of an ionic bond between them.

Lattice energy: When a cation and an anion come closer, they get attracted to each other due to the electrostatic (coulombic) force of attraction. The electrostatic force of attraction between oppositely-charged ions release a certain amount of energy and an ionic bond is formed. Larger lattice energy would favor the formation of an ionic bond.

> The lattice energy (U) of an ionic compound depends directly on the product of the ionic charges $(q_1 \times q_2)$, and inversely on the distance (r) between them:

$$U \propto \frac{q_1 \times q_2}{r}$$

where q_1 and q_2 are the charges on +ve and –ve ions respectively, and *r* is the distance between the charges q_1 and q_2 .

Small ions having a higher ionic charge will have a larger lattice energy.

\mathbf{K})Self-test Exercise 2.4

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- 1. Given that the ions in LiF and in MgO are of similar size, which compound has the stronger ionic bonding?
- 2. The radii of the sodium and potassium ions are 102 pm and 138 pm, respectively. Which compound has stronger ionic attractions, sodium chloride or potassium chloride?
- 3. In general, how does the lattice energy of an ionic compound depend on the charges and sizes of the ions?

2.2.3 Exceptions to the Octet Rule in Ionic Compounds

There are certain exceptions to the octet rule. We will discuss it here with ionic compounds.

Less than Octet (Central Atom is Deficient of Electrons)

Dear learner, remember that ions of some elements which are near to helium in the periodic table do not obey the octet rule. The tendency of these atoms (H, Li, Be and B) is to attain an arrangement of two electrons like the noble gas He, which is also a stable configuration. The hydride ion (H⁻), lithium ion (Li⁺), beryllium ion (Be²⁺) and boron ion (B³⁺) are isoelectronic with He. Although atoms with less than an octet may be stable, they will usually attempt to form a fourth bond to get eight electrons. For example, BF₃ is stable, but it will form BF₄⁻ when possible.



Activity 2.4

Dear learner, attempt to answer the following questions.

- 1. What requirements are needed for an atom to expand its valence shell?
- 2. Can you give some compounds that contain less than eight electrons around the central atom?

More than Octet (18-Electron Rule)

The ions of some transition and post-transition elements do not usually obey the octet rule.

Why don't the atoms of these elements lose electrons to achieve the noble-gas configurations of ns^2np^6 ?

Hello, you have to know about the transition metals, the 18-electron rule replaces the octet rule, due to the involvement of d orbitals of these atoms. When these atoms form positive ions, electrons are always lost first from the shells with the highest value of the principal quantum number (n). Consider the electron configurations of the ions of the transition elements iron and zinc and the post-transition elements gallium and tin.

Electron Configurations of Iron:

₃₀Fe: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d⁶ ₂₆Fe²⁺: 1s² 2s² 2p⁶ 3s² 3p⁶ 3d⁶ ₂₆Fe³⁺: 1s² 2s² 2p⁶ 3s² 3p⁶ 3d⁵ A stable ion of iron with valence shell electron configuration is 3s² 3p⁶ 3d⁵ which is not isoelectronic with a noble gas. Fe²⁺ is a well-known stable ion with a valence shell electron configuration 3s² 3p⁶ 3d⁶ which is not isoelectronic with any of the noble gases.

Electron configurations of zinc:

 $_{20}$ Zn: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d¹⁰ = [Ar] 4s²3d¹⁰

 $_{30}$ Zn²⁺: 1s² 2s² 2p⁶ 3s² 3p⁶ 3d¹⁰ = [Ar]3d¹⁰

Zn²⁺ is also not isoelectronic with any of the noble gases.

Electron configurations of gallium:

The post-transition element gallium (Ga) loses electrons first from the 4p orbital and then from the 4s orbital to from a Ga³⁺ ion as

$$_{31}$$
Ga: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s²3d¹⁰4p¹ = [Ar] 4s²3d¹⁰4p²

 $_{31}$ Ga³⁺: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s²3d¹⁰ = [Ar]3d¹⁰

On closely examining the electron configurations of Zn²⁺ and Ga³⁺, you will realize that ions have completely-filled outer subshells and a noble gas core. Their valance electron configuration can be generally represented as *ns*²*np*⁶*nd*¹⁰.

X	Self-test Exercise 2.5
1.	The heavier post-transition elements such as Pb and Sn, from which orbital (s) do they lose electrons?
2.	Can you write the electron configurations of Sn, Sn ²⁺ , and Sn ⁴⁺ ? Check if any of these tin ions are isoelectronic with any of the noble gases.
3.	Which group of ions of elements, in general, show more than the octet (18-electron rule)?
4.	How many electrons does a Se atom have to gain to have a complete octet in its valence shell?
5.	Identify the cations that obey the octet rule in the following compounds:
	a. LiF b. NaCl c. CuO d. FeCl ₃

2.2.4 Properties of Ionic Compounds

Experiment 2.1

Investigation of Solubility of Ionic Compounds

Objective: To investigate the solubility of NaCl and CuCl₂ in polar and non-polar solvents

Apparatus and Chemicals: Test tube, Bunsen burner, NaCl, CuCl₂, water, ethanol, cyclohexane, and petroleum ether

Procedure:

- 1. Place about 0.5 g each of NaCl and $CuCl_2$ in to two separate test tubes and add about 2.5 mL of water and shake both test tubes quickly.
- 2. If some residue is left in the test tubes, heat the solutions with a Bunsen burner.

SAFETY: Turn off the Bunsen burners as these chemicals are HIGHLY FLAMMABLE!

3. Repeat Step 1 using fresh ethanol, cyclohexane, and petroleum ether solvents, separately. (SAFETY: Do not heat these, as they are highly flammable!)

Observations and analysis:

Prepare an observation table in your notebook for the solubility of NaCl and $CuCl_2$ in all the four solvents at room temperature and on heating (SAFETY: only heat when water is the solvent) and record the observations.

Solvent	NaCl	CuCl ₂
Water		
Ethanol		
Petroleum ether		
Cyclohexane		

Inference/Conclusion

Interpret the observation table and give results.

Make a generalized statement about the solubility of ionic compounds in polar and non-polar solvents.

Experiment 2.2

Thermal behavior of ionic compounds

Objective: To study the effect of heat on ionic compounds

Apparatus and chemicals: Test tubes, Bunsen burner, test tubes holders, sodium chloride, and copper (II) chloride.

Procedure:

- 1. Take two hard glass test tubes and label them as A and B.
- 2. Add 0.5 g each of dry sodium chloride crystals and copper (II) chloride in test tubes A and B, respectively.
- 3. Hold these test tubes with the help of test tube holders.
- 4. Heat the tubes at the same time in the Bunsen burner flame, first slowly and then strongly. Shake them a little when heating.

Caution: Chlorine gas is produced, this would be best carried out in a fume cupboard.

Observations and analysis:

- A. Do the crystals melt?
- B. Do they have high or low melting points?

Experiment 2.3

Electrical Conductivity of Ionic Compounds

Objective: To test the electrical conductivity of molten compounds

Apparatus and chemicals: 9-volt battery, 6-watt bulb with a bulb holder, conducting wires include: beaker, tripod, wire gauze, clamp and stand, two carbon rods, lead (II) iodide or lead (II) bromide, copper (II), sulphate or sodium chloride. Caution: Lead compounds are harmful. Copper sulphate is corrosive and an irritant.

Procedure A:

- 1. Connect the circuit as shown in **Figure 2.2** in which a 9 volt DC is connected via a bulb.
- 2. Using about 2 cm depth of lead (II) iodide or lead (II) bromide in a small beaker, test the conductivity of the crystals. Do not throw the crystal away and be careful not to contaminate the sample because you will reuse it later. (Note: A fairer test of the solid compound would be to use a lump of the compound rather than its powder).
- 3. Test the conductivity of large crystals of copper (II) sulphate and sodium chloride if any of them are available in your laboratory.

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Procedure B:

- Now heat the same lead (II) iodide or lead (II) bromide (used in the above experiment) in a beaker on a tripod and wire gauze or in a boiling tube supported by a clamp and stand until it melts.
- 2. Test the conductivity of the molten compound by dipping the carbon electrodes (carbon rods) into the molten compound as shown in the **Figure 2.2**.
- 3. Repeat for the other compounds.



Figure 2.2: Conductivity in the molten state

Observations and analysis:

- A. What did you observe?
- B. Which of the compounds (molten or solid) conduct electricity? Why?
- C. All compounds contain at least two elements. Examine the names of those compounds which conduct electricity when in the molten state and decide which groups of the components of these compounds belong. Name the type of bonding that exists in the compounds used.

Experiment 2.4

Electrical Conductivity of Ionic Compounds

Objective: To test the electrical conductivity of the aqueous solutions of some common ionic compounds

Apparatus and chemicals: 9-volt battery, 6-watt bulb with a bulb holder, conducting wires, two carbon rods, H_2O , lead (II) iodide, NaCl. Caution: Lead compounds are harmful.

Procedure:

- 1. Dissolve the compound in 50 mL of water in two separate beakers.
- 2. Connect the same circuit you used in Experiment 2.3 and test the conductivity of each aqueous solution (Figure 2.3).

Observations and analysis:

- A. Predict what happens at the electrodes based on the type of the compound used for the experiment.
- B. Do you expect the same product(s) at the electrodes when electricity passes through molten and aqueous solutions of the compounds?



Figure 2.3: Electrochemical cell showing the conductivity of an aqueous solutio

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Since the ionic compounds contain ions (cations and anions) which are held together by the strong electrostatic forces of attraction, they show the following general characteristic properties:

• Ionic compounds are crystalline solids. The ionic compounds are hard and brittle in nature.

lons are arranged in a regular pattern in the crystal.

- Ionic compounds have high melting and boiling points. The melting point of sodium chloride is 1074 K (801°C) and its boiling point is 1686K (1413°C).
- Ionic compounds conduct electricity in their molten state and in aqueous solutions.



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Why ionic compounds do not conduct electricity in solid state?

CHECKLIST 2.2

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

l can...

•	define chemical bonding	
•	explain why atoms form chemical bonds	
•	illustrate chemical bonding using the octet rule	
•	describe the types of chemical bonding and the mechanisms of the bonding process.	

Section 2.3: Covalent Bonding and Molecular Geometry

Dear learner, in this section, you will study about another kind of bonding called covalent bonding. Covalent bonding is helpful in understanding the formation of molecules. Molecules having similar atoms such as hydrogen (H_2) , chlorine (Cl_2) , oxygen (O_2) , and nitrogen (N_2) are molecules of elements whereas those containing different atom like hydrochloric acid HCl, nitric acid NH₃, methane CH₄, carbon dioxide CO₂ are molecules of compounds.

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

- define covalent bonding
- explain the formation of covalent bonding using examples
- draw Lewis structures or electron dot formulas of some covalent molecules
- illustrate the formation of coordinate covalent bonding using examples
- draw resonance structures of some covalent molecules and polyatomic ions
- discuss the exceptions to the octet rule in covalent bonding
- distinguish between polar and non-polar covalent molecules
- describe the properties of covalent molecules
- describe the valence shell electron pair repulsion theory (VSEPR)
- distinguish between the bonding pairs and nonbonding pairs of electrons
- describe how electron pair arrangements and shapes of molecules can be predicted from the number of electron pairs
- explain why double bonds and lone pairs cause deviations from ideal bond angles
- explain the term dipole moment with the help of a diagram
- describe the relationship between dipole moment and molecular geometry

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- describe how bond polarities and molecular shapes combine to give molecular polarity
- predict the geometrical shapes of some simple molecules on the bases of hybridization and the nature of electron pairs
- construct models to represent shapes of some simple molecules
- define intermolecular forces
- name the different types of intermolecular forces
- explain dipole-dipole interactions
- give examples of dipole-dipole interaction
- define hydrogen bonding
- explain the effect of hydrogen bond on the properties of substances
- give reasons why hydrogen bonding is stronger than ordinary dipoledipole interactions
- explain dispersion (London) forces
- give examples of dispersion forces
- predict the strength of intermolecular forces for a given pair of molecules.

Formation of Covalent Bonding

Hello, consider the formation of a hydrogen molecule (H₂). When two isolated hydrogen atoms come closer , electrostatic interactions begin to develop between them. The two positively charged nuclei repel each other, and the two negatively charged electrons repel each other, but each nucleus attracts both electrons and each electron is attracted to both nuclei (Figure 2.4). If the attractive forces are stronger than the repulsive forces, a covalent bond is formed.

A covalent bond is formed when a pair of electrons is shared between two atoms.

Formation of a bond always results in greater electron density between the nuclei.

(B



Figure 2.4: Attractive and repulsive forces in a covalent H-H bond formation

What happens if the hydrogen atoms are too far apart? What if the hydrogen atoms are too closer?

On a graph of energy versus internuclear distance, the bond length is the H-H distance in the minimum-energy. This is the most stable arrangement (Figure 2.5).



Internuclear distance -

Figure 2.5: Covalent bond formation in H₂

There is an optimum distance between nuclei, called the bond length, where net attractive forces are maximized and the H-H molecule is most stable In the H_2 molecule, the bond length is 74 pm.

NH₃, SO₂, PCI₅, O₃, etc.

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Activity 2.5

Dear learner, attempt to answer the following questions.

- 1. Why doesn't a hydrogen atom lose an electron and form an ionic bond?
- 2. Give the main difference between ionic and covalent bond.

Representation of Covalent Bonds (Drawing Lewis Structures)

Hello, you have to know that the **Lewis structure** is also used to represent covalent bonding. The Lewis structure for a hydrogen molecule formed from hydrogen atoms is:



By sharing two electrons in a covalent bond, each hydrogen effectively has one electron pair and the stable, $1s^2$ electron configuration of helium. Similarly, fluorine has seven valence electrons, and an electron-dot structure for the F_2 molecule shows how a covalent bond can form:



The shared pairs of electrons in a molecule are called bonding pairs. The other electron pairs that stay with on atom and are not shared are called non-bonding pairs or lone pairs.

Do the fluorine atoms in a fluorine molecule (F_2) obey the octet rule?

Dear learner, note that in molecules such as O_2 , N_2 , and many others, the atoms share more than one pair of electrons, leading to the formation of **multiple covalent bonds**. Consider the formation of oxygen molecule (O_2) from the oxygen atoms. The electronic configuration of oxygen atoms is 2, 6. Now each oxygen atom needs two electrons to complete its octet. The two oxygen atoms share two electrons and complete their octet:

(F



: O : : O : sharing of 4 electrons or 2 pairs of electrons

The 4 electrons (or 2 pairs of electrons) which are shared between two atoms of oxygen are present between the two oxygen atoms. So these two pairs of shared electrons can be represented by two bonds between the oxygen atoms. An oxygen molecule can be shown as follows:

 $: \overset{\cdots}{\mathrm{O}} = \overset{\cdots}{\mathrm{O}} :$

The two oxygen atoms are said to be bonded together by two covalent bonds.

A bond consisting of two covalent bonds is known as a double bond.

Steps to writing a Lewis Formula:

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We can write the Lewis formula for a covalent compound of known geometry by using the following steps.

- Determine the total number of valence electrons. The total number of electrons for a molecule is the sum of the valence electrons for each atom. For a polyatomic anion, which has one or more extra electrons, add one electron for each unit of negative charge. For a polyatomic cation, which is missing one or more electrons, subtract one electron for each unit of positive charge.
- 2. Write the skeletal structure. The most electropositive atom usually occupies the central position. Connect bonded atoms with an electron-pair bond (a dash). Hydrogen is an exception; it is always a terminal atom, even when bonded to a more electronegative atom.
- 3. Place electron pairs around terminal atoms so that each (except hydrogen) has an octet. Assign any remaining electrons as lone-pairs around the central atom.
- 4. If at this stage, a central atom has fewer than eight electrons, a multiple bond(s) is likely. Move one or more lone-pair of electrons from a terminal atom(s) to a region between it and the central atom to form a double or a triple bond.

Example 2.1

1. Write the Lewis structure for nitrogen trifluoride (NF_3) in which all three F atoms are bonded to the N atom.

Solution: We follow the steps for writing Lewis structures.

Step 1: The outer-shell electron configurations of N and F are $2s^22p^3$ and $2s^22p^5$, respectively. Thus, there are 5 + (3 × 7), or 26, valence electrons to account for in NF₃.

Step 2: The N atom is less electronegative than F, so in the skeletal structure of NF_{3} , N goes in the center with the more electronegative F atoms bonded to it.



Step 3: We draw a single covalent bond between N and each F, and complete the octets for the F atoms. We place the remaining two electrons on N:



Because this structure satisfies the octet rule for all the atoms, step 4 is not required.

Example 2.2

Write the Lewis structure for the carbonate ion (CO_3^{2-}) .

Solution: We follow the steps for writing Lewis structures and note that this is an anion with two negative charges.

Step 1: The outer-shell electron configurations of C and O are $2s^22p^2$ and $2s^22p^4$, respectively, and the ion itself has two negative charges. So, the total number of electrons is $4 + (3 \times 6) + 2$, or 24.

Step 2: We can deduce the skeletal structure of the carbonate ion by recognizing that C is less electronegative than O. Therefore, it is most likely to occupy a central position as follows:



Step 3: We draw a single covalent bond between C and each O and comply with the octet rule for the O atoms:



This structure shows all 24 electrons.

Step 4: Although the octet rule is satisfied for the O atoms, it is not for the C atom. Therefore, we move a lone pair from one of the O atoms to form another bond with C. Now the octet rule is also satisfied for the C atom:



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Self-test Exercise 2.6

1. Determine the total number of valence electrons for the following:

a. CO_2 b. SO_4^{2-} c. NH_4^{+} d. N_2O_4

- Write a Lewis structure of
 a. nitrogen trichloride, NCl₃
 c. phosphonium ion, PH₄⁺
 b. chlorate ion, ClO₃⁻
- 3. Draw a Lewis structure for SF_4 and HCOOH (formic acid).
- 4. Which of the following atoms cannot serve as a central atom in a Lewis structure O, He, F, H, P? Explain.
- 5. Write a Lewis structure for carbonyl sulphide, COS.

Coordinate-Covalent Bonding

Dear learner, you know that covalent bonds form when two atoms each contribute one electron. However, bonds can also form when one atom donates both electrons (a lone pair) to another atom that has a vacant valence orbital. Such kind of bonds is called coordinate covalent bonds or dative bonds. For example, the ammonium ion (NH_4^+) forms when the lone-pair electrons from the nitrogen atom of ammonia, NH_3 , bond to H^+ .

$$H^{+} + : \stackrel{H}{\underset{H}{\dots}} H \longrightarrow \begin{bmatrix} H\\ H: N: H\\ H\\ H \end{bmatrix}^{+}$$

Other examples of molecules which contain coordinate-covalent bonds include NH_3 :BCl₃, H₃O⁺and AlCl₄⁻.

Coordinate covalent bonds are formed when the bonding pair of electrons is donated by just one of the bonding atoms.

Resonance Structures

Hello, you have to know that sometimes more than one valid Lewis structure is possible for a given molecule. For example, two Lewis structures are possible for ozone (O_3) :



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Structures **A** and **B** do not represent two different O_3 molecules. They are just different Lewis structures for the same molecule. However, neither one of these two Lewis structures accounts for the known bond lengths in O_3 (1.278 Å).



Bond length and bond energy measurements indicate that the two oxygenoxygen bonds in O_3 are identical, with properties that lie between those of an O-O bond and an O=O bond, something like a "one-and-a-half" bond. The molecule is shown more correctly with two Lewis structures, called resonance structures (or resonance forms), and a two-headed resonance arrow (\leftrightarrow) between them.



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Resonance structures have the same relative placement of atoms but different locations of bonding and lone electron pairs.

You can convert one resonance form to another by moving lone pairs to bonding positions, and vice versa:



The different structures used to represent the molecule or ion are called resonance forms or resonance contributors or resonance structures.
Example 2.3

Write two equivalent Lewis structures for the nitrate ion, NO_2^{-} . Describe its resonance hybrid structure.

Solution: In NO_2^- there are 5 + 2(6) + 1 = 18 valence electrons. Indicating the single bonds gives the structure

O-N-O

The remaining 14 electrons (18 - 4) can be distributed to produce these structures:



The electronic structure of the molecule is correctly shown by the average of the two. There are two equivalent N–O bonds, each one intermediate between a single and a double bond.

Self-test Exercise 2.7

- 1. Write three equivalent structures for the SO_3 molecule that obey the octet rule.
- 2. Draw Lewis structures of all the important resonance forms of each of the following:
- 3. a. NO_2F (N central) b. HNO_3 c. NO_2^-
- 4. How many structural formulas can you write for each of the following?
- 5. a. H_2O b. NH_4^+ c. CO_3^{2-} d. C_6H_6 (Benzene)
- 6. What is the difference between a resonance structure and a resonance hybrid?

Exceptions to the Octet rule in Covalent Bonding

Although many Lewis structures follow the octet rule, there are exceptions. There can be categorized into three groups:

Less than octet (central atom is deficient of electrons):

Electron-deficient compounds are compounds in which an element has an incomplete octet.

Hello, note that the central atoms of such molecules have fewer than eight electrons (below octet). This group consists of molecules containing central atoms from Groups IIA and IIIA. For example, BeCl₂, BF₃ and AlCl₃, whose Lewis formulas are shown:



More than octet (central atom has excess of electrons):

Dear learner, elements with an expanded valence (also called an expanded octet) have more than 8 electrons (often 10 or 12) in a Lewis structure. Elements in the third period and below, such as phosphorus, sulfur, and bromine, often have an expanded valence because of their larger radii (when compared to the second-row elements) and the availability of empty d orbitals in the valence shell. For example, PF_5 , SF_6 and XeF_4 :



(g





10 electrons around P

1 P 12 electrons around S



Self-test Exercise 2.8

- 1. Suggest the Lewis-like structure for NO_2 and CIO_2 molecules.
- 2. The following do not obey the octet rule. Draw a Lewis structure for each one and state the type of octet rule exception:

```
a. BH_3 b. AsF_4^- c. SeCl_4 d. PF_6^-
```

3. Which of the following atoms can expand its valance shell: F, S, H, Al?

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Polar and Non-Polar Covalent Molecules

What does the word "dipole" refer to? How does electronegativity relate to bond polarity?

When the atoms are alike in a covalent bond, as in the case of the H–H bond of H_2 , the bonding electrons are shared equally. That is, the electrons spend the same amount of time near each nucleus. The bond formed is called a non-polar covalent bond.

In a nonpolar covalent bond electrons are shared equally between two atoms.

In a covalent bond between atoms of different electronegativity's, there is an unequal sharing of an electron pair and the electrons spend more of their time around the more electronegative atom. Such a bond is said to be polar covalent bond. The H–F bond is a polar bond because fluorine is more electronegative than the hydrogen atom. The bonding electrons spend more time near the fluorine atom than the hydrogen atom. The polar nature of a bond is shown like this:

 $\mathsf{H}^{\delta^+}-\mathsf{F}^{\delta^-}$

The bond polarity can also be shown by drawing an arrow so that the head points toward the negative end (F) of the bond and the crossed tail indicates the positive end (H).

∔→ H**−−**F

For a diatomic molecule having a polar covalent bond, such as HF, we can describe a quantity called the dipole moment.

⁷ Dipole moment is a vector sum of the bond moments in a molecule.

Bond moment is a measure of polarity of a diatomic covalent bond. The dipole moment (μ) is defined as the product of the magnitude of the charge (δ) at either end of the dipole multiplied by the distance (d) that separates the charge.

μ= δ× d

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A very polar molecule is one with a large dipole moment, while a non-polar molecule will have a zero-dipole moment.

For a diatomic molecule, the bond moment is the dipole moment. The dipole moment of a polyatomic molecule (three or more atoms) depends on the geometry of the molecule. If the bond moments are equal in magnitude but opposite in direction, then they will cancel each other and the resultant dipole moment will be zero, provided that the vector sum of the bond moments is zero.

🛠) Self-test Exercise 2.9

- 1. Both CO_2 and BCI_3 have zero dipole moments, but the C = O and B CI bond moments are not zero. Explain.
- 2. Compare the dipoles in the covalent molecules HF, HCI, HBr, and HI.
- For each of the following bonds, indicate the positive end with δ⁺ and the negative end with δ⁻. Draw an arrow to show the dipole for each.
 a. N and F
 b. Si and Br
 c. C and O
 d. P and Br
 e. N and P

Properties of Covalent Compounds

Do you expect that carbon tetrachloride conduct electricity?

Experiment 2.5

Investigating the melting point of covalent compounds

Objective: To investigate the melting point of naphthalene.

Apparatus and chemicals: Thermometer, stirrer, beaker, melting point tube, naphthalene, glycerine, Bunsen burner.

Procedure:

- 1. Set up the apparatus shown in Figure 2.6.
- 2. Place a small amount of naphthalene in the melting point tube.
- 3. Attach the tube to the side of the thermometer (the liquid in the beaker will hold the tube in position).
- 4. Heat the beaker slowly. When the naphthalene melts, record the reading on the thermometer.



Figure 2.6: Determination of melting point of naphthalen

Observations and analysis:

Check your result by referring to standard melting point values. Discuss possible reasons for the differences in melting point compared to the standard melting point.

Experiment 2.6

The effects of electricity and some solvents on covalent compounds

Objective: To test the effects of electricity and some solvents on covalent compounds.

Apparatus and chemicals: Beakers, stirrer, test tubes, table sugar, cyclohexane, petroleum ether, water, alcohol, naphthalene.

Procedure A: Solubility test

- Place about 0.5 g or 5 drops each of naphthalene, table sugar, and ethanol in three separate test tubes and add about 2.5 mL of distilled water and shake the test tubes vigorously.
- 2. See a change in the amount of naphthalene, and table sugar added in each case; if some residue remains in the test tube/s, heat it on a Bunsen burner.

- 3. Observe if a separating layer is formed between the ethanol and water, it indicates ethanol is not soluble in water.
- 4. Repeat the experiment using cyclohexane and petroleum ether solvents, separately (Caution: do not heat it.)

Procedure B: Conductivity test

- Connect the circuit as shown in Figure 2.2, (Experiment 2.3) in which a 9-volt DC is connected via a bulb.
- 2. Test the conductivity of table sugar.
- 3. Dissolve a spoonful of table sugar in 50 mL of distilled water in a beaker. Test the resulting solution for electrical conductivity using the setup of of **Figure 2.3**, in **Experiment 2.4**.

Observations and analysis:

Record your observations on your notebook. Copy and complete the table below.

Property			
Substance	Solubility in water	Solubility in non-po- lar solvent	Effect of electric- ity
Naphthalene			
Table sugar			
Ethanol			

Results and Discussion:

- 1. Draw your conclusion from the above observations.
- 2. Do these substances conduct electricity either in the solid state or in the molten (liquid) state?
- 3. Is there any chemical reaction that takes place at the electrodes?

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The covalent compounds consist of molecules which are electrically neutral in nature.

Dear learner, you have to recognize that the forces of attraction present between the molecules are less strong as compared to the forces present in ionic compounds. Therefore, the properties of the covalent compounds are different from those of the ionic compounds. The characteristic properties of covalent compounds are given below:

- Covalent compounds exist as separate molecules because electrically neutral atoms form them and the forces of attraction between these molecules are relatively weak.
- Due to weak intermolecular forces, many covalent molecules or covalent compounds are liquids or gases at room temperature. However, some covalent molecules like iodine are solids at room temperature.
 - a. Liquid H_2O , $Br_{2'}$
 - b. Gas CO₂, H₂, Cl₂, NH₃.
- Covalent compounds are volatile.
- Generally, they have low melting points and boiling points.
- Covalent compounds are generally insoluble in water. Most covalent compounds are soluble in non-polar solvents.
- Non-polar covalent compounds are non-electrolytes because they do not conduct electricity.

2.3.1 Molecular Geometry

Dear learner, there is a simple procedure that enables us to predict the 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.

Molecular geometry is the three-dimensional arrangement of atoms in a molecule.

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Valence Shell Electron Pair Repulsion (VSEPR) Theory

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.

The basic principle of VSEPR theory is that the pair of valence-shell electron pairs around the central atom arrange as far away from one another as possible to minimize electron-pair repulsions.

Electron Pair Arrangement and Molecular Shape

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Hello, recognize that the electron pair arrangement is defined by the orientation/distribution of pairs of electrons, both bonding and non-bonding (lone-pair), around the central atom.

The **molecular shape** is defined by the relative positions of the atomic nuclei. Molecular shapes that occur when all the surrounding electrons are bonding differ from molecular shapes when some of the electrons are non-bonding.

The same electron set arrangement can give rise to different molecular shapes. For the prediction of geometrical shapes of molecules with the help of VSEPR theory, it is convenient to divide molecules into two categories as molecules in which the central atom has no lone pair and molecules in which the central atom has one or more lone pairs. (B

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i. Molecules in which the central atom has no lone pair

Dear learner, consider molecules that contain atoms of only two elements, A and B, of which A is the central atom. These molecules have the general formula ABx, where x is an integer 2, 3, (If x = 1, we have the diatomic molecule AB, which is linear by definition.) In the vast majority of cases, x is between 2 and 6.

In the compounds of AB₂, AB₃, AB₄, AB₅ and AB₆, the arrangement of electron pairs and the B atoms around the central atom A are: linear, trigonal planar, tetrahedral, trigonalbipyramidal and octahedral, respectively.

Such arrangements can be seen in the molecules such as $BeCl_2 (AB_2)$, $BF_3 (AB_3)$, $CH_4 (AB_4)$ and $PCl_5 (AB_5)$ as shown below **(Figure 2.7)** by their ball and stick models.





In the case of multiple (double or triple) bonds the counting is not different, because a double bond is considered as one set and a triple bond as another set of electrons. For example, in the case of CO_2 , O=C=O, there are two sets of electrons around the central atom, acetylene, C_2H_2 , H-C=C-H, two sets of electrons around each of the central atoms.

ii. molecules in which the central atom has one or more lone pairs

In such molecules there are three types of repulsive forces—those which are between bonding pairs, those which are between lone pairs, and those which are between a bonding pair and a lone pair. In general, according to the VSEPR model, the repulsive forces decrease in the following order: Lone pair vs lone pair > lone pair vs bonding pair > bonding pair vs bonding pair

Hello, note that to keep track of the total number of bonding pairs and lone pairs, we show molecules with lone pairs as AB_xE_y , where A is the central atom, B is a surrounding atom, and E is a lone pair on A. Both x and y integers; x = 2, 3, . . ., and y = 1, 2, Thus, the values of x and y indicate the number of surrounding atoms and number of lone pairs on the central atom, respectively. The simplest such molecule would be a triatomic molecule with one lone pair on the central atom and the formula is AB_yE .

AB₂E: Sulfur Dioxide (SO₂)

The Lewis structure of sulfur dioxide is:

ö=s=ö

Because VSEPR treats double bonds as though they were single, the SO_2 molecule can be viewed as consisting of three electron pairs on the central S atom. Of these, two are bonding pairs and one is a lone pair. Thus, the overall arrangement of three electron pairs is trigonal planar. But because one of the electron pairs is a lone pair, the SO_2 molecule has a "bent" shape.



Because the lone-pair versus bonding-pair repulsion is greater than the bonding-pair versus bonding-pair repulsion, the two sulfur-to-oxygen bonds are pushed together slightly and the OSO angle is less than 120°.

AB₃E: Ammonia (NH₃)

Dear learner, recognize that the ammonia molecule contains of three bonding pairs and one lone pair:



The overall arrangement of four electron pairs is tetrahedral. But in NH_3 one of the electron pairs is a lone pair, so the geometry of NH_3 is trigonal pyramidal. Because the lone pair repels the bonding pairs more strongly, the three NH bonding pairs are pushed closer together:





Activity 2.6

Dear learner, try to answer the following questions.

- 1. Why is that repulsions exerted by lone pair is greater than that of bonding pair?
- 2. Water and ammonia are triatomic molecules, the geometry of NH3 is pyramidal whereas that of H2O is bent. Why do you think this is?

AB₂E₂: Water (H₂O)

A water molecule contains two bonding pairs and two lone pairs:

н-о-н

The overall arrangement of the four electron pairs in water is tetrahedral, the same as in ammonia.



AB_4E : Sulfur Tetrafluoride (SF₄)

The Lewis structure of SF_4 is:



The central sulfur atom has five electron pairs whose arrangement is trigonal bipyramidal.

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In the SF_4 molecule, however, one of the electron pairs is a lone pair, so the molecule must have one of the following geometries:



Hello, note that in (a) the lone pair occupies an equatorial position, and in (b) it occupies an axial position. The axial position has three neighboring pairs at 90° and one at 180°, while the equatorial position has two neighboring pairs at 90° and two more at 120°. The repulsion is smaller for (a), and indeed (a) is the structure observed experimentally.

Guidelines for Applying VSEPR Model

Dear learner, you have to recognize the guidelines for applying the VSEPR model given below:

- 1. Write the Lewis structure of the molecule, considering only the electron pairs around the central atom (that is, the atom that is bonded to more than one other atom).
- Count the number of electron pairs around the central atom (bonding pairs and lone pairs). Treat double and triple bonds as though they were single bonds. If there is 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, note that a lone pair repels another lone pair or a bonding pair more strongly than a bonding pair repels another bonding pair.

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Molecular Shape and Molecular Polarity

) Why is it useful to know the polarities of molecules? Are all unsymmetrical molecules polar?

Dear learner, you have to know that many aspects of molecule's chemical behaviour can be understood if we know the geometry (*shape*) of a substance. Molecular shape affects many properties of the molecule like molecular polarity, which in turn influence melting and boiling points, solubility, and even reactivity.

Molecular polarity is created by molecules with a net imbalance of charge.

In molecules with more than two atoms, both shape and bond polarity determine the molecular polarity.

Bond Polarity and Dipole Moment

Diatomic molecules containing atoms of different elements (for example, HCl, CO, and NO) have dipole moments and are called polar molecules.

Diatomic molecules containing atoms of the same element are non-polar molecules because they do not have dipole moments.

For a molecule made up of three or more atoms, both the polarity of the bonds and the molecular geometry determine whether there is a dipole moment.

Polar bonds do not necessarily lead to polar molecules.

For example, the large electronegativity difference between C and oxygen makes each C–O bond quite polar. However, because carbon dioxide (CO₂) has a linear shape, and its bonds are directed 180° from each other, so there is no net dipole moment (μ); μ = 0 D.



Equal dipole moments with opposite direction cancel each other.

Hello, another molecule with identical atoms bonded to the central atom is water. Unlike carbon dioxide, water has a significant dipole moment (μ = 1.87 D). In each O–H bond, electron density is pulled toward the more electronegative O atom, but the bond polarities do not counter-balance each other, because the water molecule is V-shaped. Instead, the bond polarities partially reinforce each other, and the oxygen end of the molecule is considerably more negative than the other end.



) Predict whether ammonia molecule is polar or non-polar and show the direction of bond dipoles and the overall molecular dipole.

Dear learner, note that carbon dioxide and water demonstrate how molecular shape influences polarity. When two or more different molecules have the same shape, the nature of the atoms surrounding the central atom can have a major effect on the polarity of a molecule. Consider tetrachloromethane (CCl_4) and trichloromethane $(CHCl_3)$, two AB_4 type molecules: tetrahedral shape with different polarities. In CCl_4 , the surrounding atoms are all Cl atoms. Although each C–Cl bond is polar, the molecule is non-polar because the individual bond polarities counter-balance each other. In $CHCl_3$, an H atom substitutes for one of the Cl atoms, disturbing the balance and giving chloroform a significant dipole moment.

$$Cl \qquad H \\ \downarrow \\ Cl \rightleftharpoons C \rightleftharpoons Cl \qquad Cl \rightleftharpoons C \rightleftharpoons Cl \\ \downarrow \\ Cl \qquad Cl \qquad Cl \rightleftharpoons C \rightleftharpoons Cl \\ \downarrow \\ \mu = 0 D \qquad \mu = 1.02 D$$

Hello, consider the two constitutional isomers of dichloroethene $(C_2H_2Cl_2)$. They have the same molecular formula but different physical and chemical properties. VSEPR theory predicts that all the nuclei lie in the same plane with a trigonal planar molecular shape around each carbon atom.



Note that the *trans* isomer has no dipole moment (μ = 0 D) because the C – CI bond polarities balance each other. In contrast, the *cis*-isomer is polar (μ = 1.89 D) because the bond dipoles partially reinforce each other, with the molecular dipole pointing between the CI atoms.

Self-test Exercise 2.10

- Determine the molecular shape and ideal bond angles in:
 a. COCl₂
 b. PCl₃
 c. SF₂
 d. SO₄²⁻
 e. CS₂
- 2. When is the molecular shape and the electron-set arrangement the same?
- 3. Arrange the following AFn species in order of increasing F–A–F bond angles. BF3, BeF2, CF4, NF3, OF2
- 4. In the gas phase, phosphorus pentachloride exists as separate molecules. In the solid phase, however, the compound is composed of alternating PCI4+ and PCI4- ions. How does the molecular shape change as PCI5 solidifies? How does the angle change?
- 5. For molecules of general formula ABn (where n > 2), how do you determine if a molecule is polar?

2.3.2 Intermolecular Forces in Covalent Compounds

Dear learner, you have to know that there are two types of forces that hold matter together. These are intramolecular and intermolecular forces. Intramolecular force is a chemical bond (ionic, covalent or metallic) that exists within a particle (molecule or polyatomic ion) and affects the chemical property of the species.



Intermolecular forces are relatively weak as compared to intramolecular forces, because they typically involve lower charges that are farther apart. However, the strength of the intermolecular forces is important because they affect physical properties of the species such as melting point and boiling point.



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Activity 2.7

Dear learner, attempt to answer the following questions.

- 1. Name some intermolecular forces.
- 2. Why is CH4 a gas at room temperature whereas H2O is a liquid?
- 3. Ice is lighter than water. Why?

Hello, three types of attractive forces are known to exist between neutral molecules: dipole-dipole forces, London (or dispersion) forces, and hydrogen bonding forces.

The term Van der Waals forces are a general term for those intermolecular forces that include dipole-dipole and London forces.

Van der Waals forces are the weak attractive forces in a large number of substances, including Cl_2 and Br_2 .

Dipole-dipole forces

Dear learner, note that dipole-dipole forces act between the molecules possessing permanent dipole. When polar molecules are brought near one another, their partial charges act as tiny electric fields that orient them and give rise to dipole-dipole forces; the partially positive end of one molecule attracts the partially negative end of another. Ends of the dipoles possess "partial charges" and these charges are shown by Greek letter delta (δ).



This interaction is stronger than the London forces but is weaker than ion-ion interaction because only partial charges are involved. The attractive force decreases with the increase of distance between the dipoles.

Hydrogen Bonding

Dear learner, hydrogen bonding occurs in polar molecules containing hydrogen atoms bonded to highly electronegative atoms of nitrogen, oxygen, or fluorine. The atom sequence that leads to an H bond (*dotted line*) is indicated below.

Hydrogen bonds are the strongest type of attractive forces between polar covalent molecules.

What is the difference between hydrogen bonds and covalent bonds?



Examples of H-bonding

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Dispersion or London Forces

Dear learner, do you know that the electrons in a non-polar covalent molecule are distributed symmetrically? The movement of the electrons may place more of them in one part of the molecule than another, which forms a temporary dipole. These momentary dipoles align the molecules so that the positive end of one molecule is attracted to the negative end of another molecule.



These weak attractive forces between molecules resulting from the small, instantaneous dipoles that occur because of the varying positions of the electrons during their motion about nuclei are called London forces or dispersion forces.

London forces are present between all particles (atoms, ions, and molecules).

Although dispersion forces are very weak, they make it possible for non-polar molecules like CO_2 , Cl_2 , noble gases, etc. to form liquids and solids.

) London forces tend to increase with molecular weight, why do you think this is?

Hello, note that for molecules of about the same molecular weight, the strength of the dispersion forces is affected by molecular geometry (shape). Shapes that allow more points of contact have more area over which electron clouds can be distorted, so stronger attractions result. This is probably the reason why *n*-pentane has a higher boiling point (36 °C) than its isomers 2-methylbutane (28 °C), and 2, 2-dimethylpropane (9.5 °C).

Self-test Exercise 2.11

1. In which of the following substances do hydrogen bonds occur? Explain by using diagrams.

a.
$$CH_4$$
 b. CH_3CH_2OH
c. $R-CH-C-OH$ d. CH_3-O-CH_3

2. What kinds of intermolecular forces (London, dipole–dipole, hydrogen bonding) would you expect in the following substances?

a.
$$CHCI_3$$
 b. H_2O c. F_2 d. HBr

 Identify the dominant intermolecular force that is present in each of the following substance and select the substance with the higher boiling point in each pair:

4. Compare intermolecular forces with intramolecular bonding.

\checkmark CHECKLIST 2.3

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

l can...

- define covalent bonding
- explain the formation of covalent bonding using examples
- draw Lewis structures or electron dot formulas of some covalent molecules
- illustrate the formation of coordinate covalent bonding using examples
- draw resonance structures of some covalent molecules and polyatomic ions
- distinguish between polar and non-polar covalent molecules
- describe the properties of covalent molecules
- describe the valence shell electron pair repulsion theory (VSEPR)
- predict the geometrical shapes of some simple molecules on the bases of hybridization and the nature of electron pairs
- construct models to represent shapes of some simple molecules
- name the different types of intermolecular forces
- explain the effect of hydrogen bond on the properties of substances
- explain dispersion (London) forces
- predict the strength of intermolecular forces for a given pair of molecules.

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Section 2.4. Metallic Bonding

Dear learner, in the previous sections you have learnt ionic and covalent bonding. The atoms in a metal are held together in a different way. Any ideas we have about the bonding and structure of metals must be able to explain their properties. Do you remember all the properties of metals?

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

- explain how a metallic bond is formed
- explain the properties of metals related to the concept of bonding.

2.4.1 Formation of Metallic Bonding



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Which types of bond exist in solid copper?

Metallic bonds are the chemical bonds that hold atoms together in solid metals such as copper, iron, and aluminum.

In these metals, each metal atom is bonded to several neighboring atoms.

In metallic bonding the electrons freely delocalize among a lattice of positively charged metal ions.



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How is a metallic bond differ from covalent and ionic bonds?

The strength of the metallic bond depends on:

- 1. the number of electrons in the delocalized "sea" of electrons. More delocalized electrons result in a stronger bond and a higher melting point.
- 2. the packing arrangement of the metal atoms. The more closely packed the atoms are the stronger the bond is and the higher the melting point.

2.4.2 Electron-Sea Model

Hello, the electron-sea model is a very simple model, which pictures the metal as an array of metal cations in a "sea" of electrons, as illustrated in **Figure 2.8**. It proposes that all the metal atoms in a sample pool their valence electrons to form an electron "sea" that is delocalized throughout the piece. The metal ions (nuclei plus core electrons) are submerged within this electron sea in an orderly array.



Figure 2.8: The electron-sea model for the electronic structure of metals

2.4.3 Properties of Metals and Bonding

Dear learner, think of some of the things around your home which are made of metal.

Do you know which metals of these things are made from? Which properties make metals good for these uses?

The general properties of metals include malleability and ductility and most are strong and durable. Metals have the ability to stretch into wire and hammered into a thin sheet without breaking. Their strength indicates that the atoms are difficult to separate.



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To what do properties of metals like malleability and ductility refer?

Metals are good conductors of heat and electricity.

Activity 2.8

Dear learner, attempt to answer the following questions.

- 1. List four physical characteristics of a solid metal.
- 2. What is delocalization of electrons?

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

l can...

- explain how a metallic bond is formed
- explain the properties of metals related to the concept of bonding.

Section 2.5: Chemical Bonding Theories

Dear learner, do you know that **the Lewis model** is one of the earliest models of covalent bond formation? However, it fails to explain the formation of chemical bonds. It also gives no idea about the shapes of polyatomic molecules. Similarly, the VSEPR theory gives the geometry of simple molecules but, theoretically, it does not explain them and it has limited applications. Covalent bonds are more accurately explained using modern bonding theories: the valence bond theory and the molecular orbital theory.

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

- name two chemical bond theories
- explain the valence bond theory
- distinguish the Lewis model and the valence bond model
- discuss the overlapping of orbitals in covalent bond formation
- explain hybridization
- show the process of hybridization involved in some covalent molecules
- draw hybridization diagrams for the formation of sp, sp², sp³, sp³d and sp³d² hybrids

- suggest the kind of hybrid orbitals on the basis of the electron structure of the central atom
- predict the geometrical shapes of some simple molecules on the basis of hybridization and the nature of electron pairs
- discuss the hybridization involved in compounds containing multiple bonds
- explain bond length and bond strength
- explain molecular orbital theory
- describe molecular orbital-using atomic orbitals
- describe bonding and anti-bonding molecular orbitals
- draw molecular orbital energy level diagrams for homonuclear diatomic molecules
- write the electron configuration of simple molecules using the molecular orbital model
- define bond order and determine the bond order of some simple molecules and molecule-ions
- determine the stability of a molecule or an ion using its bond order
- predict magnetic properties of molecules.

The valence bond (VB) theory assumes that the electrons in a molecule occupy atomic orbitals of the individual atoms.

It gives us a picture of individual atoms taking part in the bond formation.

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The molecular orbital (MO) theory assumes the formation of molecular orbitals from the atomic orbitals.

Which theory explains all aspects of bonding?

2.5.1 Valence Bond (VB) Theory

Valence Bond Theory states that a covalent bond forms by overlapping of two atomic orbitals. The overlap region, which is between the nuclei, is occupied by a pair of electrons.

Hello, consider the formation of the H₂ molecule from two hydrogen atoms. Each atom has the electron configuration 1s¹. As the H atoms approach each other, their 1s orbitals begin to overlap and a covalent bond forms **(Figure 2.9)**.

Electron density is higher in the overlap region than anywhere else.

The build-up of negative electron charge between the positive nuclei provides the electrostatic attraction that holds the atoms together.



Figure 2.9: Atomic orbital overlap and bonding in H₂

Dear learner, when two atoms are brought closely together, the repulsion of the atomic nuclei becomes more important than the electron-nucleus attraction and the bond becomes unstable. There is a condition of optimal orbital overlap that leads to a maximum bond strength (bond energy) at a particular internuclear distance (bond length). The strength of the bond between any two atoms depends on the extent of the overlap between the two orbitals.

Hello, do you know that the extent of overlap depends on the shapes and directions of the orbitals? An s orbital is spherical, but p and d orbitals have more electron density in one direction than in another. Thus, whenever possible, a bond involving p or d orbitals will be oriented in the direction that maximizes overlap.

In the HF bond, for example, the 1s orbital of H overlaps the half-filled 2p orbital of F along the long axis of that orbital **(Figure 2.10).** Any other direction would result in less overlap and, thus, a weaker bond.

A single bond consists of two electrons of opposite spin.

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Hence, the two atomic orbitals that give rise to a single bond can have no more than two electrons in total.



Figure 2.10: Orbital and spin pairing in the formation of a HF molecule

Similarly, in the F–F bond of $F_{2^{\prime}}$ the two 2p orbitals interact end-to-end, that is along the orbital axes, to attain maximum overlap (Figure 2.11).



Figure 2.11: Orbital and spin pairing in the formation of a F_2 molecule

When two atomic orbitals overlap to build up electron density along the two axis between the two nuclei, the resulting localized bond is called a sigma bond (σ).



2. Identify the overlapping orbitals that form the single bond in a BrCl molecule.

Dear learner, so far we have considered only sigma bonds, bonds in which the highest electron density lies along the bond axis in between the bonded nuclei. The shape of p atomic orbitals allows them to overlap not only in a head-to-head fashion. If they are oriented in a parallel position, they can overlap sideways or laterally as shown in **Figure 2.12.** A bond formed in this way is called **pi-bond** (π).

UNIT TWO





Hello, note that the directional nature of p orbital allows to overlap in two ways: **end-to-end** or **sideways**. These two modes give rise to the two types of covalent bonds: **sigma (\sigma)** and **pi (\pi) bond**, respectively. However, s - s and s - p will always overlap along the nuclear axis, resulting only in sigma bonds.



All of the electron density in a π -bond is found in lobes above and below the line connecting the atomic centers: one above and one below the sigma bond axis. One π -bond holds two electrons that occupy both regions of the bond.

A double-bond always consists of one σ and one π bond. A triple bond consists of one σ and two π bonds.



Dear learner, try to answer the following questions

- 1. Which bond is stronger: sigma (σ) or pi (π)? Why?
- 2. Compare the bond strength and length of single, double, and triple bonds.

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Hybridization of Orbitals

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Dear learner, did you recognize that the number of covalent bonds a carbon atom supposed to form is different from that of carbon actually forms? This discrepancy and molecular structure, specifically bond angle, can be explained using the idea of hybridization.

Hybridization is an imaginary mixing process in which the orbitals of an atom rearrange themselves to form new orbitals called **hybrid** orbitals.

Hello, consider covalent bond formation of carbon atom. From its ground state configuration, two unpaired electrons in the 2p subshell are observed.



Ground state C:

 \checkmark The simplest stable hydrocarbon is methane, CH₄.

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 electron configuration is that of an excited state having energy greater than the configuration in the ground state.



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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.

⁷ 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. Note that it does not explain the reason for the shape.



Dear learner, answer the following by saying true or false.

- 1. Two σ bonds comprise a double bond.
- 2. A triple bond consists of one π bond and two σ bonds.
- 3. Bonds formed from atomic s orbitals are always σ bonds.
- 4. A π -bond consists of two pairs of electrons.
- 5. End-to-end overlap results in a bond with electron density above and below the bond axis

sp hybrid orbitals

Dear learner, recognize that the mixing of one s and one p orbital results in the formation of two equivalent sp hybrid orbitals.

Each sp hybrid orbitals has 50% s-character and 50% p-character.

Such a molecule in which the central atom is sp hybridized and linked directly to two other central atoms possesses linear geometry. For example, Be in BeCl₂ molecule has an sp hybridization.

Hello, consider BeCl₂: The ground state electronic configuration of Be is 1s²2s². In the exited state one of the 2s-electrons is promoted to vacant 2p orbital to account for its bivalency **Figure 2.13**. One 2s and one 2p-orbital gets hybridized to form two sp hybridized orbitals as shown in **Figure 2.14**.

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Figure 2.14: The formation of two equivalent sp hybride orbitals in Be

The two sp hybrid orbitals face opposite directions, forming an angle of 180°.

Each of the sp hybridized orbitals overlaps with the 2p-orbital of chlorine axially and forms two BeCl sigma bonds. This is shown in **Figure 2.15.**



Figure 2. 15: The sp hybrid orbitals in gaseous BeCl, hybrid orbitals

sp² hybridization

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Hello, note that in sp² hybridization one s and two p-orbitals are mixed in order to form three equivalent sp² hybridized orbitals. For example, in BCI_3 molecule, the ground state electronic configuration of central boron atom is $1s^22s^22p^1$. In the excited state, one of the 2s electrons is promoted to vacant 2p orbital (Figure 2.16) as a result boron has three unpaired electrons. These three orbitals (one 2s and two 2p) hybridize to form three sp² hybrid orbitals as shown in Figure 2.16.



Figure 2.16: Excitation of boron 2s electrons and sp² hybrid orbitals formation

The three equivalent hybrid orbitals formed in a trigonal planar arrangement and overlap with 2p orbitals of chlorine to form three B-Cl bonds. Therefore, in BCl₃ (Figure 2.17), the geometry is trigonal planar with ClBCl bond angle of 120°.



Figure 2.17: sp² hybrid orbitals and bonding in BCl₃

sp³ hybridization

Dear learner, sp³ type of hybridization can be explained by taking the example of CH₄ molecule in which there is mixing of one s-orbital and three *p*-orbitals (Figure 2.18) of the valence shell to form four sp³ hybrid orbital of equivalent energies and shape.

Can you determine the percent of s-character and p-character in each sp3 hybrid orbital?

Hello, note that the four sp³ hybrid orbitals so formed are directed towards the four corners of the tetrahedron. The angle between sp³ hybrid orbital is 109.5° as shown in **Figure 2.19**.



Figure 2.18: Excitation of carbon 2s electrons and sp³ hybrid orbitals formation



Section 2.19: sp³ hybrid orbitals and bonding in methane

Dear learner, have you recognized that the structure of NH_3 and H_2O molecules can also be explained with the help of sp^3 hybridization. In NH_3 , the valence shell (outer) electronic configuration of nitrogen in the ground state is $2s^22p_x^{-1}2p_y^{-1}2p_z^{-1}$. In the formation of NH_3 one 2s orbital and three 2p orbitals of nitrogen are mixed up by forming four hybrid orbitals (sp^3) of equivalent energy. One of the sp^3 hybrid orbitals has a lone pair of electrons and the other three sp^3 orbitals are half-filled. These three half-filled hybrid orbitals overlap with 1s orbitals of hydrogen atoms to form three N–H sigma (σ) bonds.

The force of repulsion between a lone pair and a bond pair is more than the force of repulsion between two bond pairs of electrons.

So, the molecule gets distorted and the bond angle is reduced to 107° from 109.5°. The geometry of such a molecule will be pyramidal as shown in **Figure 2.20**.



Figure 2.20: sp^3 hybrid orbitals and bonding in NH₃

Hello, in the case of H₂O molecule, the four oxygen orbitals (one 2s and three 2p) undergo sp³ hybridization forming four sp³ hybrid orbitals out of which two contain one electron each and the other two contain a pair of electrons. These four sp³ hybrid orbitals acquire a tetrahedral geometry, with two corners occupied by hydrogen atoms while the other two by the lone pairs. The bond angle in this case is reduced to 104.5° from 109.5° (Figure 2.21) and the molecule thus acquires a V-shape or angular geometry.



 H_2O Figure 2.21: sp³ hybrid orbitals and bonding in H_2O

Why is the HOH bond angle in water less than the HNH bond angle in NH_3 ?

Hybridization of elements involving d orbitals

Dear learner, the elements present in the third period contain d orbitals in addition to s and p orbitals. The energy of the 3d orbitals is comparable to the energy of the 3s and 3p orbitals. The energy of 3d orbitals is also comparable to those of 4s and 4p orbitals. As a consequence, the hybridization involving either 3s, 3p and 3d or 3d, 4s and 4p is possible.

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Do you expect hybrid orbitals formation by mixing 3p, 3d, and 4s orbitals?

i. Formation of PCl₅ (sp³d hybridization):

The ground state and the excited state outer electronic configurations of phosphorus (Z=15) are represented below in **Figure 2.22a**.



Figure 2.22(a): sp³d hybrid orbitals filled by electron pairs donated by five CI atoms

Now the five orbitals (i.e., one s, three p and one d orbitals) are available for hybridization to yield a set of five sp³d hybrid orbitals which are directed towards the five corners of a **trigonal bipyramidal** as in the **Figure 2.2b**.



Figure 2.22(b): Trigonal bipyramidal geometry of PCl₅ molecule

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[′] All the bond angles in trigonal bipyramidal geometry are not equivalent.

Three P–Cl bond lie in one plane and make an angle of 120° with each other; these bonds are termed as **equatorial bonds**. The remaining two P–Cl bonds– one lying above and the other lying below the equatorial plane, make an angle of 90° with the plane. These bonds are called **axial bonds**.

ξ) Activity 2.12

Dear learner, why do you think that the axial bonds are slightly longer and weaker than the equatorial bonds?

ii. Formation of SF₆ (sp³d² hybridization):

Dear learner, in SF₆ the central sulphur atom has the ground state outer electronic configuration $3s^23p^4$. In the exited state the available six orbitals i.e., one s, three p and two d orbitals are singly occupied by electrons (Figure 2.23a). These orbitals hybridize to form six new sp³d² hybrid orbitals, which are projected towards the six corners of a *regular octahedron* in SF₆. These six sp³d² hybrid orbitals overlap with singly occupied orbitals of fluorine atoms to form six S–F sigma bonds. Thus, SF₆ molecule has a regular octahedral geometry as shown in Figure 2.23b.



Figure 2.23a: sp³d² hybridization



Figure 2.23b: Octahedral geometry of SF, molecule

One hybrid orbital is produced for every single atomic orbital involved.

Dear learner, you are describing the probable hybridization scheme for a structure, follow the following four steps:

- 1. Write a reasonable Lewis structure for the chemical.
- 2. Use VSEPR theory to predict the electron-set arrangement of the central atom.
- 3. Select the hybridization scheme that corresponds to the VSEPR prediction.
- 4. Describe the orbital overlap.

Example 2.4

Describe a hybridization scheme for the central atom of iodine pentafluoride, IF₅.

Solution:

The reasonable Lewis structure for IF₅ is



VSEPR predicts an octahedral electron-set arrangement for six electron pairs (AB $_{\rm 5}$ E type)

The hybridization scheme corresponding to **octahedral** electron arrangement is *sp*³d²

The six sp^3d^2 hybrid orbitals are directed to the corners of an octahedron, but one of the orbitals is occupied by a lone pair of electrons. The resulting molecular shape is that of a **square pyramid** with bond angles of approximately 90°.
🛠) Self-test Exercise 2.12

- List the three types of hybrid orbitals that may be formed by an atom with only s and p orbitals in its valence shell. Draw the shapes of the hybrid orbitals produced.
- 2. Describe a hybridization scheme for the central atom and molecular geometry of the triiodide ion, I₃⁻.

Multiple bonds and hybridization



Can the hybridization theory describe the geometry of molecules containing double and triple bonds?

Multiple bonds consist of σ and π bonds.

Dear learner, consider the **ethene**, C_2H_4 , and **acetylene**, C_2H_2 molecules. The Lewis structure of ethene, C_2H_4 , shows us that each carbon atom is surrounded by one other carbon atom and two hydrogen atoms.



The three bonding regions form a *trigonal planar* electron-pair geometry. Thus, we expect the σ bonds from each carbon atom are formed using a set of sp² hybrid orbitals that result from hybridization of two of the 2p orbitals and the 2s orbital (Figure 2.24). These orbitals form the C–H single bonds and the σ bond in the C = C double bond (Figure 2.25). The π bond in the C= C double bond (Figure 2.25). The π bond in the C= C double bond results from the overlap of the third (remaining) 2p orbital on each carbon atom that is not involved in hybridization. This unhybridized p orbital (lobes shown in red and blue in Figure 2.27 is perpendicular to the plane of the sp² hybrid orbitals. Thus, the hybridized 2p orbitals overlap in a side-by-side fashion, above and below the internuclear axis and form a π bond.







Figure 2.25: Bond formation in the ethene molecule, C_2H_4 : (a) five σ bonds and (b) one π bond

) Why does rotation around single (σ) bonds occur easily, but is much more difficult for multiple bonds?

Hello, note that in molecules with sp hybrid orbitals, two unhybridized p orbitals remain on the atom (Figure 2.26). We find this situation in acetylene, H-C=C-H, which is a *linear molecule*. The two linear sp hybrid orbitals of a carbon atom lie in a straight line and the two unhybridized p orbitals lie at perpendicular angles (Figure 2.26). The sp hybrid orbitals of the two carbon atoms overlap end to end to form a σ bond between the carbon atoms (Figure 2.27). The remaining sp orbitals form σ bonds with two hydrogen atoms. The two unhybridized p orbitals per carbon are positioned such that they overlap side by side and, hence, form two π bonds. The two carbon atoms of acetylene are bound together by one σ bond and two π bonds, giving a *triple bond*.







Figure 2.27 (a) The formation of σ bonds in an acetylene molecule, C₂H₂, and (b) π bonds.

Self-test Exercise 2.13

- Describe the hybridization of the central atoms N and S in molecules and the molecular geometry of:
 a. NO₂
 b. SO₂
- 2. Discuss the bonding in ammonium and sulphate ions: predict the ideal bond angles, bond length, shape of the ions, and the number of σ and π bonds.
- 3. Prepare sketches of the overlaps of the following atomic orbitals: a. s with s; b. s with p along the bond axis; c. p with p along the bond axis (head-on overlap); d. p with p perpendicular to the bond axis (side-on overlap).

2.5.2 Molecular Orbital Theory (MOT)

Dear learner, the molecular orbital theory assumes that when atoms come together, their orbitals, not only overlap, but are also simultaneously transformed into new orbitals. These new orbitals called, molecular orbitals (MOs), play the same role for molecules that atomic orbitals play for atoms.

Atomic orbitals are the allowed states for an electron moving in the field of one nucleus, whereas molecular orbitals are the allowed states for an electron moving in the field of several nuclei (atoms present in the molecules).

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Which theory (VBT or MOT) exactly explains all aspects of bonding?

Bonding and Anti-Bonding Molecular Orbitals

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In contrast to VB theory, in which one bonding orbital is formed as a result of the interaction of two atomic orbitals, in MO theory two molecular orbitals result from the combination of two atomic orbitals.

The molecular orbitals that results from the addition of the atomic orbitals that overlap is called a bonding orbital.

In a bonding molecular orbital, because the electron density is concentrated between the two nuclei, the energy of the system is lowered compared with that of isolated atomic orbitals.

A bonding molecular orbital is always lower in energy (more stable) than the component of atomic orbitals.

The molecular orbital that results from the subtraction of atomic orbitals that overlap is called an antibonding molecular orbital.

Hello, in antibonding molecular orbitals, the electron density is concentrated away from the region between the two nuclei. That is, anti-bonding molecular orbitals have a region of zero electron density (a node) between the nuclei. The net effect of having a low electron density between the nuclei is that the two nuclei repel each other.

An antibonding molecular orbital is always higher in energy (less stable) than the component of atomic orbitals.

When two atomic orbitals overlap end-to-end, they form two σ MOs. Consider the H₂ molecule, which has two H atoms and therefore two 1s AOs. The two 1s atomic orbitals combine (Figure 2.28) to produce two σ MOs, which differ in energy and location. One of the σ MOs is a bonding orbital, denoted σ_{1s} , the other is an anti-bonding orbital denoted σ^*_{1s} . The σ_{1s} MO has a lower energy than the original 1s AOs, while the σ^*_{1s} MO has a higher energy. (B

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Electron Configuration of Diatomic Molecules

Dear learner, electrons occupy molecular orbitals following the same rules that apply to the filling of atomic orbitals: the *aufbau* principle and the *Pauli Exclusion Principle* and *Hund's* rule are obeyed. The electronic structure of the molecule is obtained by feeding the appropriate number of electrons into the new molecular orbitals. For example, both electrons in H₂ will go into the lower energy orbital denoted by $(\sigma_{1s})^2$. There are no more electrons in H₂ so the σ_{1s}^* orbital remains empty in the ground state. **Figure 2.28** is the molecular orbital diagram for H₂.

A molecular-orbital diagram shows the relative energy and number of electrons in each MO, as well as the atomic orbitals from which they form.





Figure 2.28: Molecular orbital diagram and bonding in the $\rm H_{2}$ molecule

The number of molecular orbitals must equal to the number of atomic orbitals available for combination.

MO theory uses the term "bond order" to indicate whether a covalent bond is single (bond order = 1), double (bond order = 2) or triple (bond order = 3).

Bond order is defined as the number of electrons in bonding MOs minus the number in antibonding MOs, divided by two:

Thus, for H_2 , the bond order is $\frac{1}{2}(2 - 0) = 1$. A bond order which is greater than zero indicates that the molecule is stable relative to the separate atoms, whereas a bond order of zero implies no net stability.

In general, the higher the bond order, the stronger is the bond.

Self-test Exercise 2.14

- 1. Which theory (MOT or VBT) better in explaining covalent bond formation?
- 2. Draw molecular orbital energy diagrams for:

a. C₂ b. C₂⁺

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Dear learner, recognize that the order of energy of molecular orbitals has been determined mainly from spectroscopic data.

a. In simple homonuclear diatomic molecules (total electrons = 14 or less) the order is:

$$\sigma_{1s}\sigma_{1s}^*\sigma_{2s}\sigma_{2s}^*\left(\pi_{2p_y} = \pi_{2p_z}\right)\sigma_{2p_x}\left(\pi_{2p_y}^* = \pi_{2p_z}^*\right)\sigma_{2p_x}^*$$

b. For simple homonuclear diatomic molecules (total electrons greater than 14) the

 $(\pi_{2p_y} = \pi_{2p_z})$ comes after σ_{2p_x} and the order is:

$$\sigma_{1s}\sigma_{1s}^*\sigma_{2s}\sigma_{2s}\sigma_{2p_x}^*(\pi_{2p_y} = \pi_{2p_z})(\pi_{2p_y}^* = \pi_{2p_z}^*)\sigma_{2p_x}^*$$

The $2p_y$ atomic orbital give π bonding and π antibonding (π^*) MOs of the same energy as those produced from $2p_z$ orbitals. The π_{2p_y} and π_{2p_z} orbitals are said to be double degenerate, and similarly $\pi^*_{2p_y}$ and $\pi^*_{2p_z}$ are double degenerate.

Example 2.5

1. Use the molecular orbital theory to derive the electron configuration of:

a. H_2^+ b. Li_2 c. He_2^+ d. O_2

Which of these molecules or molecular ions exist?

Solution:

- a. H_2^+ molecular ion. This may be considered as a combination of H atom and a H⁺ ion, giving one electron to be accommodated in a MO. The bond order is $\frac{1}{2}(1-0) = \frac{1}{2}$, so we predict that H_2^+ does exist. The electron configuration is σ_{1s}^{-1} .
- b. Li_2 molecule. Each Li atom has two electrons in its inner shell, and one in its outer shell, making a total of six electrons arranged: $(\sigma_{1s})^2 (\sigma_{1s}^*)^2 (\sigma_{2s})^2$. We ignore the inner electrons here because, in general, only outer (valence) orbitals interact enough to form molecular orbitals. The bond order is $\frac{1}{2}(2-0) = 1$, so we predict that Li_2 does exist.
- c. He_2^+ Molecular ion. The electron configuration is $(\sigma_{1s})^2 (\sigma_{1s}^*)^1$. The filled bonding orbital gives stabilization whilst the half-filled gives destabilization. The bond order is $\frac{1}{2} (2-1) = \frac{1}{2}$, there is some stabilization, so the He_2^+ should exist, and it has been observed spectroscopically.
- d. O₂ molecule each oxygen has eight electrons, making a total of sixteen for the molecule. These are arranged as follows

$$(\sigma_{1s})^{2}(\sigma_{1s}^{*})^{2}(\sigma_{2s})^{2}(\sigma_{2s}^{*})^{2}(\sigma_{2p_{x}})^{2}\begin{cases} (\pi_{2p_{y}})^{2} \\ (\pi_{2p_{x}})^{2} \end{cases} \begin{cases} (\pi_{2p_{y}}^{*})^{1} \\ (\pi_{2p_{z}}^{*})^{1} \end{cases}$$

The antibonding $\pi_{2p_y}^*$ and $\pi_{2p_z}^*$ orbitals are singly occupied in accordance with **Hund's rule.** As in previous examples, the inner shell does not participate in bonding and the bonding and antibonding 2s orbitals cancel each other. A σ bond results from the filling of $(\sigma_{2p_x})^2$. Two half π bonds arise form π_{2p_y} and π_{2p_z} bonding and antibonding.

Therefore, $1\sigma + \frac{1}{2}\pi + \frac{1}{2}\pi = 2$ bonds are formed.

Self-test Exercise 2.15

 Use the molecular orbital theory and derive the electron configuration of the following molecules. Identify those which exist and those which do not exist.

a. He₂ b. Be₂ c. B₂

2. Use the MO theory to predict the bond order and the number of unpaired electrons in NO and CO.

Magnetic Properties

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Molecules with unpaired electrons exhibit paramagnetic properties, while those in which all the electrons are paired, there is **diamagnetism**.

Paramagnetic species is attracted by an external magnetic field, whereas diamagnetic species is not attracted by a magnetic field.

) Can you predict magnetic properties of diatomic molecules by employing VBT?

The antibonding π_{2py} and π_{2pz} orbitals for O_2 are singly occupied in accordance with Hund's rule. Unpaired electrons give rise to paramagnetsim. Since there are two unpaired electrons with parallel spins, this explains why oxygen is paramagnetic. The MO theory successfully predicts the paramagnetisim of O_2 , a fact not even thought of with the VB representation of the oxygen molecule (O = O).

Self-test Exercise 2.16

- 1. What is the bond order for CN^{-} , CN and CN^{+} ?
- 2. Which homonuclear diatomic molecules from the second period elements, besides O₂, should be paramagnetic?

CHECKLIST 2.5

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

I can...

- explain the valence bond theory and molecular orbital theory
- distinguish the Lewis model and the valence bond model
- discuss the overlapping of orbitals in covalent bond formation
- explain hybridization
- show the process of hybridization involved in some covalent molecules
- draw hybridization diagrams for the formation of sp, sp2, sp3, sp3d and sp3d2 hybrids
- suggest the kind of hybrid orbitals on the basis of the electron structure of the central atom
- predict the geometrical shapes of some simple molecules on the basis of hybridization and the nature of electron pairs
- discuss the hybridization involved in compounds containing multiple bonds
- explain bond length and bond strength
- explain molecular orbital theory
- describe molecular orbital-using atomic orbitals
- draw molecular orbital energy level diagrams for homonuclear diatomic molecules
- write the electron configuration of simple molecules using the molecular orbital model
- define bond order and determine the bond order of some simple molecules and molecule-ions
- determine the stability of a molecule or an ion using its bond order
- predict magnetic properties of molecules.

Section 2.6: Types of Crystal

Dear learner, in this section you will study types of crystalline solids and the attraction forces exist within crystalline solids.

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

- define a crystal
- name the four types of crystalline solid and give examples
- mention the types of attractive force that exist within each type of crystalline solid
- describe the properties of each type of crystalline solid
- build a model of sodium chloride crystal structure.

Solids can be crystalline or amorphous. A crystalline solid is composed of one or more crystals.

Each crystal has a well-defined, ordered structure in three dimensions. Sodium chloride (table salt) and sucrose (table sugar) are examples of crystalline substances.



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Are metals crystalline or amorphous solids?

An amorphous solid has a disordered structure; it lacks the welldefined arrangement of basic units (atoms, molecules, or ions) found in a crystal.

A glass is an amorphous solid obtained by cooling a liquid rapidly enough that its basic units are "frozen" in random positions before they can assume an ordered crystalline arrangement.

A structural unit of a crystalline solid has a characteristic repetitive pattern.

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Do you expect similar repetitive pattern in amorphous solids?

The crystalline solids can be classified as ionic, molecular, covalent network or metallic.

Ionic Crystals

Ionic crystals consist of ions held together by ionic bonds (Figure 2.29).

Hello, the structure of an ionic crystal depends on the charges on the cation, anion and on their radii. The properties of ionic solids are direct consequences of the strong inter-ionic forces, which create a high lattice energy. Ionic solids have high melting points, an indication of the strong attraction force holding the ions together.



Figure 2.29: Ball-and stick diagrams of the unit cells of (a) NaCl(s) and (b) CsCl(s)

) Can you build the model of crystal sodium chloride with the help of balls and sticks of any material available in the surrounding?

Activity 2.13

Dear learner, try to answer the following questions.

- 1. In covalent compounds such as water and ammonia, which type bonding can exist (intermolecular or intramolecular bonding)?
- 2. Can you give some examples of ionic solids we use in everyday life?

Molecular Crystals

Molecular solids are made up of discrete molecules that interact via intermolecular forces.

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Various combinations of *dipole-dipole*, *dispersion* and *hydrogen-bonding forces* are operative in molecular solids, which accounts for their wide range of physical properties. Dispersion forces are the principal forces acting in non-polar substances, so melting points generally increase with molar mass. Among polar molecules, dipole-dipole forces and where ever possible, hydrogen-bonding dominate.

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Why do molecular crystals have much lower melting points than ionic, metallic and network covalent solids?

The fundamental unit of a molecular solid is molecules or monatomic elements.

Molecular crystals are usually transparent, brittle, and break easily when stressed. They are usually non-conductors of heat and electricity and usually have low melting points.

Can you give some examples of molecular crystals we use in our everyday life?

Covalent Network Crystals

Hello, in this type of crystalline solids, separate particles is not present. Instead, strong covalent bonds link the atoms together throughout the network of covalent solid. As a consequence of the strong bonding, all these substances have extremely high melting and boiling points, but their conductivity and hardness depend on the nature of their bonding. The two common crystalline forms of elemental carbon (graphite and diamond) are examples of network covalent solids (Figure 2.30).



Figure 2.30: The structures of (a) diamond and (b) graphite

In a covalent network solid, the whole crystal is one giant molecule. These crystals are usually hard, non-conductors of heat and electricity, and have high melting points.

Metallic Crystals

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The strong metallic bonding forces hold individual atoms together in metallic solids. Bonding in metals can be explained as a network of positive ions immersed in a sea of electrons.

> The electrons in the valence shell of the metal atoms are highly delocalized. For this reason, metals are very good conductors of electricity.

Metallic bonding forces are stronger than those arising from intermolecular forces, so metallic solids have higher melting points than molecular solids. Metallic crystals are opaque with reflective surfaces. They are ductile and malleable, good conductors of heat and electricity, and they usually have high melting points.

Self-test Exercise 2.17

- Identify the type of crystalline solid (metallic, network covalent, ionic, or molecular) formed by each of the following substances:
 - a. SiO_2 b. KCl c. Cu d. CO_2 e. C (diamond)
- 2. Classify each substance in the table as either a metallic, ionic, molecular, or covalent network solid:

Substance	Appearance	Melting	Electrical	Solubility	
		Point	Conductivity	in Water	
Х	brittle, white	e 800 °C only if melted/		soluble	
			dissolved		
Y	shiny, malleable	1100 °C	High	insoluble	
Z	hard, colorless 3550 °C None		insoluble		

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

I can...

٠	explain the valence bond theory and molecular orbital theory	
•	distinguish the Lewis model and the valence bond model	
•	discuss the overlapping of orbitals in covalent bond formation	
•	explain hybridization	
•	show the process of hybridization involved in some covalent molecules	
•	draw hybridization diagrams for the formation of sp, sp2, sp3, sp3d and sp3d2 hybrids	
•	suggest the kind of hybrid orbitals on the basis of the electron structure of the central atom	
•	predict the geometrical shapes of some simple molecules on the basis of hybridization and the na-ture of electron pairs	
		151

- discuss the hybridization involved in compounds containing multiple bonds
- explain bond length and bond strength
- explain molecular orbital theory
- describe molecular orbital-using atomic orbitals
- draw molecular orbital energy level diagrams for homonuclear diatomic molecules
- write the electron configuration of simple molecules using the molecular orbital model
- define bond order and determine the bond order of some simple molecules and molecule-ions
- determine the stability of a molecule or an ion using its bond order
- predict magnetic properties of molecules.

UNIT SUMMARY

- Lewis electron-dot formulas are simple representations of the valenceshell electrons of atoms in molecules and ions.
- An ionic bond is a strong attractive force holding ions together. An ionic bond can form between two atoms by the transfer of electrons from the valence shell of one atom to the valence shell of the other.
- A covalent bond is a strong attractive force that holds two atoms together by their sharing of electrons. These bonding electrons are attracted simultaneously to both atomic nuclei, and they spend part of the time near one atom and part of the time near the other.
- In some cases of covalent bonding, one atom appears to provide both electrons in the bonding pair; the bond is known as coordinatecovalent bond or dative bond.
- Molecular geometry refers to the spatial arrangement of atoms in a molecule.
- The valence-shell electron-pair repulsion (VSEPR) model is a simple model for predicting molecular geometries.
- In the valence bond method (VB) a covalent bond is viewed as the overlap of atomic orbitals of the bonded atoms in a region between the atomic nuclei.
- In the VB theory, hybridized atomic orbitals are formed by the combination and rearrangement of orbitals of the same atom.
- Hybridized orbitals are all of equal energy and electron density, and the number of hybridized orbitals is equal to the number of pure atomic orbitals that combine.
- Molecular orbital theory describes bonding in terms of the combination and rearrangement of atomic orbitals to form orbitals that are associated with the molecule as a whole.
- Bonding molecular orbitals increase electron density between the nuclei and are lower in energy than individual atomic orbitals. Antibonding molecular orbitals have a region of zero electron density between the nuclei, and an energy level higher than that of the individual atomic orbitals.

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5	Self-assessment Questions						
Pa	Part I: Multiple Choice Questions: Choose the correct answer f	rom the given					
alt	alternatives.						
1. Which of the following is not a property of ionic bonds?							
	A. loss of electrons C. sharing of electron	ons					
	B. gain of electrons D. transfer of electro	ons					
2.	2. Which one of the following does not comply with the octet r	rule?					
(a. PCI_3 b. CBr_4 c. NF_3 d. AsF_5						
3.	3. A π (pi) bond is the result of the:						
	A. OVERIAD OF TWO'S ORDITAIS						
	B. overlap of an s and a p orbital						
	C. overlap of two p orbitals along their axes						
	D. sidewise overlap of two parallel p orbitals						
4.	4. The maximum number of hydrogen bonds that a molecule of	of water can					
	have is:						
	A. 1 b.2 c.3 d.4						
5.	5. Which of the following is planar?						
	A. NO_3^{-} b. H_3O^+ c. SO_3^{2-} d. PF_3						
6.	6. Among the following mixtures, dipole-dipole as the major int	eraction, is					
	present in:						
	A. Denzene ana emanoi						
	B. acetonitrile and acetone						
C. KCl and water							
	D. benzene and carbon tetrachloride						
7.	7. The main axis of a diatomic molecule is Z. Atomic orbitals $p_{\rm x}$	and p_y					
	overlap to form which of the following orbital?						
	A. a. π-molecular orbital C. c. δ-molecular or	bital					
	B. b. σ -molecular orbital D. d. no bond will for	orm					
8.	8. The structure of IF ₇ is:	d					
	A. peniagonal pipyramia C. Trigonal pipyrami	U					
0	 b. square pyramia b. Uctahedral c. square pyramia d. Uctahedral 	and process					
У.	 In which of the following molecules, is a co-ordinate covalent is 	oona present?					

	Α.	F ₂	b. O,		c. N	H	d. C	Cl
10.	ln v	∠ vhich of the	followina c	ompounds	are 8	3 Valence	e electrons pre	esent in
	every atom?							
	А.	IF8	$b. C_2 H_4$		c. Sil	F₄	d. KH	1
11.	Мо	lecules are l	held togeth	ner in a crys	tal by	/:		
	Α.	hydrogen b	oonds		C.	Van de	r Waal's attrac	ction
	B. electrostatic attraction D. dipole-dipole attraction						on	
12.	Find	d the molec	ule with the	e maximum	dipo	le mome	ent:	
	Α.	CH_4	b. NH	H ₃	(c. CO ₂		d. NF ₃
13.	Wh	ich of the m	olecules do	pes not hav	e a p	permane	nt dipole mom	ient?
	Α.	SO ₃	$b.SO_2$	с. Н ₂	S		d. CS ₂	
14.	The	e F-S-F bond	angles in SF	are	_•			
	Α.	09°28'	b. 120° only	c. 9	0° an	d 120°	d. 90° and 18	80°
15.	15. A neutral molecule having the general formula AB_3 has two unshared pair							
	of electrons on A. What is the hybridization of A?							
	Α.	sp	b. sp ²	c.sp ³	d. :	sp³d		
Pa	rt II:	Answer the	following c	questions				
16.	Wri	te Lewis dot	symbols for	r the followi	ng at	toms and	d ions: Mg, Na,	B, N, Br,
Be, Li ⁺ , Cl ⁻ , O ²⁻ , Mg ²⁺ and N ³⁻								
17.	Wh	ich one is sti	ronger: σ-bo	ond or π-bo	nd? V	Why?		
18.	18. Explain why a BeH, molecule has a zero dipole moment although the							
	Be-H bonds are polar?							

- 19. How can one non-polar molecule induce a dipole in a nearby non-polar molecule?
- 20. Describe the change in hybridization of an aluminum atom (if any) during the reaction:

 $A|C|_3$ + $C|^- \rightarrow A|C|_4^-$

21. Apart from tetrahedral geometry, another possible geometry for CH₄ is square planar with the four H atoms at the corners of the square and C atom at its center. Explain why CH₄ is not square planar?

Answer key for Activities of Unit Two

Activity 2.1

- Atoms combine with each other only if the process leads to lowering of energy. In the process each atom attains a stable outer octet of electrons. The electron configuration and the strength of the nucleus-electron attraction(s) determine the properties of an atom, the type and strength of chemical bonds form.
- 2. Elements who have 1, 2 and 3 valence electrons are known as metals, now for stability, they have to complete their octet for this they lose their valence electrons to complete their octet because it is easier to lose electrons than gain it. On the other hand, non metals have 4, 5, 6 or 7 valence electrons and can only gain electrons to attain octet configuration and thus they form anions.

Activity 2.2

Sodium and chlorine as well as lithium and oxygen can form ionic compounds but not nitrogen and fluorine.

a.
$$\begin{bmatrix} K \end{bmatrix}^+$$
 b. $\begin{bmatrix} \vdots S \\ \vdots S \end{bmatrix}^{2-}$ c. $\begin{bmatrix} Al \end{bmatrix}^{3+}$ d. $\begin{bmatrix} \vdots F \\ \vdots S \end{bmatrix}^{2-}$



- When an atom expands its octet, it does so by making use of empty d orbitals that are available in the valence level of the atom doing the expanding. Atoms that do not have empty valence level d orbitals cannot expand their octets. Second period atoms cannot have more than eight valence electrons around the central atom, but atoms in the third (and below) period can.
- 2. BeCl2, BF3 and AlCl3 are examples

E Activity 2.5

- 1. Hydrogen doesn't form ionic bonds because it has an average electronegativity of 2.1. Simply put, there aren't any atoms that are electronegative to completely take its electrons, nor there any atoms with low enough electronegativity can that hydrogen take their electrons.
- 2. In an ionic bond, there is complete transfer of electrons, while in a covalent bond, there is sharing of electrons. In an ionic bond, the atoms are bound by attraction of opposite ions, whereas, in a covalent bond, atoms are bound by sharing electrons. In covalent bonding, the molecular geometry around each atom is determined by VSEPR rules, whereas, in ionic solids, the geometry follows maximum packing rules.



- Because they are only attracted by one nucleus, lone pairs of electrons take up more space than bonded electrons because they are spread out at a shorter distance from that nucleus. Bonding electrons are further from the nucleus, but they are more confined therefore they aren't as dispersed.
- 2. Water has bent shape due to repulsion between two bonding pairs and two lone pairs which make bond angle of 104.5 ° from ideal tetrahedral arrangement which is 109°. Whereas ammonia has bond angle 107.5 due to repulsion between one lone pair and three bond pairs.



- 1. Dipole dipole, hydrogen bond, London (dispersion) forces, etc.
- 2. Hydrogen bonds (strongest) hold water molecules in a rigid structure keeping them very close to each other, whereas the weak London (dispersion) forces in methane cannot bring distant particles close to each other. This is one of the reasons that the particles of gases are far apart.
- 3. The hydrogen bond in ice holds the water molecules in a rigid but open structure and more volume is occupied. As ice melts, some of the hydrogen bonds are broken. Water molecules move into some of the holes in the rigid structure of ice. This means that liquid water at 0 °C is denser than ice.

Activity 2.8

- 1. a. They have high electrical and thermal conductivity.
 - b. They have bright lustrous appearance.
 - c. They are malleable and ductile.
 - d. They are hard
- 2. Electrons which are not confined to a particular atom (mobile electrons).



2. 4p of Br overlap with 3p of chlorine

🕞 Activity 2.10

- In sigma bonds, the large overlap of the orbital involves the removal of a large amount of energy. While in pi bonds the extent of overlapping is less than sigma bond. Therefore, sigma bond is stronger than pi bond.
- 2. A double bond between two atoms is stronger and shorter than a single bond between the same two atoms. A triple bond is even stronger/shorter.



Activity 2.12

This is because of the equatorial bonds lie in one plane and make an angle of 120° with each other; whereas, the axial bonds-one lying above and the other lying below the equatorial plane, make an angle of 90° with the plane. As the axial bond pairs suffer more repulsive interaction from the equatorial bond pairs, axial bonds have been found to be slightly longer and hence slightly weaker than the equatorial bonds.



- 1. Both intramolecular and intermolecular bonding.
- 2. Table salt (sodium chloride), baking soda (sodium bicarbonate) are examples of ionic solids.

ਾਰ) Answer key for self-test Exercises

Self-test Exercise 2.1

- a. When the two atoms are in contact, potassium readily transfers its outer electron to chlorine which readily accepts it, resulting in both atoms achieving a state of eight outermost electrons. With this electron transfer, the ionic bond in KCl is formed.
 - b. When magnesium reacts with oxygen to form magnesium oxide, the magnesium atom loses two electrons to form the Mg²⁺ cation, which has electronic configuration 2, 8. Oxygen atom gain the two electrons to form the O²⁻ anion with configuration 2, 8, 8. Then, Mg²⁺ cation interact with O²⁻ anion to form ionic bond.
 - c. Two sodium atoms losing two electrons; one oxygen atom gaining the two lost electrons to form ionic bond.
- 2. Only calcium because chlorine is a non-metal, and calcium is a metal. An ionic bond is between a metal and a non-metal. Carbon, oxygen and silicon are non-metals.
- 3. The net attractive force that holds positive and negative ions together in a crystalline solid
 - a. Ca²⁺ and Cl⁻ b. Mg²⁺ and O²⁻ c. Al³⁺ and O²⁻





- 1. The ions in MgO contains stronger ionic bond than LiF because the charges on Mg and O are twice that of Li and F of similar size.
- 2. The ionic attractions sodium chloride is stronger than that of potassium chloride because the size of sodium is smaller than that of potassium.
- 3. Lattice energy increases as the charge on the ion increases and the size decreases.

Self-test Exercise 2.5

- 1. The heavier post-transition elements like Pb and Sn lose the p electrons or both the p and s electrons from the valence shell.
- 2. ₀Sn: [Kr] 5s²4d¹⁰5p²
 - ₅₀Sn²⁺: [Kr] 5s²4d¹⁰

₅₀Sn⁴⁺: [Kr] 4d¹⁰

Neither of these configurations are noble gas configurations.

- The ions of some transition elements and post-transition elements do not usually obey the octet rule. Generally, these properties are exhibited by ions of elements from
 - i. Group IB and Group IIB (transition elements) and
 - ii. Group IIIA and Group IVA (heavier post-transition elements)
- 4. The Se atom has to gain two electrons in order have a complete octet in its valence shell.
- 5. a. both Li and F in LiF obey the octet rule
 - b. both Na and Cl in NaCl obey the octet rule
 - c. O in CuO obeys the octet rule but Cu does not
 - d. Cl in FeCl₃ obeys the octet rule but not Fe

Self-test Exercise 2.6

1. a.
$$CO_2$$
 has (1 × 4) + (2 × 6) = 16

b. SO_4^{2-} has $(1 \times 6) + (4 \times 6) + 2 = 32$

c. NH_4^+ has $(1 \times 5) + (4 \times 1) - 1 = 8$ d. N_2O_4 has $(2 \times 5) + (4 \times 6) = 34$

4. H and F because they cannot form more than one covalent bond.



Self-test Exercise 2.9

- From their electronegativity data, one can predict that there is a bond moment in both CO₂ and BCl₃. But from the molecular geometry of these molecules (CO₂ linear and BCl₃ trigonal planar), it can be predicted that the bond moments cancel to give a non-polar molecule to produce a zero resultant dipole moment.
- 2. The dipoles in decreasing order is HF > HI > HBr > HCI.



Self-test Exercise 2.10

- a. trigonal planar, 120°
 b. trigonal pyramidal, 109.5°
 c. bent, 98.3°
 d. Tetrahedral, 109.5°
 e. Linear, 180°
- 2. When there is no lone pair of electron(s) on the central atom.
- 3. $OF_2 < NF_3 < CF_4 < BF_3 < BeF_2$
- 4. $PCl_{_5}(g)$ (AX₅ type) has a trigonal bipyramidal molecular shape. In the solid phase, $PCl_{_4}^+$ (AX₄ type) is tetrahedral, and $PCl_{_6}^-$ (AX₆ type) is seesaw in shape.
- 5. By the vector sum of bond moments.

Self-test Exercise 2.11

- 1. a. In b and c because hydrogen is bonded to highly electronegative atoms oxygen and nitrogen.
- 2. a. Dipole-dipole forces and London forces.
 - b. London forces, dipole-dipole forces, and hydrogen bonding.
 - c. London forces.
 - d. London forces and dipole-dipole forces.
- a. ethanol (hydrogen bonding)b. hexane (Dispersion forces)c. magnesium chloride (ionic)d. methyl amine (hydrogen bonding)
- 4. Intermolecular forces act between molecules or between molecules and ions. Whereas, intramolecular forces are chemical bonds that hold atoms together in a molecule. Generally, intermolecular forces are much weaker than bonding forces.



1. Molecular orbital theory is better than valence bond theory in terms of correctly explain the bonding.







17. In a sigma bond, the electrons are in orbitals between the nuclei of the bonding atoms (in the plane of electron density), where electron density is greatest. The end-to-end axial overlap of electrons attracts the nuclei

to form a strong bond. In a pi bond, the p-orbitals overlap side-to-side above and below the nuclei (equatorial plane), so this attraction is more of a "sideways" pull as opposed to a direct line of pulling. Additionally, the election density is spread out over a greater area. Thus, sigma bonds are stronger than pi bonds.

- 18. Although the Be-H bonds are polar having the same dipole moment, but due to the linear structure of BeH₂, the bond dipoles of the two Be-H bonds cancel each other. Therefore, the resultant dipole moment of BeH₂ molecule is zero.
- 19.It arises from temporary variations in electron density in atoms and molecules. At any instant, the electron distribution may be unsymmetrical and hence produce an instantaneous dipole. This can cause an induced transient dipole in the neighboring molecule and cause the molecules to be attracted.
- 20. The hybridization of aluminum atom in $AICl_3$ is sp^2 whereas in $AICl_4^-$, the hybridization of aluminum is sp^3 .
- 21. According to VESPR theory, if CH₄ were square planar, the bond angle would be 90°. For tetrahedral structure, the bond angle is 109.5°. Therefore, in square planar structure, repulsion between bond pairs would be more and thus the stability will be less.



UNITS

PHYSICAL STATES OF MATTER

Unit Overview

Dear learner, in this unit you will learn about the three physical states of matter gases, liquids, and solids. It introduces the kinetic molecular theory of matter and properties of matter. Different gas laws that govern the properties of gaseous molecules and the effects of variables such as temperature, pressure, volume, and number of moles are presented in detail. The properties of liquids such as changes in the physical properties of liquids by boiling, condensation, evaporation, and vapor pressure are also included. Finally, the properties of the solid state are also discussed. Concepts such as melting point, freezing point, sublimation, determination of melting point, and phase changes are presented in this unit.

Unit Outcomes

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

- discuss the kinetic molecular theory and properties of the three physical states of matter
- explain the behaviour of gases by using the variables volume, temperature, pressure and number of moles
- analyse the gas laws
- develop skills in solving problems to which the gas laws apply
- perform activities to illustrate gas laws

- distinguish terms like ideal gas, diffusion, evaporation, boiling, condensation, vapour pressure, boiling point, molar heat of vaporization, molar heat of condensation, melting, fusion, sublimation, melting point, freezing point, molar heat of fusion, and molar heat of solidification
- carryout experiments to determine the boiling points of liquids and the melting point of solids
- demonstrate an experiment to show phase changes.
- describe scientific enquiry skills along this unit: inferring, predicting, classifying, comparing and contrasting, communicating, asking questions and making generalization.

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The Required Study Time

5 weeks

If you set an average of one hour per day, you will be able to complete this unit in four weeks. Try your best to answer self-assessment questions, the experimental activities, and the unit review questions. Cross check your observations for the experimental activities and answers for the self-assessment and unit review questions with the suggested answers at the end of the Module.

If you do not get a right answer to a particular self-assessment question in your first attempt, you should not get discouraged, but instead, attempt to answer it over and over again. If you still do not get it right after several attempts, then you should rather seek help from your friend or even your tutor.

Unit Learning Strategies

- You should monitor your performance through planning, monitoring, and self-regulation
- You should manage the learning environment and available resources
- You should try all self-test, self-assessments and written assignments
- You should use additional reference materials for better understanding contents in the module.
- You should perform experimental activities to understand better about the lesson given in the module
- You should consult your tutor for further clarifications on some contents in the module.

Section 3.1: Introduction



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

- describe the physical states of matter
- give examples for each of the three physical states of matter
- compare and contrast the three physical states of matter

Activity 3.1

Dear learner, attempt to answer the following questions based on daily life experinance.

- Sometimes during rainy season you observe ice rain. What is the source of the ice rain (precipitation)? What makes it different from water?
- 2. If you take some of amount of water in evaporating dish and gently heat for some time, what kind of change do you observe?



What is matter? Can you mention the three common states of matter?

Dear learner, from your previous lessons you do remember that all objects around us are called matter. Matter is defined as anything occupying space and having mass, which is the material of the Universe. It exists in three common states: gas, liquid and solid. For example, water can exist in the three physical states of matter as follows:

- Steam (water vapor): gaseous state of water
- Water: liquid state of water
- Ice: solid state of water



Dear learner, most solids change to liquids and most liquids change to gases as they are heated. Liquids and gases are known as fluids because they flow freely. Solids and liquids are referred to as condensed states because they have much higher densities than gases.

In our everyday life, we experience changes in the state of matter. For example, ice melts and water freezes; water boils and steam condenses.

Note that the states of matter can be changed to solids, liquids, gases, or plasma by varying temperature and/or pressure.

Solid

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What are the main characteristics of solids?

Dear learner, a solid is rigid; it has a fixed volume and a fixed shape. Solids cannot be compressed. They have higher densities than liquids because their particles are usually packed closer than those in liquids. The particles of solids are also highly organized and stay in a fixed position.

Metals, chalk, sand, soil, and stone are examples of solids.

Liquid

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What is liquid and its characteristics?

A liquid has a definite volume, but has no a definite shape. Liquid takes the shape of its container. It is very slightly compressible and has the ability to flow.
Water, ethanol, mercury, bromine, and oil are examples of liquids.

Gas

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What are the distinguishing properties of gases?

A gas has neither definite volume nor definite shape; it takes on the shape and volume of its container. Gas is highly compressible because most of the volume of a gas is composed of the large amounts of empty space between the gas particles.

(B)

Air, hydrogen, oxygen, carbon dioxide, and nitrogen are gases.

Plasma

What distinguishes plasma from other states of matter?

Dear learner, plasma is the fourth physical state of matter that exists at a high temperature (million degrees Celsius). This state of matter, a gaseous mixture of positive ions and electrons, is called plasma. In another word, plasma is an ionized gas. A particle diagram of a solid, liquid, gas and plasma is given in **Figure 3.1**.







- 2. How is placed and command the intervence of the
- 3. How is plasma different from gas?

🕖 CHECKLIST 3.1

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

l can ...

- describe the physical states of matter
- give examples for each of the three physical states of matter
- compare and contrast the three physical states of matter

Section 3.2: Kinetic Theory and Properties of Matter

Hello, in this section you will learn about the kinetic theory of matter and how the properties of matter are expressed in terms of the kinetic theory.

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

- state the kinetic theory of matter
- explain the properties of matter in terms of the kinetic theory

The Kinetic Theory of Matter

) What does the kinetic theory describe?

Dear learner, according to the kinetic theory of matter, every substance consists of a very large number of very small particles called ions, atoms and molecules. These particles are in a state of continuous and random motion with all possible velocities. The motion of the particles increases with a rise in temperature.



What are the assumptions of the kinetic molecular theory of matter?

Generally, the kinetic theory of matter is based on the following assumptions:

- 1. All matter is composed of particles that are constantly moving.
- 2. All particles have kinetic energy and potential energy. Molecules in the solid phase have the least amount of energy, while gas particles have the greatest amount of energy.
- 3. The difference between the three states of matter is the energy contents they have and the motion of the particles. A change in phase may occur

when the energy of the particles is changed.

- 4. The temperature of a substance is a measure of the average kinetic energy of the particles.
- There are spaces between particles of matter. The average amount of empty space between molecules gets progressively larger as a sample of matter moves from the solid to the liquid and gas phases.
- 6. There are attractive forces between atoms/molecules, and these become stronger as the particles move closer together. These attractive forces are called intermolecular forces.

Activity 3.2

Dear learner, consider the following substances; one existing in the solid state, Cu; the second in the liquid state, H2O; and the third in the gaseous state, O2 at room temperature.

1. Use the following table to explain the motion, distance and attraction between particles.

Substances	Motion of particles	Distance between particles	Attraction between particles
Cu			
H ₂ O			
O ₂			

2. In which state of matter do the particles possess the highest kinetic energy?

Properties of Matter

Dear learner, the properties of these three states of matter: a solid, liquid, and gas, are explained in terms of the kinetic theory as follows:

Properties of Gases



Can you state the properties of gases according to the kinetic theory?

From the kinetic molecular theory, the following general properties of gases can be summarized.

- 1. Gases have no fixed shape and fixed volume. They assume the volume and shape of their containers.
- 2. Gases can be easily compressed. This happens due to the large spaces between the particles of gases.
- 3. Gases have low densities compared to liquids and solids. A small mass of a gas occupies a large volume and hence, results in a very low density.
- 4. Gases exert pressure in all directions. Gases that are confined in a container exert pressure on the walls of their container.
- 5. Gases easily flow and diffuse through one another. A gas moves freely and randomly throughout in a given space.

Properties of Liquids

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Liquids have the following properties:

- Liquids have a definite volume, but have no definite shape. They assume the shapes of their containers. Lack of a definite shape for liquid substances arises from low intermolecular forces of attraction between their particles as compared to that of solids.
- Liquids have higher densities than gases. Their density is a result of the close arrangement of liquid particles. Thus, the particles are closer in liquid than in their gaseous state. This accounts for the higher densities of liquids as compared to gases.
- 3. Liquids are slightly compressible. The available free space between particles in liquids is very little as a result, liquids resist an applied external force.
- 4. Liquids are fluids. Most liquids naturally flow downwards due to gravitational force. But, the fluidity of liquids is much slower than gases.

A fluid is a substance that can easily flow.

Properties of Solids



What are the properties of solid?

Solids have the following properties

- Solids have a definite shape and definite volume. These are happened due to the presence of the strong force of attraction between the particles in a solid.
- 2. Solids generally have higher densities than gases and liquids. The particles of solids are very close to each other. This closeness of particles makes solids to have more particles (mass) per unit volume.
- 3. Solids are extremely difficult to compress.
- 4. Solids are not fluid because particles of solids are rigidly held in position by strong forces that restrict them to move.



Arrange the three states of matter in terms of their increasing:

- a. intermolecular force
- b. density
- c. compressibility
- d. kinetic energy

) CHECKLIST 3.2

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

l can ...

- an state the kinetic theory of matter
- explain the properties of matter in terms of the kinetic theory



Section 3.3: The Gaseous State

Dear learner, in this section you will learn about assumptions of kinetic molecular theory of gases; the properties of gases according to the kinetic molecular theory; how variables such as V (volume), T (Temperature), P (pressure) and n (number of moles) affect the behaviour of gases; different gas laws and their application.

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

- explain the assumptions of kinetic molecular theory of gases
- describe the properties of gases using kinetic molecular theory
- describe the behavior of gases by using the variables V (volume), T (Temperature), P (pressure) and n (number of moles)
- state Boyle's law
- perform an activity to show changes in volume and pressure of gases to illustrates Boyle's law
- apply Boyle's law in solving problems
- state Charles' law
- conduct an activity to show changes in volume and temperature of gases to illustrate Charles' law
- apply Charles' law in solving problems
- derive combined gas law equation from Boyle's law and Charles' law
- use the combined gas law to calculate changes in volume, pressure or temperature
- define an ideal gas
- derive an ideal gas equation from Boyle's law, Charles' law and Avogadro's law
- compare the nature of real gases with ideal gases
- solve problems related to ideal gas equation
- define diffusion
- state Graham's law of diffusion
- carry out an activity to compare the rate of diffusion of two different gases
- apply Graham's law of diffusion in solving problems.

The Kinetic Molecular Theory of Gases

Dear learner, the Kinetic molecular theory is a simple model for understanding the behavior of gases. This model predicts the correct behavior of most gases. Like other models, the kinetic molecular theory is not perfect.



What are the assumptions of the kinetic molecular theory of gases?

Assumptions of the kinetic molecular theory of gases:

- 1. A gas is a collection of particles (molecules) in constant, straight-line motion.
- 2. Gas particles do not attract or repel each other, they do not interact. The particles collide with each other and with the surfaces around them, but they bounce back from these collisions like pool balls.
- 3. There is a lot of space between gas particles compared with the size of the particles themselves. That means, the volume of the particles is negligible compared to the total volume of the gas. This assumption explains why gases are so easily compressed and why they can mix so readily.
- 4. The average kinetic energy of gas particles is proportional to the temperature of the gas in kelvins. This means, as the temperature increases, the particles move faster and therefore have more energy. The mathematically expressed as:

$$KE = \frac{1}{2}mv^2$$

Where KE is kinetic energy, m is mass and v is velocity of gas molecules. According to this formula different molecules have different velocities and thus, have different kinetic energies.

Note that the gaseous state has generally the following physical properties.

(P

- Gases are highly compressible.
- Gases exert pressure equally in all directions.
- Gases have much lower density than the solids and liquids.
- The volume and the shape of gases are not fixed. These assume volume and shape of the container.
- Gases mix evenly and completely.

Activity 3.3

Dear learner, consider three water samples which are taken in containers: A, B and C. The three water samples are kept at different temperature: 120 °C, 20 °C, and -10 °C, respectively and constant atmospheric pressure.

- a. Which sample molecules has the highest kinetic energy?
- b. Which sample has the greatest density?
- c. Which sample has the most regular arrangement of molecules?

Section 3.4: The Gas Laws

Dear learner, the gas laws are the results of various research conducted for several centuries on the physical properties of gases. The first reliable measurement on the properties of gases was made by Robert Boyle in 1662 (an Anglo-Irish scientist). The observation of Boyle's and other scientists led to the development of the gas laws.



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Why were gas laws discovered?

The gas laws express mathematical relationships between the volume, temperature, pressure, and quantity of a gas.

) How do you measure pressure, volume and temperature of gaseous molecules?

Pressure: the pressure exerted by a gas sample is defined as the force per unit area that results from the collisions of gas particles with surrounding surfaces.

Pressure =
$$\frac{Force}{Area}$$

Pressure is one of the measurable properties of gases. Thus, it can be expressed in unit of atmosphere, Pascal, torr, millimeter of mercury. The SI unit of pressure is Pascal (Pa).

 $1 Pa = 1 N/m^2$ and

1 atm = 760 mmHg = 76 cmHg = 760 torr = 101325 Pa = 101.325 kPa

Volume: volume is the space occupied by a substance. The SI unit of volume is the cubic meter (m³). Volume is also expressed in cubic centimeter (cm³) and cubic decimeter (dm³). Other common units of volume are milliliter (mL) and liter (L).

$$1 \text{ cm}^3 = (1 \times 10^{-2} \text{ m})^3 = 1 \times 10^{-6} \text{ m}^3$$

$$1 \text{ dm}^3 = (1 \times 10^{-1} \text{ m})^3 = 1 \times 10^{-3} \text{ m}^3 = 1 \text{ L}$$

A liter is equivalent to one cubic decimeter, as follows

 $1 L = 1000 mL = 1000 cm^3 = 1 dm^3$

Temperature: temperature is the degree of hotness or coldness of a body. Three temperature scales are commonly used. These are °F (degrees Fahrenheit), °C (degrees Celsius) and K (Kelvin). In calculations, the Kelvin scale is usually used. The three units are interconverted as follows:

$$K = °C + 273$$

°C = (°F - 32) 5/9

Self-test Exercise 3.3

Convert the following:

- a. 500 mmHg into atm, torr, and cmHg
- b. 100 dm³ into mL, cm³, L, m³
- c. 54 °C into K and °F.

How do you determine molar volume and standard conditions?

The conditions of a pressure of 1 atmosphere and a temperature of 0 °C (273 K) are called standard temperature and pressure (STP) for gases. At STP the volume of one mole of any gas is 22.4 liters. This volume is known as the molar volume of gases.

i. Boyle's Law

How do you mathematically express Boyle's law?

Hello, the first quantitative experiment on gases was performed by Robert Boyle (1662). His experiment helped to understand the relationship between the volume and pressure of a fixed amount of a gas at constant temperature. On the basis of his experiments, at constant temperature, the pressure of a fixed amount of gas varies inversely with its volume. This is shown in **Figure 3.2**.



Figure 3.2: The relation between pressure and volume

) How can you graphically show the relationships between volumepressure at constant temperature?

Values of volume-pressure relationships are given in Table 3.1.

Table 3.1: Pressure	and	volume	data	for	a	gas	at	constant	mass	and	tem-
perature.											

Pressure (atm)	Volume (mL)	PV
0.5	1200	600
1.0	600	600
2.0	300	600
3.0	200	600
4.0	150	600
5.0	120	600
6.0	100	600

Figure 3.3 (a and b) shows the graphical representation of the relationship between p and v, using the data of Table 3.1.

(B

Figure 3.3: Graphical representations of Boyle's Law, using Table 3.1 data. (a) V versus P and (b) V versus 1/P

Boyle's law states that at a constant temperature the volume of a fixed mass of gas is inversely proportional to the pressure.

Mathematically expressed as:

$$V \alpha \frac{1}{P}$$

 $V = k \left(\frac{1}{P} \right)$ or $PV = k$

Where V is volume, P is pressure and k proportionality constant.

This relationship is true because if the pressure increases, the volume decreases, but the product PV is always equal to the same constant, k.

Thus, for two different sets of conditions, we can say:

$$P_1V_1 = k = P_2V_2$$

 $P_1V_1 = P_2V_2$

Where P_1 and V_1 are the initial pressure and volume of the gas, and P_2 and V_2 are the final volume and pressure.

Dear learner, to understand the relationship between pressure and volume you need to perform the following experiment. To conduct the experiment, visit a high school in your vicinity and request Grade 11 chemistry teacher to support you in performing or demonstrating the experiment. Note that the experiment should be performed only in the laboratory, or do not attempt to perform at your home. In case, if high school is not available in your vicinity, request your tutor to get support and advices.

Experiment 3.1

The effect of pressure on the volume of gas

Objective: To observe the relationships between the volume and pressure of a gas at constant temperature.

Apparatus: U-tubes, ruler, rubber tube, burette, glass tube.

Chemicals: Mercury.

Procedure:

- Join two tubes by a rubber tubing to give a U-arrangement as shown in Figure 3.4 and then partially fill these two tubes with mercury.
- 2. Put a ruler in the middle of the tube.
- 3. The first arm of the tube (A) contains air and is sealed by a tap.
- 4. By moving the second arm of the tube (B) up and down, the volume of air in the first tube can be varied.
- 5. The pressure exerted on the air is obtained from the difference in height of mercury in the two arms of the tube.



Figure 3.4: The effect of pressure on the volume of a gas at constant temperature

Observations and analysis:

- 1. Plot a graph taking pressure on the vertical axis versus volume on the horizontal axis and comment on the shape of the graph.
- 2. What can you conclude from this experiment?

Example 3.1

Sulfur dioxide (SO₂), that forms acid rain, is found in the exhaust of petrol and diesel vehicles, and power plants. Consider a 1.5 L sample of gaseous SO_2 at a pressure of 5.6 kPa. If the pressure is changed to 15.0 kPa at a constant temperature, what will be the new volume of the gas?

Solution:

We are given: Initial conditions: $P_1 = 5.6$ kPa atm and $V_1 = 1.5$ L

Final conditions: $P_2 = 150.0$ kPa and $V_2 = ?$

Use Boyles' law equation: $P_1V_1 = P_2V_2$

Thus,
$$V_2 = \frac{P_1 V_1}{P_2}$$

= $\frac{5.6 \text{ kPa} \times 1.5 \text{ L}}{150 \text{ kPa}} = 0.56 \text{ L}$

Self-test Exercise 3.4

- 1. A certain gas occupies a volume of 10.0 m³ at a pressure of 100.0 kPa. If its volume is increased to 20 m³, what would be the new pressure of the gas assuming temperature remains constant?
- A cylinder equipped with a moveable piston has an applied pressure of
 4.0 atm and a volume of 6.0 L. What is the volume of the cylinder if the
 applied pressure is decreased to 1.0 atm?

ii. Charles' Law

Dear learner, following Boyle's findings, other scientists continued to study the properties of gases. One of these scientists was a Jacques Charles (a French physicist), the first person who flight a balloon filled with a hydrogen gas.



Why does the plot of the volume of a gas versus its temperature (K) at constant pressure give straight line?

In his experiment, Charles (1787), observed that at constant pressure the volume of a gas increases linearly with its temperature. This means, a plot of the volume of a gas versus its temperature (K) gives a straight line. This behavior is shown for samples of several gases in **Figure 3.5**.



Figure 3.5: Plots of V versus T (K) for several gases

Why do the slopes of the lines, in Figure 3.5, different?

The slopes of the lines in this graph are different because the samples contain different numbers of moles of gases. However, the volumes of all the gases extrapolate to zero at the same temperature, -273 °C. On the Kelvin temperature scale this point is defined as 0 K, which leads to the following relationship between the Kelvin and Celsius scales:

At temperatures below 0 K, the extrapolated volumes of gases would become negative. However, in reality a gas cannot have a negative volume, this implies that 0 K has a special significance.

0 K (-273 °C) is considered as absolute zero, which is theoretically the lowest attainable temperature, but practically cannot be attainable.

Can you state the Charles' law?

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Charles' law states that the volume of a fixed mass of gas at constant pressure varies directly with the Kelvin temperature.

$$V \alpha$$
 1

k = VT

The value of T is the Kelvin temperature, V is volume and k is a constant.



The ratio V to T for any set of volume-temperature values always equals the same k.

Thus, for two different sets of conditions, we can say:

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

Example 3.2

1. A sample of gas at 15 °C and 1 atm has a volume of 2.58 L. What volume will this gas occupy at 38 °C and 1 atm?

Solution:

We are given: Initial conditions: $T_1 = 15 \text{ °C} = 15 + 273 = 288 \text{ K}$, $V_1 = 2.58 \text{ L}$

Final Conditions: $T_2 = 38 \text{ °C} = 311 \text{ K}, V_2 = ?$

By rearranging Charles' equation V_2 can be calculated as

$$V_{2} = \frac{V_{1}T_{2}}{T_{1}}$$
$$= \frac{2.58 \text{ L} \times 311 \text{ K}}{288 \text{ K}} = 2.78 \text{ L}$$

2. A sample of gas has a volume of 2.80 L at an unknown temperature. When the sample is submerged in ice water at 0 °C, its volume decreases to 2.57 L. What was its initial temperature (in kelvins and in Celsius)?

Solution:

We are given: Initial conditions: $V_1 = 2.80$ L, T_1 ?

inal conditions: V $_{\rm 2}$ = 2.57 L, T $_{\rm 2}$ = 0 °C = 273 K

$$T_1 = \frac{T_2 V_1}{V_2} = \frac{273 \text{ K} \times 2.80 \text{ L}}{2.57 \text{ L}} = 297.43 \text{ K}$$

In °C, 297.43 - 273 = 24.43 °C

Dear learner, to get acquainted with the volume-temperature relationship of a gas at constant pressure you need to perform the following experiment. To conduct the experiment, visit a high school in your vicinity and request Grade 11 chemistry teacher to support you in performing or demonstrating the experiment. Note that the experiment should be performed only in the laboratory. Do not attempt to perform at your home. If you couldn't find a high school in your vicinity, request your tutor to get support and advices.

Experiment 3.2

The effect of temperature on the volume of gas

Objective: To observe the changes in volume of a gas as temperature changes.

Apparatus: Round bottomed flask, beaker, delivery tube and burner.

Procedure:

- 1. Set up the apparatus as shown in Figure 3.6.
- 2. Warm the flask gently with a low Bunsen flame.
- 3. Cool the flask and note what happens.

Air

4. Record your observation.



Beaker

Bunsen burner

Round bottom flask

Figure 3.6: Relationship between temperature and volume of a gas

Observations and analysis:

- 1. What do you observe from the experiment?
- 2. What is your conclusion from this activity?

Self-test Exercise 3.5

- At constant pressure, by what fraction of its volume will a quantity of gas change if the temperature changes from -173 °C to 27 °C? At what temperature will the volume of a gas be
 - a. halved,
 - b. doubled, and
 - c. tripled at constant pressure if the original temperature is 17 °C?
- 2. At 25°C and 1 atm a gas occupies a volume of 1.5 dm³. What volume will it occupy at 100°C and 1 atm?

iii. Gay Lussac's Law

Activity 3.4

Dear learner, attempt to answer the following questions.

- 1. Why does the pressure of car tires increase on a hot day but decrease on a cold morning?
- 2. Why do pressurized containers, such as deodorant or spray-paint cans, have warning labels stating that they must be kept away from fire and stored in a cool environment?

) What is the relationship between pressure and temperature of a gas at constant volume?

Dear learner, Joseph Louis Gay-Lussac (a French scientist) studied how the pressure and temperature of a fixed amount of gas at constant volume are related. The relationship that he established is called Gay-Lussac's law or the pressure-temperature law.

Gay-Lussac's law: states that at constant volume, pressure of a fixed amount of a gas varies directly with the temperature.

Mathematically:

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



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The ratio P to T for any set of volume-temperature values always equals the same \mathbf{k} . That is:

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

Example 3.3

At a temperature of 25 °C, the pressure of the gas in a deodorant can is 3 atm. Calculate the pressure of the gas when it is heated to 845 °C.

Solution:

We are given: Initial conditions: $T_1 = 25 \text{ °C} = 25 + 273 = 298 \text{ K}$, $P_1 = 3.0 \text{ L}$

Final conditions: $T_2 = 845 \text{ °C} = 1,118 \text{ K}, P_2 = ?$

By rearranging Charles' equation, Equation 3.11, P_2 can be calculated:

 $P_2 = \frac{P_1 T_2}{T_1} = \frac{3.0 \text{ atm } \times 1,118 \text{ K}}{298 \text{ K}} = 11.23 \text{ atm}$

Self-test Exercise 3.6

- The pressure of a gas in a cylinder when heated to a temperature of 250 K is 1.5 atm. What is the initial temperature of the gas if its initial pressure was 1.0 atm?
- 2. List some examples of items that we use in our everyday life that obey Gay-Lussac's Law,

The Combined Gas Law

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How do you relate if the three variables: pressure, volume, and temperature of a gas undergo simultaneous changes?

Dear learner, a gas can undergo simultaneous changes in temperature, pressure, and volume. In such cases, the three variables should be considered at the same time. Thus, the three gas laws we have learned can be brought together into a single equation known as the combined gas law.

Combined gas law: expresses the relationship between pressure, volume, and temperature of a fixed amount of gas.

Mathematically:

 $\frac{PV}{\tau} = k$ where, k is a proportionality constant

The ratio PV to T for any set of pressure, volume and temperature values always equals the same k. Thus,

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

Where P_1 , V_1 and T_1 are the initial pressure, volume and temperature; P_2 , V_2 and T_2 are the final pressure, volume and temperature of the gas, respectively.

Example 3.4

A sample of gas occupies 12.0 liters at 240.0 °C under a pressure of 80.0 KPa. At what temperature would the gas occupy 15.0 liters if the pressure were increased to 107.0 KPa?

Solution:

We are given: Initial conditions: $T_{_1}$ = 240 °C = 513 K, $P_{_1}$ = 80.0 KPa, and $V_{_1}$ = 12.0 L

Final conditions: $T_2 = ? V_2 = 15.0 L$ and $P_2 = 107.0 KPa$

By rearranging the combined gas law Equation, we solve for T_2

$$T_{2} = \frac{P_{2}V_{2}T_{1}}{P_{1}V_{1}} = \frac{107 \text{ kPa} \times 15 \text{ L} \times 513 \text{ K}}{80.0 \text{ kPa} \times 12 \text{ L}} = 857.67 \text{ K} \approx 858 \text{ K} = 585 \text{ }^{\circ}\text{C}$$

Self-test Exercise 3.7

- 1. If a 50 cm³ sample of gas exerts a pressure of 60.0 kPa at 35 °C, what volume will it occupy at STP (0 °C and 1 atm)?
- 2. A 280 mL sample of neon exerts a pressure of 660 Torr at 26 °C. At what temperature in °C would it exert a pressure of 940 torr in a volume of 440 mL?

Avogadro's law

How do you mathematically express Avogadro's law?

In 1811, Amedeo Avogadro (an Italian scientist) suggested that at the same temperature and pressure, equal volumes of gases contain equal numbers of moles. This observation is called Avogadro's law, shown as:

$$V \alpha n$$

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

Where V is the volume of the gas, n is the number of moles of gas particles, and k is the proportionality constant.

This equation shows that for a gas at constant temperature and pressure, the volume is directly proportional to the number of moles (molecules) of gas. Similarly, equal numbers of molecules of different gases compared at the same temperature and pressure occupy equal volumes.

For two samples of gas at the same temperature and pressure, the relationship between volumes and numbers of moles can be represented as:

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

The volume occupied by a mole of gas at STP, is called the standard molar volume, which nearly constant for all gases.

The standard molar volume of an ideal gas is equal to 22.4 L per mole at STP. For an ideal gas: 1 mol gas = 22.4 L gas

Example 3.5

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Suppose we have a 12.2 L sample containing 0.50 moles oxygen gas (O_2) at a pressure of 1 atm and a temperature of 25 °C. If all this O_2 were converted to ozone (O_3) at the same temperature and pressure, what would be the volume of the ozone?

Solution:

The balanced equation for the reaction is

 $3O_2(g) \longrightarrow 2O_3(g)$

To calculate the moles of O_3 produced, we must use the appropriate mole ratio: Mol of O_3 (g) = 0.5 mol $O_2 \times \frac{2 \mod O_3}{3 \mod O_2} = 0.33 \mod$

Then, to determine volume of 0.33 mol $O_{_3}$ the following Equation will be used.

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Where $V_{_1}$ is the volume of $n_{_1}$ moles of $O_{_2}$ gas and $V_{_2}$ is the volume of $n_{_2}$ moles of $O_{_3}$ gas.

In this case we have

 $n_{1} = 0.5 \text{ mol} \qquad n_{2} = 0.33 \text{ mol}$ $V_{1} = 12.2 \text{ L} \qquad V_{2} = ?$ Thus, $V_{2} = \frac{V_{1}n_{2}}{n_{1}} = \frac{12.2 \text{ L} \times 0.33 \text{ mol}}{05 \text{ mol}} = 8.052 \text{ L} = 8.0 \text{ L}$

Self-test Exercise 3.8

One mole of a gas occupies 27.0 L, and its density is 1.41 g/L at a particular temperature and pressure.

- a. What is its molecular weight?
- b. What is the density of the gas at STP?

vi. The Ideal Gas Equation



What is an ideal gas and how do you derive its equation of state?

Dear learner, an ideal gas is a hypothetical gas that obeys the gas laws. Real gases only obey the ideal gas laws closely at high temperature and low pressure. Under these conditions, their particles are very far apart.



$$V \alpha n \frac{1}{P}$$

 $V = Rn \frac{T}{P}$

Where **R**, is proportionality constant called the ideal gas constant. By rearranging this Equation, we obtain the following ideal gas equation:

$$VP = nRT$$

Thus, the ideal gas equation describes the relationship among the four variables P, V, T and n. An ideal gas is a gas whose pressure, volume and temperature behavior can be completely explained by the ideal gas equation.

At STP (n = 1 mol, V = 22.4 L, T = 273 K (0 °C) and P = 1 atm), the values of R can be calculated from the ideal gas equation:

$$R = \frac{PV}{nT} = \frac{1 \text{ atm } \times 24 \text{ L}}{1 \text{ mol } \times 273 \text{ K}} = 0.082 \text{ atm. L/mol. K}$$

Example 3.6

 What pressure, in atm, is exerted by 54.0 g of Xe in a 1.00 L flask at 20.0 °C? (Molar mass of Xe = 131.3 g/mol)

Solution: By rearranging ideal gas Equation we obtain:

$$P = Rn \frac{T}{V}$$

But, n = m/M, where m is given mass and M is molar mass of a substance. Thus,

$$P = R \frac{mT}{MV}$$

$$P = 0.082 \quad \frac{\text{atm} \pm}{\text{mol} \text{ K}} \frac{54 \text{ g} \times 293 \text{ K}}{131.3 \frac{\text{g}}{\text{mol}}} = 9.88 \text{ atm}$$

2. At STP, 0.280 L of a gas weighs 0.400 g. Calculate the molar mass of the gas.

Solution:

We are given: V = 0.280 L, m = 0.400 g, T = 273 K and P = 1 atm So to solve the molar mass of the gas, we use the following formula

$$M = R \frac{mT}{VP}$$

= 0.082 $\frac{\text{atm} \pm}{\text{mol} + \frac{0.400 \text{ g} \times 73 \text{ K}}{0.280 \pm \times 1 \text{ atm}}}$ = 31.98 g/mol

$\boldsymbol{\zeta}$)Self-test Exercise 3.9

- The density of a gas at a pressure of 1.34 atm and a temperature of 303
 K is found to be 1.77 g/L. What is the molar mass of this gas?
- A helium-filled weather balloon has a volume of 7240 ft³. How many grams of helium would be required to inflate this balloon to a pressure of 745 torr at 21 °C? (1 ft³ = 28.3 L)

Graham's Law of Diffusion

Dear learner, you have earlier learned that the gas molecules are in constant, rapid, random motion and occupy quickly throughout any container.

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The spreading of gas molecules throughout the container is called diffusion.

Why does one gas diffuse faster than the other?

In 1829, Thomas Graham (a Scottish chemist) studied the rate of diffusion of different gases. He found that gases having low densities diffuse faster than gases that have higher densities. Based on his observation, he postulated the relationship between the density of a gas and its rate of diffusion, which is known as Graham's law of diffusion.

Graham's law of diffusion states that at constant temperature and pressure, the rate of diffusion of a gas, r, is inversely proportional to the square root of its density, d, or molar mass, M.

Mathematically it is expressed as:

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$$r \alpha \sqrt{\frac{1}{d}}$$
 or $r \alpha \sqrt{\frac{1}{M}}$

Where **r** is the rate of diffusion, **d** is the density and **M** is the molecular mass of the gas. For two gases (Gas 1 and Gas 2), their rates of diffusion can be given as:

$$r \alpha \sqrt{\frac{1}{d_1}}$$
 or $r \alpha \sqrt{\frac{1}{M_1}}$
 $r \alpha \sqrt{\frac{1}{d_2}}$ or $r \alpha \sqrt{\frac{1}{M_2}}$

By rearranging these Equations, we obtain the following equation:

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} \qquad \text{or} \qquad \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

Where, r_1 , d_1 and M_1 represent the rate of diffusion, density and molecular mass of gas. r_2 , d_2 and M_2 represent the rate of diffusion, density and molecular mass of gas 2.

Activity 3.6

Dear learner, why do you think that a hydrogen-filled balloons deflate over time faster than air-filled balloons? (Hint: A balloon has many invisible tiny holes.)

Dear learner, to understand more about the rate of diffusion of gaseous you are advised to perform the experiment given below. To conduct the experiment, visit a high school in your vicinity and request Grade 11 chemistry teacher to help you in performing or demonstrating the experiment. Note that the experiment should be performed only in the laboratory. If you couldn't find a high school in your vicinity, request your tutor to get support and advices.

Experiment 3.3

Determination of diffusion of gases

Objective: To compare the rate of diffusion of two different gases.

Apparatus: Porous pot, cork, delivery tube, U-tube, and beaker

Chemicals: Colored water, hydrogen and CO₂.

Procedure:

- 1. Set up the apparatus as shown in Figure 3.7.
- 2. Fit the porous pot with a cork and delivery tube
- 3. Attach the delivery tube to a U-tube containing coloured water
- 4. Place the porous pot in beaker of hydrogen gas
- 5. Watch the change in the level of the colored water in the U-tube
- 6. Pressure in pot increases as hydrogen or diffuses in quicker than the air diffuses out
- 7. Repeat the experiment by putting the porous pot in beaker of carbon dioxide



Figure 3.7: The determination of diffusion of gases

Observations and analysis:

- 1. In which case pressure in the pot decreases?
- 2. How did you observe the changes in pressure?
- 3. Write your conclusions about the experiment.

Example 3.7

 Which gas will diffuse faster, oxygen or carbon dioxide? What is the relative rate of their diffusions? Solution:

The molecular weight of CO_2 is 44 g/mol and that of O_2 is 32g/mol. Therefore, NH_3 diffuses faster than CO_2 .

- Let the rate of diffusion of O₂ be
- Let the rate of diffusion of CO₂ be

We can calculate the rate of diffusion using the following equation:

$$\frac{r_{1}}{r_{2}} = \sqrt{\frac{M_{2}}{M_{1}}}$$
$$\frac{r_{O_{2}}}{r_{CO_{2}}} = \sqrt{\frac{M_{CO_{2}}}{M_{O_{2}}}} = \sqrt{\frac{44 \text{ g/mol}}{32 \text{ g/mol}}} = 0.73$$

This means, rate of diffusion of O_2 is 0.73 times that of CO_2 .

2. The rate of diffusion of methane (CH_4) is twice that of an unknown gas. What is the molecular mass of the gas?

Solution:

Let and rx be the rates of diffusion of CH4 and the unknown gas as well as and Mx be the molecular masses of CH4 and the unknown gas, respectively. The rate of diffusion of CH4 is two times faster than the unknown gas. This can be written mathematically as $r_{CH_4} = 2r_x$.

Now, substitute $2r_x$ in place of r_{CH_4} and solve for M_x using Graham's law equation:

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$
$$\frac{r_{CH_4}}{r_x} = \sqrt{\frac{M_x}{M_{CH_4}}}$$

By substituting $2r_x$ in place of r_{CH_4} we obtain

$$\frac{2\tau_{x}}{\tau_{x}} = \sqrt{\frac{M_{x}}{M_{CH_{4}}}}$$
Thus, 2 = $\sqrt{\frac{M_{x}}{M_{CH_{4}}}}$

By rearranging this equation, we obtain the value of M_{x}

$$M_x = 2^2 M_{CH_4} = 4 \times 16 \frac{g}{mol} = 64 g/mol$$

Therefore, the molecular mass of the unknown gas is 64.

The rate at which a gas diffuses is also inversely proportional to the time if it takes to move a certain distance.

Mathematically:

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$$r \alpha \frac{1}{t}$$

If two different gases (gas 1 and gas 2) under the same temperature and pressure diffuse through a porous container, then the time required to diffuse for the two gases can be given by the following formula:

$$\frac{r_{1}}{r_{2}} = \frac{t_{2}}{t_{1}} = \sqrt{\frac{M_{2}}{M_{1}}}$$

Where t_1 and t_2 are the time taken, r_1 and r_2 are the rates, M_1 and M_2 are the molecular masses of gas 1 and gas 2, respectively.

Self-test Exercise 3.10

- A sample of ethane, C₂H₆, diffuses at a rate of 3.6 x 10⁻⁶ mol/h. An unknown gas, under the same diffuses at a rate of 1.3 x 10⁻⁶ mol/hr. Calculate the molar mass of the unknown gas.
- 2. Which gas in each of the following pairs diffuses more rapidly, and what are the relative rates of diffusion?

a. Kr and O_2

b. N_2 and acetylene (C_2H_2)

CHECKLIST 3.3

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

l can ...

- explain the assumptions of kinetic molecular theory of gases
- describe the properties of gases using kinetic molecular theory
- describe the behavior of gases by using the variables V (volume), T (Temperature), P (pressure) and n (number of moles)

٠	state Boyle's and perform an activity to show changes in volume and	
	pressure of gases to illustrates Boyle's law	
•	apply Boyle's law in solving problems	
•	state Charles' law, and conduct an activity to show changes in	
	volume and temperature of gases to illustrate Charles' law	
•	apply Charles' law in solving problems	
•	derive combined gas law equation from Boyle's law and Charles' law	
•	use the combined gas law to calculate changes in volume, pressure	
	or temperature	
•	define an ideal gas, and derive an ideal gas equation from Boyle's	
	law, Charles' law and Avogadro's law	
•	compare the nature of real gases with ideal gases	
•	solve problems related to ideal gas equation	
•	define diffusion, and state Graham's law of diffusion	
•	carry out an activity to compare the rate of diffusion of two different gases	
•	apply Graham's law of diffusion in solving problems.	

Section 3.5: The Liquid State

Dear learner, in this section you will learn about the physical properties of liquids and different terms including evaporation, boiling, condensation, vapor pressure; boiling point, molar heat of vaporization and molar heat of condensation. To learn more about the properties of liquids you will also perform some practical activities.

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

- explain the terms: evaporation, boiling, condensation, vapor pressure; boiling point, molar heat of vaporization and molar heat of condensation
- carry out an activity to demonstrate the concept of vapor pressure
- carry out an activity to determine the boiling points of water and ethanol.

Dear learner, from your previous lessons you have learned that liquids have a definite volume, but an indefinite shape. They take the shape of their containers to the level they fill. Liquids are denser than gases, but less dense than solids.

Why liquids are more ordered than gases?

As in a gas, particles in a liquid are in constant motion. However, the particles in a liquid are closer together than those in a gas. The attractive forces between particles in a liquid are stronger than between particles of a gas. This attraction between liquid particles is caused by the intermolecular forces of attraction such as dipole-dipole forces, London dispersion forces, and hydrogen bonding.

Liquids are more ordered than gases because of the stronger intermolecular forces and the lower mobility of liquid particles.

Energy Changes in Liquids



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What is evaporation and how it takes place?

Dear learner, the process by which a liquid change to a gas is known as vaporization or evaporation. Evaporation is the process by which molecules on the surface of a liquid break away and goes into the gas phase (Figure 3.8). Molecules on the surface of a liquid are held less tightly than those in the interior, so the most energetic molecule can break away into the gas.



Figure 3.8: Evaporation of liquid

How can evaporation related to energy of molecules?

Evaporation is explained in terms of the energy that the molecules on the surface of the liquid have. In an open container, evaporation continues until all of the liquid enters the gas (vapor) phase (Figure 3.8). Most of the molecules that escape into the vapor phase do not collide with the surface of the liquid and return to the liquid phase. Instead, they will diffuse through the gas phase away from the container and equilibrium cannot be established. Under these conditions, the liquid will continue to evaporate until it has disappeared.

However, liquids in a closed container behave differently. The volume of the liquid decreases for some time, and then, remains unchanged. This is because in closed containers, the vapor cannot escape. As more molecules leave the liquid, more gaseous molecules collide with the container walls, with each other, and with the liquid surface and return to the original liquid state. The process of when a vapor returns to the liquid state is called condensation.

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Chemistry Grade 11 Distance Module -I

) Activity 3.7

Dear learner, you usually observe a fog (mist) on the bathroom mirror after you shower with hot water. Why it gets foggy (misty)?

'Evaporation and condensation are opposing processes. Evaporation is a liquid turning into a gas, and condensation is a gas turning into a liquid.

For example, when liquid water is initially put into a closed container, more evaporation happens than condensation because there are so few gaseous water molecules in the space above the water (Figure 3.9a). However, as the number of gaseous water molecules increases, the rate of condensation also increases (Figure 3.9b). At the point where the rates of condensation and evaporation become equal (Figure 3.9c), dynamic equilibrium is reached and the number of gaseous water molecules above the liquid remains constant.

The vapor pressure of a liquid is the partial pressure of its vapor in dynamic equilibrium with its liquid.



Figure 3.9: Evaporation and condensation: (a) when water is first put into a closed container, water molecules begin to evaporate; (b) Evaporation continues, but condensation also begins to occur; and (c) Dynamic equilibrium: rate of evaporation equals the rate of condensation



The rate of evaporation of a liquid can be affected by three factors:

- temperature
- intermolecular forces, and
- surface area of the liquid.

Temperature

P How does temperature affect the rate of evaporation of a liquid?

An increase in temperature increases the average kinetic energy of the molecules and thus increases the tendency to change into the gaseous state. Some liquids evaporate readily at room temperature. Such liquids are said to be volatile. Volatile liquids have relatively weak forces of attraction between particles. Liquids such as formaldehyde, ethyl alcohol, mercury, and benzene are volatile liquids.

Liquids that do not vaporize easily at a given temperature are said to be nonvolatile. They have relatively stronger attractive forces between their molecules. Liquids such as motor oil, edible oil, glycerin, water, and molten ionic compounds are nonvolatile.

Vapor pressure

What is vapor pressure of a liquid?

The partial pressure of the vapor above a liquid is called vapor pressure. The vapor pressure of a liquid depends up on the temperature. At a given temperature, vapor pressure is constant. The vapor pressures of liquids always increase as temperature increases because the rate of evaporation increases with increasing temperature.

Intermolecular forces

?) What is the impact of intermolecular forces on vapor pressure of liquid?

Hello, vapor pressure depends also on the intermolecular forces between the molecules of the liquid. The stronger the intermolecular forces, the lower will be the vapor pressure of the liquid. Because fewer molecules will have enough kinetic energy, it overcomes the attractive force at a given temperature.

For example, water and ethanol have relatively low vapor pressure because the strong hydrogen bonding in these liquids, accounting for their unusually low vapor pressures. However, liquids with low intermolecular forces have high vapor pressures at room temperature. For example, diethyl ether, a non-polar molecule with relatively weak dispersion forces, has a relatively higher vapor pressure.

Surface area of the liquid

What is surface area of the liquid?

Evaporation occurs at the liquids surface, where it interfaces with the air surrounding it. Liquid surface area is the amount of liquid that is exposed to the atmosphere (air).

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How does surface area of a liquid affect the rate of evaporation?

As you previously learned, liquids have no definite shape; instead, they take the shape of containers in which they are contained. So, the number of liquid molecules present at the surface increases as the surface area of the container (or the surface area occupied by the liquid) increases. This, in turn, increases the number of molecules that separate from the liquid surface and change into gases at a specific moment, increasing the rate of evaporation.

However, for liquids having the same surface area, the rate of evaporation depends on the two factors: temperature and the strengths of intermolecular attractions.

Boiling and Boiling Point

) What is boiling?

Dear learner, heating a liquid always increases its vapor pressure. When a liquid is heated to a sufficiently high temperature under a given applied (usually atmospheric) pressure, bubbles of vapor begin to form below the surface. If the vapor pressure inside the bubbles is less than the applied pressure on the surface of the liquid, the bubbles collapse as soon as they form. However, if the temperature is raised sufficiently, the vapor pressure is high enough that the bubbles can persist and rise to the surface, and burst, releasing the vapor into the air. This process is called boiling.

In short, boiling is the change of a liquid to bubbles of vapor that appear throughout the liquids. It is the conversion of liquid to vapor within the liquid as well as at its surface. During evaporation, only molecules at the surface escape into the vapor phase.

Can you define boiling point?

(B

Hello learner, if the temperature of the liquid is increased, the vapor pressure also increases. Finally, the boiling point is reached. The boiling point is the temperature at which its vapor pressure of a liquid is equal to the external atmospheric pressure. The normal boiling point is the temperature at which the vapor pressure of a liquid is equal to exactly one atmosphere.

For instance, water boils at 100 °C at 1.0 atmospheric pressure and thus, its normal boiling point of water is 100 °C. Nevertheless, the boiling point of water at 0.83 atm is 95 °C. Generally, when the pressure exerted on a liquid is varied, the boiling point of a liquid also varies.

Note that as heat energy is added to vaporize a pure liquid at its boiling point, the temperature remains constant until the liquid is completely changed to vapor. Then, the temperature begins to rise after the liquid is completely changed to vapor.

Dear learner, to understand about the vapor pressure and the boiling point of liquid, perform the following experiments 3.4 and 3.5, respectively. To conduct the experiment, visit a high school in your vicinity and request a teacher who is teaching Grade 11 chemistry to help you in performing or demonstrating the experiment. Note that the experiment should be performed only in the laboratory. If you couldn't find a high school in your vicinity, request your tutor to get support and advices.

Experiment 3.4

Observing the vapor pressure of liquid

Objective: To observe the vapor pressure of liquid.

Apparatus: Erlenmeyer flask, rubber bung, U-tube and burner.

Procedure:

- 1. Set up the apparatus as shown in Figure 3.10.
- 2. Add about 100 mL of water into the Erlenmeyer flask and put a stopper. Heat the flask to expel the air above the water in the flask.
- 3. Half fill the U-tube with water.
- 4. Connect the U-tube to Erlenmeyer flask and note the water level in the two arms of the U-tube.
- 5. Heat the flask gently and observe the water level changes in the arms of the U-tube.





Observations and analysis:

What do you observe? Give an explanation for the observation.

Experiment 3.5

Determining of boiling points

Objectives: To determine the boiling point of water and ethanol.

Apparatus: Test tube, stopper, thermometer, beaker, burner, clamp, stand and base.

Procedure:

- 1. Half fill the test tube with a sample of pure water and add some.
- 2. Take a rubber stopper and pierce a thin opening on the side of the rubber stopper to allow the vapor to escape.
- 3. Fit the thermometer with the rubber stopper and insert it in the test tube.
- 4. Put the test tube in a beaker containing oil as shown in Figure 3.11.
- 5. Heat the oil in beaker gently and record the temperature at which the water boils.
- 6. Repeat the experiment by replacing the water sample by ethanol.



Figure 3.11: Determination of boiling point

Observations and analysis:

- 1. What is the boiling point of water and ethanol that you obtained from these experiments?
- 2. Does the temperature from the thermometer reading increase after the water and ethanol started to boils?
- 3. Explain why the thermometer was not put into the liquid.
- 4. What is the purpose of adding porcelain chips?


Dear learner, boiling of a liquid requires a certain amount of heat energy to break the forces of attraction between the molecules. The amount of heat energy necessary to bring about the vaporization of a fixed amount of a liquid at a fixed temperature is called the heat of vaporization. For example, the heat of vaporization per mole of water at 298 K and 1 atmosphere is 44.0 kJ. This is called the molar heat of vaporization (ΔH_{yap}) of water.

The molar heat of vaporization (ΔH_{vap}) is the amount of heat needed to convert 1 mole of a liquid at its boiling point to a gas.

The molar heat of vaporization is equal to the amount of energy that is released when 1 mole of vapor condenses to liquid at the condensing point of a vapor.

) What is condensation?

Condensation is the change of a gas to a liquid at the boiling point of the liquid. During condensation heat is released. Thus, the amount of heat realised when 1 mole of a gas is converted to a liquid at its condensation point is called the molar heat of condensation (ΔH_{cond}).



(B

What is the relationship between molar heat of vaporization and molar heat of condensation?

Molar heat of vaporization (ΔH_{vap}) and molar heat of condensation (ΔH_{cond}) are equal in magnitude but opposite in sign:

 $\Delta {\rm H}_{\rm vap} = - \Delta {\rm H}_{\rm cond}$

Note that vaporization is an endothermic process whereas condensation is an exothermic process.



- 2. What are the effects of impurity on the boiling point of liquids?
- 3. Why does the boiling point of liquids decrease as altitude increases?

✓ CHECKLIST 3.4

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

l can ...

- explain the terms such as evaporation, boiling, condensation, vapor pressure, boiling point, molar heat of vaporization and molar heat of condensation
- carry out an activity to demonstrate the concept of vapor pressure
- carry out an activity to determine the boiling points of water and ethanol.

Section 3.6: The Solid State

Dear learner, in this section you will learn about the physical properties of solids. When you learn this section give emphasis to the following terms: phase change, melting, fusion, sublimation, melting point, freezing point, molar heat of solidification. To more understand the properties of solids, you should perform the suggested practical activities.

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

- describe phase changes
- explain the terms melting, fusion, sublimation, melting point, freezing point, molar heat of solidification
- explain temperature changes associated with phase changes
- determine melting point of ice
- demonstrate an experiment to show the phase changes from ice to liquid water and then to water vapor.

?

(B

Why do you relate freezing point with melting point?

Dear learner, the temperature at which a pure liquid changes to a crystalline solid, or freezes, is called the freezing point; it is identical to the melting point.

The melting or freezing point occurs at the temperature where the liquid and solid are in dynamic equilibrium.

$(\ref{eq: the difference between melting and freezing?}$

When a solid is continuously heated the ordered crystalline structure of solids will be disturbed. As a result, particles gradually get freedom of motion and melting (or fusion) take place. Melting is the process of converting solid into the liquid. The temperature at which a crystalline solid change to a liquid or melts is called the melting point.

On the contrary, when a liquid is cooled, its molecules come closer to one another and thus, the intermolecular forces of attraction between them get stronger. As a result, its particles arrange themselves into a regular pattern and then converted to a solid. This process is called freezing or solidification. For instance, ice melts at 0 °C and water freezes at 0 °C. Ice and water coexist in equilibrium at 0 °C as follows:



(P

Describe heat of fusion and why it is in needed?

Dear learner, the amount of heat needed to convert one gram of solid to liquid at the melting point is called heat of fusion. The molar heat of fusion (ΔH_{fus}) is the quantity of heat needed to convert one mole of a solid at its melting point to the liquid state.

For example, the molar heat of fusion of ice is 6.01 kJ at 0°C. This is the amount of energy needed to break the attractive forces in the solid, ice, at its melting point.

Melting requires the supply of energy; therefore, it is an endothermic process.

During the process of solidification, the amount of heat liberated is exactly equal to the heat of fusion. The heat liberated is called the heat of solidification or heat of crystallization.

The molar heat of crystallization (ΔH_{cryst}) is the quantity of energy that is removed from one mole of a liquid to convert it to the solid state at its freezing point.

 $\Delta H_{cryst} = -\Delta H_{fus}$



What is sublimation? Which substances can sublimate?

Dear learner, some solids evaporate directly from the solid to the vapor state without passing through the liquid state. This process is called sublimation that is the change of solid to vapor. The opposite of sublimation is deposition that is the change of vapor to solid. The process can be expressed as follows:

The enthalpy of sublimation (ΔH_{sub}) is the energy required to convert one mole of a substance from the solid to the gaseous state. Sublimation is an endothermic process. The enthalpy of deposition, ΔH_{dep} , is the energy released when one mole of a substance from the gaseous state change to solid.

The heat (enthalpy) of sublimation is related to the enthalpies of fusion and vaporization by:

$$\Delta H_{sub} = \Delta H_{fus} + \Delta H_{vap}$$

$$\Delta H_{sub} = -\Delta H_{dep}$$



Dear learner, attempt to answer the following question

- 1. When the crystals of iodine are warmed, they disappear into vapors without being changed into liquid.
- 2. When ethyl alcohol (ethanol) is taken in an open container it disappears after sometime.

Heating curve

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What is heating curve? What do you learn from the heating curve?

Dear learner, a heating curve is a plot of the temperature versus the amount of heat added (Figure 3.12). It is commonly used to show the relationship between phase changes and enthalpy for a given substance. There are two main observations that you can learn from the heating curve:

- i. regions where the temperature increases as heat is added, and
- ii. plateaus where the temperature stays constant.

Heat curve is at plateaus that a phase change occurs.



Figure 3.12: A heating curve

Hello, to understand more phase changes you should perform the following Experiment 3.6. To conduct the experiment, visit a high school in your vicinity and request a teacher who is teaching Grade 11 chemistry to help you in performing or demonstrating the experiment. Note that the experiment should be performed only in the laboratory. If you couldn't find a high school in your vicinity, request your tutor to get support and advices.

Experiment 3.6

Demonstration of phase changes

Objective: to show the phase changes from ice to water and then to water vapour.

Apparatus: Beakers, thermometer, glass rod, burner, clamp stand, tripod stand and wire gauze.

Procedure:

- 1. Take a 250 mL beaker and half-filled with the dry crushed ice given by your teacher
- 2. Place the beaker on a wire gauze kept over a tripod stand.
- 3. Suspend a thermometer from the clamp stand such that the bulb of the thermometer is completely surrounded by ice as shown in **Figure 3.13**.
- 4. Heat the ice cubes and stir continuously for uniform heating.
- 5. Record the temperature (T_1) when ice starts melting.
- 6. Heat continuously till ice melts completely. Record this temperature (T_2) .

- 7. Continue heating, record at the temperature (T_3) when water begins to boil.
- 8. Continue to heat the water for at least 3 minutes after the water begins to boil, record the temperature (T_4) after 3 minutes.



Figure 3.13: Phase changes from ice to water and then to water vapour Observations and analysis:

- 1. At any time did the temperature seem to rise at or near a constant rate?
- 2. When did the temperature change the slowest?
- 3. When did the temperature change the fastest?
- 4. Is there any difference between T_1 and T_2 as well as between T_3 and T_4 ?

Self-test Exercise 3.12

- 1. Draw a cooling curve.
- 2. What is the relationship and difference between a heating curve and a cooling curve?

CHECKLIST 3.5

In the boxes provided for each of the following tasks put a tick mark ($\sqrt{}$) if you can perform or concepts you are familiar with them.

l can ...

- describe phase changes
- explain the terms melting, fusion, sublimation, melting point, freezing point, molar heat of solidification
- explain temperature changes associated with phase changes
- determine melting point of ice
- performed experiment and observed the phase changes from ice to liquid water and then to water vapor.

Unit Summary

- Matter is anything that has mass and occupies space.
- Matter exists in one of the three states: a solid, a liquid or a gas.
- Solids have a definite volume and a definite shape.
- A liquid has a definite mass, but doesn't have a definite shape.
- A gas has neither a definite volume nor a definite shape.
- According to Boyle's law at fixed temperature the volume of a fixed mass of gas is inversely proportional to the pressure.
- Charles' law states that at constant pressure the volume of a fixed mass of gas is directly proportional to the temperature.
- According to Gay-Lussac's law the pressure of a fixed amount of a gas varies directly with the temperature.
- Combined gas law states that the ratio PV to T for any set of pressure, volume and temperature values always equals the same k.
- Avogadro's stated that at the same temperature and pressure, equal volumes of gases contain equal numbers of moles.
- The ideal gas law is a combination of Boyle's law, Charles' law and Avogadro's law.
- At constant temperature and pressure, the rate of diffusion of a gas, r, is inversely proportional to the square root of its density, d, or molar mass, M.

• Energy is released when a less ordered phase is converted to a more ordered phase as follows:

	•				
•	Phase Change	•	Physical Process	•	Energy Change
•	Fusion (melting)	•	Solid \rightarrow Liquid	•	Energy is absorbed
	Vaporization		Liquid \rightarrow Gas		Energy is absorbed
	Sublimation		Solid \rightarrow Gas		Energy is absorbed
	Freezing		$Liquid \to Solid$		Energy is released
	Condensation		Gas \rightarrow Liquid		Energy is released
	Deposition		$Gas \rightarrow Solid$		Energy is released

🛠 SELF-ASSESSMENT

Part I: Write what each of these refers to:

- 1. They have a definite shape and a definite volume.
- 2. They can take the shape of their container.
- 3. They can move in all direction at high speed.
- 4. They can easily diffuse through each other.
- 5. They have less density relative to the other states.

Part II: Complete the following statements

- 6. Liquids and gases have no definite ______.
- 7. The freezing point of a liquid is the same as ______.
- The temperature at which a crystalline solid is directly converted to a gas is called ______.
- 9. The lowest attainable temperature is ______.
- "At the fixed temperature and pressure, equal volumes of different gases have equal numbers of molecules." This statement is postulated by______.

Part III: Multiple choice questions

Choose the correct answer from the given alternatives

- 11. Which of the following would take up the entire shape of the container when it is enclosed within?
 - a. Gas c. Solid
 - b. Liquid d. d. both a and b
- 12. If a gas expands at constant temperature, it indicates that:
 - a. Number of the molecules of gas increases
 - b. Kinetic energy of molecules decreases
 - c. Pressure of the gas increases
 - d. Kinetic energy of molecules remains the same
- 13. A sample of oxygen occupies 47.2 L under a pressure of 1240 torr at 25°C. What volume would it occupy at 25°C if the pressure were decreased to 730 torr?
 - a. 27.8 L c. 29.3 L
 - b. 32.3 L d. d. 80.2 L

14. If helium (He) diffuses through a porous barrier at a rate of 4.0 moles per minute, at what rate (in moles per minute) would oxygen (O₂) gas diffuse?

- a. 1.4 c. 0.7
- b. 2.0 d. 0.5
- 15. What will happen when water freezes?
 - a. Water changes from a gas to a liquid state
 - b. Water changes from a liquid to a solid state
 - c. Water changes from a solid to a gas state
 - d. Water changes from a gas to a plasma state

Part IV: Solve the following problems

- 16. A vessel of 120 mL capacity contains a certain amount of gas at 35 °C and 1.2 bar pressure. The gas is transferred to another vessel of volume 180 mL at 35 °C. What would be its pressure?
- 17. 34.05 mL of phosphorus vapor weighs 0.0625 g at 546 °C and 0.1 bar pressure. What is the molar mass of phosphorus?
- 18. If 0.500 mole of nitrogen gas occupies a volume of 11.2 L at 0°C; what volume will 2.00 mole of nitrogen gas occupy at the same temperature

and pressure?

- 19. A sample of a certain gas has a volume of 1.25 L at -125°C and 5.0 atm. The gas is compressed at 50.0 atm to a volume of 325 mL. What is the final temperature of the helium gas in °C?
- 20. The temperature of 2.5 L of a gas initially at STP is raised to 250°C at constant volume; calculate the final pressure of the gas in atmosphere.

Module I Written Assignment Questions

Dear learner, you are supposed to work on the following assignment questions after completing Module II. The questions were designed from module 1 contents. Please, attempt to answer all the questions on separate answer sheet and submit it to your tutor for correction as a part of your assessment.

- 1. How does Dalton's atomic theory account for the law of mass conservation and the law of definite proportions?
- 2. There are four naturally occurring isotopes of strontium:
 - a. How many protons, neutrons, and electrons are in Sr-87?
 - b. What is the most abundant isotope in a strontium sample?
 - c. How many neutrons are in Sr-84?
 - d. Why don't any of the isotopes of strontium have the atomic mass of 87.62 amu listed on the periodic table?
- 3. One type of electromagnetic radiation has a frequency of 102 MHz, another type has a wavelength of 3.0 10-10 m, and another type of electromagnetic radiation has photons with energy equal to 2.08 10-19 J/ photon. Identify each type of electromagnetic radiation and place them in order of increasing photon energy and increasing frequency.
- 4. Calculate the wavelength of light emitted when each of the following transitions occur in the hydrogen atom.

a. n = 5 to n = 2 b. n = 4 to n = 3 c. n = 2 to n = 1

5. Give the maximum number of electrons in an atom that can have these quantum numbers

a. n = 3 b. n = 4, $m_{\ell=-1}$ c. n = 3, $\ell_{=} 2$ d. n = 2, $\ell_{=} 1, m_{\ell} = 0, m_{s} = -\frac{1}{2}$

- 6. Write the ground-state electron configurations for the following elements using Aufabu's building principle and orbital diagram.
 - a. Ga (Z = 31) b. Kr (Z = 36) c. Sr (Z = 38) d. Sn (Z = 50)
- 7. Which of the following elements Li, Be, N, and F:
 - a. is an alkaline earth metal?
 - b. has the largest atomic size?
 - c. has the highest ionization energy?
 - d. is found in Group 5A (15)?
 - e. has the most metallic character?
- 8. Complete the following table:

Name	Atomic	Number of	Number of	Number of
	Symbol	Protons	Neutrons	Electrons
		28	34	
Magnesium			14	

What common characteristics do elements within a group of the periodic table have?

9. 1. For each of the following, draw the Lewis structure and determine the shape:

d. NH_2CI (N is the central atom) e. TeO_2

10. Identify the errors in each of the following Lewis structures and draw the correct formula:

a. :Ci=O=Ci: b.
$$H-C-H$$
 c. $H-N=O-H$

- 11. Classify each of the following as ionic or molecular, and name each:
 - a. Na₂O b. NF₃ c. BaCO₃ d. WO₂ e. CS₂ f. Cs₃PO₄ g. K(AuCl₄)

12. Indicate the major type of intermolecular forces: ionic bonds, dipole-dipole attractions, hydrogen bonds, dispersion forces that occurs between particles of the following:

a. NF_3 b. CIF c. Br_2 d. Cs_2O e. C_4H_{10} f. CH_3OH

- 13. Predict the shape and polarity of each of the following molecules, which have polar covalent bonds:
 - a. A central atom with three identical bonded atoms and one lone pair.
 - b. A central atom with two bonded atoms and two lone pairs.
 - c. A central atom with four identical bonded atoms and no lone pairs.
 - d. A central atom with four bonded atoms that are not identical and no lone pairs.
- 14.State the following gas laws in words and also in the form of an equation: Boyle's law, Charles's law, Avogadro's law, Gay-Lussac's law, Graham's law. In each case, indicate the conditions under which the law is applicable, and give the units for each quantity in the equation.
- 15. Explain the reason why the density of a gas much lower than that of a liquid or solid under atmospheric conditions.
- 16. A 2.50-L container is filled with 175 g argon.
 - a. If the pressure is 10.0 atm, what is the temperature?
 - b. If the temperature is 225 K, what is the pressure?
- 17. A cylinder of oxygen gas contains 80 g O_2 . If the volume of the cylinder is 8.58 L, what is the pressure of the O_2 if the gas temperature is 21°C?
- 18. Which gas in each of the following pairs diffuses more rapidly, and what are the relative rates of diffusion?
 - a. SO_2 and CO_2 b. O_2 and N_2 d. CO and N_2
- 19. Explain the differences between the following processes?
 - a. Melting point and boiling point
 - b. Condensation and solidification
 - c. Sublimation and Deposition
 - d. Heat of condensation and heat of fusion

ANSWER KEYS TO ACTIVITIES

Activity 3.1

- The source of ice rain (precipitation) is water vapor. The water vapor in the air must be cooled to allow water to be condensed into cloud droplets. These droplets then grow to form precipitation particles. Ice is solid, but water is liquid. Ice has rigid structure and but water has not. Ice has low density than water.
- 2. Up on gently heating, the water in the dish evaporates or changes to vapor.



1. The motion, distance and attraction between particles of these substances are summarized as follows

Particles	Motion ofmole- cules	Distances betweenmol- ecules	Attraction between molecule
Cu(solid)	low	very close to each other	high
H ₂ O (liquid)	medium	close to each other	medium
O ₂ (gas)	high	very far apart	low

2. In the gaseous state, the molecules possess the highest kinetic energy, since the motion of particles of gases is higher as compared to those of liquids and solids.



- a. A is gas and thus it has the highest kinetic energy. Because, when the temperature of an object increases, the average kinetic energy of its particles increases.
- b. B is liquid has the highest density, because liquid water has greater density than ice (solid water)
- c. C is ice (solid) and thus, it has the most regular arrangement of molecules.

E Activity 3.4

- The pressure in automobile tires increases on a hot day because the pressure of a gas is directly proportional to the temperature. Thus, automobile tires increase when driving on hot roads or in high ambient temperatures. Tires absorb heat when exposed to direct sunlight, causing their air chamber temperature to rise.
- 2. Because at higher temperatures, their pressure will rise, and thus, they can explode.

E Activity 3.5

- When too much air is added to a balloon, the number of molecules and the pressure inside the balloon increase, but the volume remains the same. Finally, the balloon bursts since its volume cannot increase any further with the increase in molecules and pressure.
- 2. The volume and temperature are constant, but its pressure and number of moles gases particles increase, causing the balloon to burst.
- 3. Yes, we can.

Activity 3.6

A light gas diffuses faster than a heavy gas. Thus, the balloon filled with hydrogen deflates over time more than the balloon filled with air because hydrogen gas diffuses faster through its invisible holes than the balloon filled with air. It happens since hydrogen is lighter than air.

E Activity 3.7

Bathroom mirror get foggy (misty) when steam from hot water condenses on the mirror. Steam emerging from hot water can condense on a colder surface.



- An iodine crystal vaporizes directly into the gaseous state when it is warmed. The direct change of a substance from the solid to the gaseous state is known as sublimation. Therefore, iodine undergoes sublimation.
- 2. When ethyl alcohol is placed in an open container, it disappears after some time because it is a volatile substance.



2. The pressure increases with decreasing volume at constant temperature, i.e. V $\alpha = \frac{1}{P}$.

Experiment 3.2

- As the flask is heated, the air molecules inside the flask expand and move through the capillary tube and form bubbles in the beaker of water.
- 2. The experiment is an illustration of the increase in the volume of a gas with an increase in temperature.

Experiment 3.3

The experiment demonstration of diffusion of gases is presented as follows:

- 1. Pressure in pot decrease in the case of CO_2 as air diffuses out quicker than carbon dioxide diffuses in.
- 2. Pressure changes were observed by levels of coloured water in the limbs of the U-tube.
- 3. Generally, lighter gases have higher diffusion rate than heavier gases.



As the water in the flask is heated, its molecules escape to create pressure above the surface of the liquid. This is called vapor pressure of the liquid. The vapor pressure of the liquid pushes the liquid to a certain height in the U- tube.



Experiment 3.5

- 1. The temperature rises steadily as the water is heated. The boiling point of water actually depends on altitude and atmospheric pressure.
- 2. Record the thermometer reading while preforming the experiment.
- 3. Once boiling starts, the temperature of the water remains constant (at its boiling point).
- 4. Porcelain chips are added as boiling chips. The purpose of adding boiling chips is to avoid the sudden bumping of the boiling liquid.



Experiment 3.6

The experiment demonstrates phase change, and involves determination of melting point of solids, using melting ice. Relate your observation and analysis with the heat curve (Figure 3.13).

Answer Keys to Self-test Exercises

Self-test Exercise 3.1

- Nitrogen and oxygen can exist as liquid and solid at very low temperatures and high pressures.
- 2. Let students hold discussions in groups about the three states of matter in terms of space between their particles, distance between them, volume and shape, and compressibility and density. 6
- 3. Plasma is different from gas because it is made up of groups of positively and negatively charged particles. For example, in neon gas, its electrons are all bound to the nucleus. But in a neon plasma, its electrons are free to move around the system

Self-test Exercise 3.2

- a. Gas < Liquid < Solid.
- b. Gas < Liquid < Solid.
- c. Solid < Liquid < Gas.
- d. Solid < Liquid < Gas.

Self-test Exercise 3.3

- a. 500 mmHg = 0.66 atm = 500 torr = 50 cmHg
- b. $100 \text{ dm}^3 = 100,000 \text{ mL} = 100,000 \text{ cm}^3 = 100 \text{ L} = 0.1 \text{ m}^3$
- c. 54 °C = 327 K = 129.2 °F

Self-test Exercise 3.4

- 1. 200 kpa
- 2. 0.75 L = 750 mL

Self-test Exercise 3.5

1. Given: Initial conditions:
$$T_1 = -173 \text{ °C} = 100 \text{ K}$$
, $V_1 = V_1$
Final conditions: $T_2 = 27 \text{ °C} = 300 \text{ K}$, $V_2 = V_1$
 $V_2 = T_2 \frac{V_1}{T_1} = 300 \text{ K} \frac{V_1}{100 \text{ K}} = 3V_1$
a. Given: Initial conditions: $T_1 = 17 \text{ °C} = 290 \text{ K}$, $V_1 = V_1$
Final conditions: $T_2 = ?$, $V_2 = \frac{1}{2}V_1$
 $T_2 = \frac{T_1}{V_1} V_2 = \frac{290 \text{ K}}{V_1} \frac{1}{2} \sqrt{T_1} = 145 \text{ K} = -128 \text{ °C}$
b. $V_2 = 2V_1$
 $T_2 = \frac{T_1}{V_1} V_2 = \frac{290 \text{ K}}{V_1} 2V_1 = 580 \text{ K} = 307 \text{ °C}$
c. $V2 = 3V1$
 $T_2 = \frac{T_1}{V_1} V_2 = \frac{290 \text{ K}}{V_1} 3 \sqrt{T_1} = 870 \text{ K} = 597 \text{ °C}$
2. Initial conditions: $T_1 = 25 \text{ C} = 298 \text{ K}$, $V_1 = 1.5 \text{ dm3}$
Final conditions: $T_2 = 100 \text{ °C} = 373 \text{ K}$, $V_2 = ?$
 $V_2 = T_2 \frac{V_1}{T_1} = 373 \text{ K} \frac{1.5 \text{ dm}^3}{298 \text{ K}} = 3V_1 = 1.88 \text{ dm}^3$

Self-test Exercise 3.6

1. Initial conditions: $T_1 = ?$, $P_1 = 1$ atm Final conditions: $T_2 = 250$ K, $p_2 = 1.5$ atm

$$T_1 = T_2 \frac{P_1}{P_1} = 250 \text{ K} \frac{1 \text{ atm}}{1.5 \text{ atm}} = 166.67 \text{ K}$$

2. a. Tire pressure: Automobile tire pressure drops on a cold day and rises on a hot day.

If you put too much air in your tires when they are cold, they could over-pressurize when they heat up. Similarly, if your tires read the proper pressure when they are hot, they will be underinflated when it's cold.

- b. Pressure cooker: Applying heat to a pressure cooker increases the pressure inside the device. Increasing pressure <u>raises the boiling point</u> of water, shortening cooking times. Because the container is sealed, flavors aren't lost to the air with steam.
- c. Aerosol can: The reason you shouldn't store aerosol cans under hot conditions or dispose of them by burning is because heating the can increases the pressure of its contents, potentially causing the can to burst.
- d. Water heater: An electric water heater is a lot like a pressure cooker. A pressure-relief valve prevents steam from accumulating. If the valve malfunctions, heat drives up the steam pressure inside the heater, eventually bursting it.

Self-test Exercise 3.7

1. Given: Initial conditions: $T_1 = 35 \text{ °C} = 308 \text{ K}$, $V_1 = 50 \text{ cm}^3 P_1 = 60.0 \text{ kPa}$ Final conditions: $T_2 = 0 \text{ °C} = 273 \text{ K}$, $P_2 = 1 \text{ atm} = 101.325 \text{ kPa}$, $V_2 = ?$ $V_2 = \frac{T_2}{P_2} \times \frac{P_1 V_1}{T_1} = \frac{273 \text{ K}}{101.325 \text{ kPa}} \times \frac{60 \text{ kPa} \times 50 \text{ cm}^3}{308 \text{ K}} = 26.24 \text{ cm}^3$

2. Given: Initial conditions: $T_1 = 26 \text{ °C} = 299 \text{ K}$, $P_1 = 660 \text{ torr}$, $V_1 = 280 \text{ mL}$ Final conditions: $T_2 = ?$, $P_2 = 940 \text{ torr}$, $V_2 = 440 \text{ mL}$ $T_2 = \frac{P_2 V_2 T_1}{P_1 V_1} = \frac{940 \text{ torr} \times 440 \text{ mL} \times 299 \text{ K}}{660 \text{ torr} \times 280 \text{ mL}} = 669.19 \text{ K} = 396.19 \text{ °C}$

Self-test Exercise 3.8

i. Molecular weight of a gas = density $\left(\frac{g}{L}\right) \times \frac{Volume(L)}{mole}$

= 1.41
$$\frac{g}{L}$$
 × 27 $\frac{L}{mol}$ = 38.07 $\frac{g}{mol}$

ii. At STP, 1 mol of the gas, 38.07 g, would occupy 22.4 L, and its density would be



- $1.62 r_{Kr}$)
- b. N_2 (M = 28 g/mol) and C_2H_2 (26 g/mol): Both of them have almost equal rate of diffusion since they have nearly equal molecular mass.

Evaporation	Boiling
• Evaporation is slow process.	• Boiling is a fast process.
 Evaporation takes place from the surface of the liquid. 	Boiling takes place from all the parts of the liquid
 Evaporation takes place at all temperatures. 	 Boiling takes place at a fixed temperature on heating.

3. At higher altitudes, air pressure is lower. Thus, at high altitude, it takes less energy to bring a liquid to the boiling point. Less energy means less heat, which means a liquid will boil at a lower temperature at a higher altitude

Self-test Exercise 3.12

1. Cooling curve

(£)=



2. Heating curves show how the temperature changes as a substance is heated up. However, cooling curves are the opposite. They show how the temperature changes as a substance is cooled down. Just like heating curves, cooling curves have horizontal flat parts where the state changes from gas to liquid, or from liquid to solid.

	Answer to Self-Assessment Questions				
	Part				
	runi				
1.	Solid	4.	Gas		
2.	Liquid	5.	Gas		
3.	Gas				
	Part II				
6.	Shape	9.	0 K or -273 °C		
7.	Melting points	10.	Avogadro		
8.	Sublimation points				
	Part III				
11.	a	14.	a		
12.	d	15.	a		
13.	d				
	Part IV				
16	6. 0.8 bar	19	. 884.8 К = 111.8 °С		
17	7. 1, 232.7 g/mol	20). 1.92 atm		
18	3. 44.8 L				

References

- 1. Zumdahl, SS & Zumdahl, SA 2007, Chemistry, 7th edn, Houghton Mifflin Company, New York, Boston.
- 2. Petrucci, RH, Herring, FG, Madura, JD & Bissonnette, C 2017, General Chemistry: Principles and Modern Applications, 7th edn, Pearson, Canada, Toronto.
- 3. Brady, JE & Senese, F 2009, Chemistry: Matter and Its Changes, 5th edn, John Wiley & Sons, USA.
- 4. Ebbing, DD & Gammon, SD. 2007, General Chemistry, 11th edn, Cengage Learning, USA, Boston.
- 5. Silberberg MS 2007, Principles of General Chemistry, 1st edn, Mc-Graw-Hill, USA.
- 6. Vining, WJ, Young, SM, Day, R & Botch, B 2016, General Chemistry, Cengage Learning, USA, Boston.
- 7. Zumdahl SS & DeCoste DJ 2010, Introductory Chemistry, 7th edn, rooks/Cole, Cengage Learning, USA.
- 8. McQuarrie, DA., Rock, PA & Gallogly, EB 2011, General Chemistry, 4th edn, University Science Book, Canada, Mill Valley.
- 9. Tro, NJ 2011, Introductory Chemistry, 4th edn, Pearson, USA.
- 10. Chang, R & Overby, J 2008, General Chemistry: The Essential Concepts, 6th edn, McGraw-Hill, New York.